

*VISUALIZED*  
**CHEMISTRY**

**LEMKIN**



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# VISUALIZED CHEMISTRY

INCLUDING  
CHEMISTRY IN WARFARE

BY

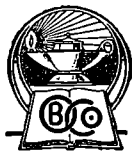
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## PREFACE

THE purpose of VISUALIZED CHEMISTRY is to present the subject-matter of high school chemistry in a manner that is concise, up-to-date, interesting, and well within the grasp of the average student. With this end in view, the author has employed various features and devices which, he believes, should be of material assistance to the student in mastering the subject.

*Conciseness.*—Every effort has been made to achieve a lucid, concise style that will arouse and maintain student interest. Extraneous details have been eliminated, so that attention is focused sharply on the essential facts. In many instances, the text has been arranged in tabular form, thus providing graphic recapitulations of important topics. These qualities make the book ideal for review purposes or for use as a supplementary text in conjunction with any larger book. At the same time, VISUALIZED CHEMISTRY can be used very effectively as a basic text, for it meets fully the latest requirements of the New York State Board of Regents, the College Entrance Examination Board, the American Chemical Society, and the syllabi of the most progressive secondary schools throughout the United States.

*Motivations.*—Each chapter begins with an introductory statement designed to give the student a proper approach to the material to be studied. In chapters containing purely factual material, interest is aroused by compact summaries of the importance and uses of the elements and compounds discussed. The chapters dealing with theories, laws, calculations, etc., are motivated by means of brief descriptions of the manner in which the concepts involved have affected the development of the science.

*Visualizations.*—VISUALIZED CHEMISTRY makes effective use of one of the most basic of all teaching aids—appeal to the eye. An attempt has been made to depart from the routine illustrations found in most textbooks, and to introduce a dramatic element that encourages the student to form a mental *picture* of the material he is studying. This is manifest not only in the unique full-page visualizations, but also in the standard diagrams.

*Timeliness.*—Every topic in VISUALIZED CHEMISTRY is treated so that it includes all the most recent discoveries and concepts which leading authorities regard as fundamentally sound. Chapter 27, "Frontiers of Chemistry," presents an up-to-the-minute survey of the advance of chemistry, both theoretical and applied, and emphasizes the prospects for future progress. Chapter 29 provides thumbnail biographical sketches of distinguished chemists, of our own day and of earlier generations.

*Electron Theory.*—Now regarded as an indispensable aid in teaching high school chemistry, the Electron Theory is explained fully in an early chapter, and is used as an integrating concept throughout the text. The drawings illustrating atomic structure are particularly effective.

*Applied Chemistry.*—The text provides ample material on practical applications of chemistry in industry, agriculture, the home, etc. The relation of the subject to familiar, everyday activities is stressed throughout.

*Exercise Material.*—The thought questions following each chapter have been carefully designed to provide ample drill material, as well as to stimulate independent activity. These questions are supplemented by a variety of objective tests. Complete recent examination papers, set by the New York State Board of Regents, are included.

*Appendix.*—The Appendix contains a variety of valuable information, such as important definitions, chemical tests, lists of alloys and common substances, etc. Much of this material is in tabular form, thus facilitating ready reference.

The author makes grateful acknowledgment to the Editor, Mr. Maurice U. Ames, Chairman of the Department of Physical Sciences, George Washington High School, New York City, who gave most generously of his time, and whose suggestions, criticism and general guidance were invaluable in the preparation of the manuscript and illustrations, and in the reading of the proofs.

W. L.

# CONTENTS

CHAP:	PAGE
1. INTRODUCTION . . . . .	1
2. OXYGEN . . . . .	9
3. HYDROGEN . . . . .	20
4. WATER AND SOLUTION . . . . .	29
5. ATOMS, MOLECULES, FORMULAS, AND EQUATIONS . . . . .	45
6. THE ELECTRON THEORY. . . . .	55
7. THE IONIZATION THEORY . . . . .	65
8. OXIDES, ACIDS, BASES, AND SALTS . . . . .	76
9. THE HALOGENS . . . . .	85
10. THE ATMOSPHERE . . . . .	103
11. COMPOUNDS OF NITROGEN . . . . .	114
12. SULFUR AND SULFUR COMPOUNDS . . . . .	127
13. CARBON AND SOME CARBIDES . . . . .	143
14. OXIDES OF CARBON . . . . .	153
15. LIQUID AND GASEOUS FUELS . . . . .	163
16. ORGANIC CHEMISTRY. . . . .	174
17. METALS AND NON-METALS . . . . .	191
18. SODIUM, POTASSIUM, AND THEIR COMPOUNDS . . . . .	199
19. CALCIUM AND ITS COMPOUNDS . . . . .	210
20. IRON AND ITS COMPOUNDS . . . . .	219
21. ALUMINUM AND ITS COMPOUNDS . . . . .	234
22. COPPER AND ITS COMPOUNDS . . . . .	240
23. OTHER METALS . . . . .	244
24. FOODS . . . . .	253
25. TEXTILES. . . . .	261
26. COLLOIDAL STATE OF MATTER. . . . .	267

CHAP.	PAGE
27. FRONTIERS OF CHEMISTRY . . . . .	273
28. CHEMICAL CALCULATIONS . . . . .	279
29. CHEMISTRY IN WARFARE . . . . .	302
30. BRIEF BIOGRAPHIES . . . . .	321

## APPENDIX

IMPORTANT DEFINITIONS . . . . .	328
GENERAL SOLUBILITY RULES . . . . .	338
CHEMICAL TESTS . . . . .	339
ACTIVITY (ELECTROCHEMICAL) SERIES . . . . .	342
ALLOYS . . . . .	343
COMMON SUBSTANCES . . . . .	344
PHYSICAL CONSTANTS OF COMMON GASES . . . . .	346
THE NINETY-TWO ELEMENTS . . . . .	347
SOME HELPFUL BOOKS FOR FURTHER STUDY . . . . .	348
RECENT EXAMINATION PAPERS . . . . .	349
INDEX . . . . .	vii



# VISUALIZED CHEMISTRY

## CHAPTER 1

### INTRODUCTION

**Definition.**—Chemistry is the science that deals with (1) the different kinds of matter, (2) the properties and uses of matter, (3) the changes which matter undergoes, and (4) the conditions which influence these changes.

**Why Study Chemistry?**—Chemistry plays an important part in the creation and utilization of everything found in our environment. The essentials of life—air, water, food, clothing, and building materials—either are the products of chemical processes or take part in chemical changes which are of vital importance to all of us. Throughout our daily routine, we make contact, in one form or another, with countless applications of chemistry. In fact, during the twenty-four hours of the day, there is scarcely a single object coming within the range of our five senses which does not have some direct association with chemistry. In addition to this material aspect of the science, there is another, and perhaps more important, reason for studying chemistry. It develops the *scientific habit of mind*, which is the attitude of examining problems critically, systematically, and without prejudice. It teaches the value of rejecting preconceived notions and of tackling a question with only one desire—to arrive at the truth.

**Matter.**—Matter is anything that occupies space and has weight. A special kind of matter, such as water, iron, wood, and air, is called a *substance*.

**States of Matter.**—Matter may exist in three states or forms: *solid*, *liquid*, and *gaseous*. The table on page 2 indicates how these three states differ from each other.

Many substances may exist in all three states at different temperatures. For example, water is a liquid under ordinary conditions. When heated, it changes to a gas (*steam*); when cooled, it becomes a solid (*ice*).

SOLID	LIQUID	GAS
Has a definite shape.	Takes the shape of the container.	Has no definite shape.
Has a definite volume.	Has a definite volume.	Expands indefinitely to fill the container.
<i>Examples</i> wood steel glass ice	<i>Examples</i> water alcohol gasoline mercury	<i>Examples</i> air hydrogen ammonia steam

**Properties of Matter.**—The properties of a substance are the characteristics by which we recognize or identify it. Properties are divided into two groups:

**Physical properties** are those characteristics of a substance which can be observed without chemical action, and are readily perceptible by our senses.

*Examples*

Chalk is *white*.  
Sugar is *sweet*.  
Salt is *soluble* in water.  
Diamond is *hard*.  
Copper *conducts* electricity.

**Chemical properties** are those characteristics of a substance which can be determined by observing its action, or lack of action, when in contact with chemical agents.

*Examples*

Hydrogen *burns* in oxygen.  
Nitrogen *does not burn*.  
Acids *neutralize* bases.  
Sodium *reacts* with water.  
Iron *rusts* in moist air.

**Changes in Matter.**—Matter may undergo two kinds of changes.

A **physical change** is any action which does not alter the chemical composition of a substance.

*Examples*

*Breaking* of glass.  
*Melting* of ice.  
*Chopping* of wood.  
*Tearing* of paper.  
*Magnetizing* of iron.

A **chemical change** is any action which alters the composition of a substance and results in the formation of one or more new substances.

*Examples*

*Souring* of milk.  
*Rusting* of iron.  
*Digesting* of food.  
*Burning* of coal.  
*Explosion* of gunpowder.

**Kinds of Matter.**—Substances are classified as *elements*, *compounds*, and *mixtures*.

**Elements.**—An element is a substance which cannot be decomposed into simpler substances by any known means. *Examples:* sulfur, oxygen, carbon, iron, silver, mercury.

**Symbols.**—Each element is represented by a symbol, consisting of one or two letters of the English or Latin name of the element. The table below gives some common elements and their symbols.

ELEMENT	SYMBOL	ELEMENT	SYMBOL
Aluminum	Al	Hydrogen	H
Calcium	Ca	Iron	Fe
Carbon	C	Mercury	Hg
Chlorine	Cl	Oxygen	O
Gold	Au	Sodium	Na

**Classification of Elements.**—Elements may be conveniently grouped into two general classes, *metals* and *non-metals*. The table below summarizes the important physical and chemical properties of these two classes.

	METALS	NON-METALS
<b>Physical Properties</b>	Lustrous. Good conductors of heat and electricity. Malleable and ductile. Generally high density. Solids at ordinary temperatures (except mercury).	Non-lustrous. Poor conductors of heat and electricity. Not malleable or ductile. Generally low density. Solids and gases at ordinary temperatures (except bromine).
<b>Chemical Properties</b>	Combine with non-metals. Form bases. Form positive ions (page 66). Replace hydrogen of acids. Lend electrons (page 58).	Combine with metals. Form acids. Form negative ions.  Do not replace hydrogen of acids. Borrow electrons.
<b>Examples</b>	Iron Lead Copper	Sulfur Chlorine Oxygen

**Compounds.**—A compound is a substance consisting of two or more elements chemically united, that is, combined in such a manner that each element loses its characteristic properties. Every compound has its own individual properties by which it

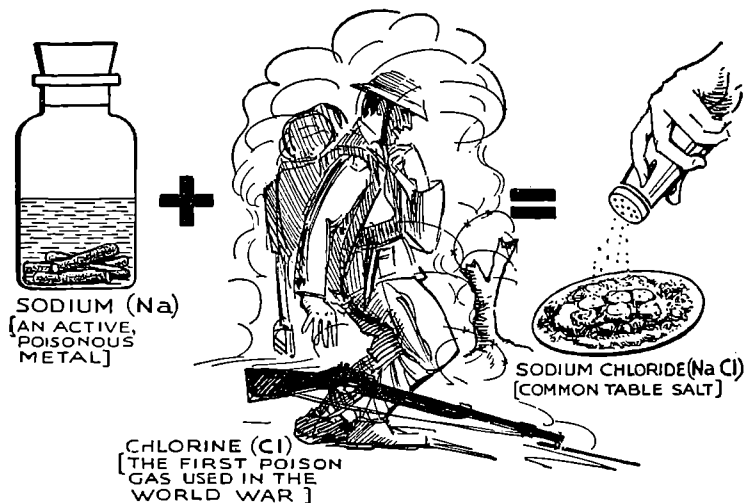


FIG. 1. Two active and poisonous elements combine, forming a stable and harmless compound.

may be distinguished from other compounds. It always contains the same elements in the same proportion by weight (*Law of Definite Proportions*). Its composition may be represented by a formula, as indicated in the table below.

COMPOUND	ELEMENTS PRESENT	FORMULA
Water	Hydrogen and oxygen	H <sub>2</sub> O
Calcium oxide	Calcium and oxygen	CaO
Sodium chloride	Sodium and chlorine	NaCl
Sulfuric acid	Hydrogen, sulfur, and oxygen	H <sub>2</sub> SO <sub>4</sub>

**Mixtures.**—A mixture is a substance consisting of two or more elements or compounds that are *not* chemically united. *Examples:* air, soil, flour, petroleum, talcum powder.

**Comparison of Mixtures and Compounds.**—The table below summarizes the chief points of difference between a mixture and a compound.

MIXTURE	COMPOUND
Ingredients are not chemically united.	Ingredients are chemically united.
Ingredients retain their individual properties.	Ingredients do not retain their individual properties.
Can be separated into its ingredients by mechanical means.	Can be separated into its ingredients only by chemical action.
Has no definite composition by weight.	Has a definite composition by weight.

**Law of Conservation of Matter.**—*Matter can neither be created nor destroyed, but it may be changed from one form into another.* Expressed differently, the total weight of a substance after a chemical action is exactly the same as it was before. This fundamental truth was first demonstrated by Lavoisier. He proved that when a substance burns in air it combines with oxygen and

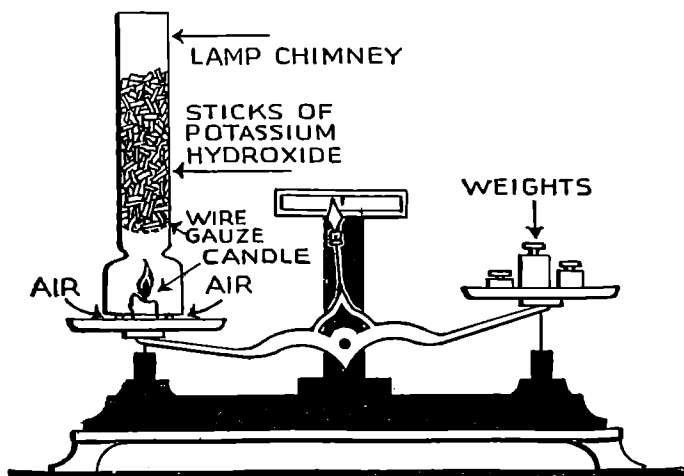


FIG. 2. When a candle burns, it combines with the oxygen of the air, giving off gaseous products. If these products are absorbed by potassium hydroxide, as shown above, the total weight on the pan at the left will become greater than that at the right, because of the addition of the combined oxygen.

becomes heavier by the weight of the oxygen with which it has combined. Fig. 2 illustrates a simple experiment proving that the products formed when a candle burns weigh more than the candle itself, the additional weight being due to the oxygen which unites with the ingredients of the candle.

**Energy.**—Energy is the ability to do work. The table below lists the six common forms of energy.

FORM OF ENERGY	ILLUSTRATED BY
Heat	Burning coal
Mechanical	Steam engine
Electrical	Dynamo
Chemical	Storage battery
Light	Electric bulb
Sound	Electric bell

**Law of Conservation of Energy.**—*Energy can neither be created nor destroyed, but it may be changed from one form into another.* Fig. 3 illustrates various transformations of energy. Throughout

HEAT → MECHANICAL → ELECTRICAL → LIGHT AND CHEMICAL

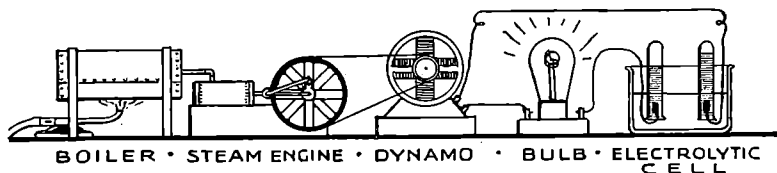


FIG. 3. Energy may be changed from one to another of its various forms. In these transformations, the total amount of energy remains constant. This illustrates the *Law of Conservation of Energy*.

the changes which take place we are dealing with a *constant* quantity of energy. There is no gain or loss—only transformation.

### QUESTIONS

1. Define: (a) chemistry, (b) matter, (c) energy.
2. State two reasons why an intelligent person should have a knowledge of chemistry.
3. Explain how one can change (a) a liquid to a gas, (b) a liquid to a solid, (c) a gas to a solid.

4. (a) Distinguish between a physical property and a chemical property. (b) Give two examples of each.

5. State some simple physical or chemical property by which you could identify each of the following substances: salt, glass, chalk, butter, rubber, gold, sulfur, gasoline, cork, lead.

6. Describe experiments by which you could show the difference between a physical change and a chemical change.

7. Classify the following as physical or chemical changes: (a) evaporation of water, (b) digestion of food, (c) fermentation of sugar, (d) breaking of glass, (e) melting of ice, (f) baking of bread, (g) rusting of iron, (h) peeling of potatoes, (i) boiling of potatoes, (j) tarnishing of bronze.

8. Classify the following substances as elements, compounds or mixtures: oxygen, gasoline, sulfur, sugar, gunpowder, iron, air, water, glass, diamond.

9. Why is it difficult to prove the Law of Conservation of Matter in the case of a burning candle?

10. Trace the steps in the transformation of energy from coal to the heat of an electric toaster.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Matter occupies . . . . . and has . . . . .

2. Three physical states of matter are . . . . .

3. We recognize a substance by observing its . . . . . and . . . . . properties.

4. A mixture can be separated into its components by . . . . . means.

5. A compound can be separated into its constituents only by . . . . . means.

6. Three forms of energy are . . . . .

7. Dissolving salt in water is a . . . . . change.

8. In general, metals are . . . . . (good, poor) conductors of heat and electricity.

9. The Law of Conservation of Matter states that matter can neither be . . . . . nor . . . . .

10. The true nature of burning was first demonstrated by . . . . .

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

A	B
( ) Water	1. Has no definite size or shape
( ) Light	2. Chemical property
( ) Copper	3. Form of energy
( ) A gas	4. Conducts electricity
( ) Carbon burns in air	5. Mostly solids
( ) Tearing paper	6. Rusting of iron
( ) Kerosene	7. Physical property
( ) Density	8. Symbol
( ) Metals	9. Compound
( ) Cl	10. Sulfur
	11. Mixture
	12. Physical change



## CHAPTER 2

### OXYGEN

**Importance.**—Of the ninety-two elements known to man, there is not a single one that is more important to his existence than oxygen. He breathes it every minute of the day and night, and is thus enabled to carry on his life processes. He uses it to help burn his fuels, for cooking, warmth, and generating power to operate his machinery. He depends on oxygen to decay organic waste materials, destroy germs, and purify water. In hospitals, submarines, mines, and airplanes, he puts oxygen to vitally important uses, where ordinary air would be inadequate for the emergencies which arise.

**Occurrence.**—Oxygen is the most abundant element. It occurs both free and combined with other elements.

1. In the *free state*, oxygen is found chiefly in the atmosphere, which is made up of approximately 21% oxygen and 78% nitrogen by volume, together with small amounts of other gases.

To a limited extent, free oxygen occurs also in water, since the pressure of the atmosphere on the vast bodies of water covering the earth's surface forces a small quantity of air into solution.

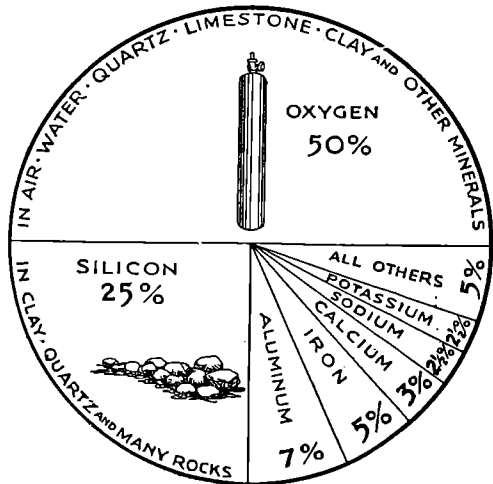
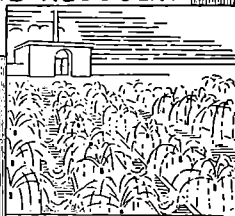
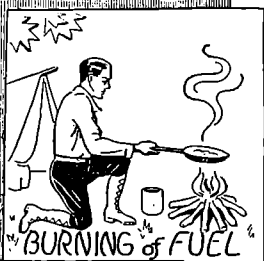


Fig. 4. The composition of the earth's crust.

# OXYGEN and its USES



RESPIRATION IN MAN-ANIMAL AND PLANT

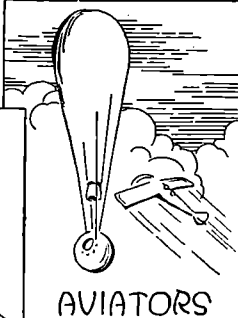
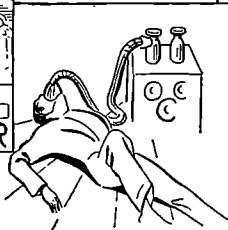


FIG. 5.

2. In the *combined state*, oxygen is found in water, in plant and animal substances, and in the earth. Together with carbon, hydrogen, and other elements, oxygen enters into the composition of thousands of compounds such as starch, cellulose, sugar, fat, protein, etc. Enormous quantities of oxygen are locked up in sand, clay, granite, limestone, and hundreds of other rocks and minerals. In fact, about 50% of the earth's crust is oxygen (Fig. 4).

### Preparation of Oxygen

1. *Decomposition of Mercuric Oxide (Historical Method).*—In 1774, Joseph Priestley, an English chemist and minister, heated red oxide of mercury in a glass tube and obtained a gas which caused a burning candle to burn more brightly than it did in air. At the same time, he discovered that metallic mercury was deposited on the inside of the tube. This method of preparing oxygen is still used as a laboratory demonstration. The chemical action that takes place may be expressed by the following equation:

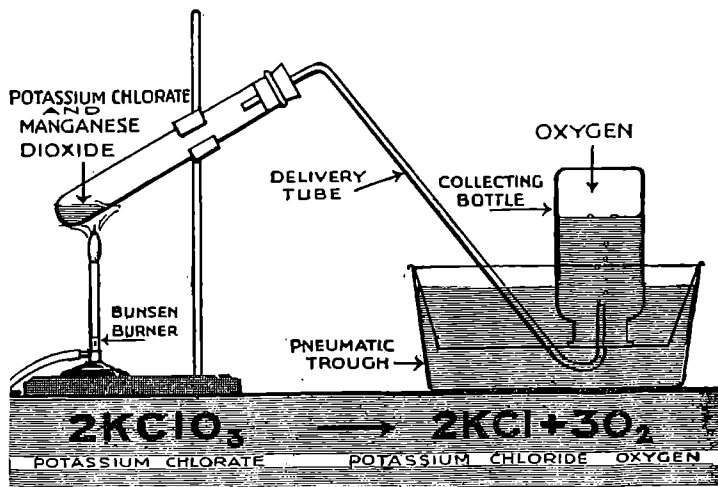
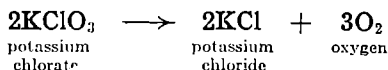


FIG. 6. Laboratory preparation of oxygen.

2. *Decomposition of Potassium Chlorate (Laboratory Method).*—Oxygen is readily prepared in the laboratory by gently heating a

mixture of potassium chlorate ( $\text{KClO}_3$ ) and manganese dioxide ( $\text{MnO}_2$ ). The gas is collected by displacement of water as shown in Fig. 6. The equation for the reaction is:



The manganese dioxide does not appear in the above equation because it does not enter into the reaction. Its function is to help generate the oxygen faster and at a lower temperature. In doing so, the manganese dioxide is itself not permanently changed. Such a substance is called a *catalytic agent* or a *catalyst*.

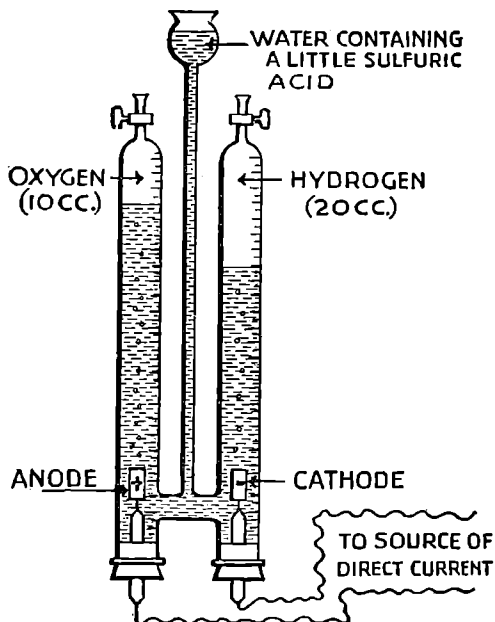
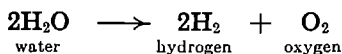


FIG. 7. Hoffman apparatus for the electrolysis of water.

3. *Electrolysis of Water (Commercial Method).*—If a little sulfuric acid is added to water to make it electrically conductive, and a current is passed through the solution, the water is decomposed into oxygen and hydrogen. The oxygen collects over

the positive electrode (*anode*), while the hydrogen collects over the negative electrode (*cathode*). The volume of hydrogen thus obtained is always twice as great as the volume of oxygen. The Hoffman apparatus (Fig. 7) is a convenient device for demonstrating this process of electrolysis. The reaction is:

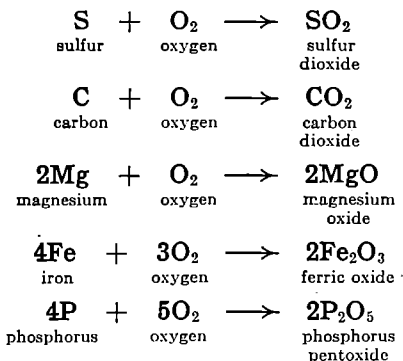


4. *Distillation of Liquid Air (Commercial Method).*—Air may be converted into a liquid by subjecting it to a high pressure and a low temperature. When liquid air is allowed to boil, the nitrogen, having a lower boiling point, escapes first. This leaves nearly pure oxygen.

### Properties of Oxygen

1. *Physical Properties.*—Oxygen is a gas under ordinary conditions of temperature and pressure. It has no color, odor, or taste, is slightly heavier than air, and is only sparingly soluble in water. It can be liquefied and solidified by subjecting it to an extremely low temperature and a high pressure.

2. *Chemical Properties.*—Oxygen does not burn, but it supports the combustion of other substances with great vigor. It is very active chemically, combining with nearly all elements to form *oxides*:



**Oxidation.**—Oxidation is the chemical union of a substance with oxygen. There are two varieties of oxidation, depending on the speed of the action and, therefore, on the manner in which the resulting energy is released.

1. *Slow oxidation* is oxidation which goes on so slowly as to produce no light and no noticeable heat. The rusting of iron and the rotting of wood are examples.

2. *Rapid oxidation* is oxidation which goes on so rapidly as to produce light and noticeable heat. The burning of wood and the explosion of gunpowder are examples.

It is difficult to detect the heat evolved during slow oxidation, because it is soon dissipated to the surrounding matter. However, careful measurements show that the amount of heat given off by a substance undergoing oxidation is always the same, regardless of whether the oxidation is slow or rapid.

**Combustion.**—Combustion is any chemical action producing heat and light. It is usually a rapid oxidation, although oxygen is not the only gas that supports combustion. For example, sodium burns vigorously in chlorine, producing sodium chloride.

**Kindling Temperature.**—The kindling temperature of a substance is the lowest temperature at which the substance begins to burn. For example, yellow phosphorus has an extremely low kindling temperature (36°C.), and must therefore be kept under water. On the other hand, coal has a relatively high kindling temperature. In making a fire in a stove we start by first igniting paper which has a fairly low kindling temperature. The heat generated is sufficient to raise the wood to its kindling temperature, which, in turn, heats the coal until the kindling temperature of that substance is reached.

**Spontaneous Combustion.**—Spontaneous combustion is active burning caused by the gradual accumulation of heat that results from slow oxidation. As the substance oxidizes slowly, a small amount of heat is generated. If this heat is not permitted to escape, it gradually raises the temperature of the substance until the kindling temperature is reached. At this point the substance begins to burn. Some cases in which spontaneous combustion may occur are:

1. Heaps of oily or paint-covered rags left in a poorly ventilated closet.
2. "Green" hay stored in a barn before it is thoroughly dry.
3. Soft coal heaped in large piles.

**Oxidation and Its Control.**—Oxidation, although usually of great benefit to us, may sometimes be destructive. The table below summarizes how we may control various types of undesirable oxidation.

TYPE OF OXIDATION	NECESSARY CONDITIONS	METHODS OF STOPPING OR PREVENTING
<b>Fire</b> (rapid oxidation)	A combustible material (fuel).  Air or oxygen.  Sufficient heat to raise the fuel to its kindling temperature.	Remove or scatter the material.  Shut off the supply of air or oxygen by the use of water, carbon dioxide or some other non-combustible substance.  Cool below the kindling temperature by the use of water.
<b>Decay</b> (slow oxidation of organic matter)	Animal or vegetable matter. Oxygen. Moisture. Bacteria. Warmth.	Prevent the growth of bacteria. Foods are preserved by canning, refrigeration, drying, smoking and salting.
<b>Rusting</b> (slow oxidation of a metal)	A metal. Oxygen of the air, sometimes aided by carbon dioxide and moisture.	Exclude air and moisture from the metal. This may be accomplished by oiling, painting, varnishing, enameling, plating, galvanizing, or tinning.

**Uses of Oxygen.**—In the diluted form in which it exists in the air, oxygen has a number of important uses, associated principally with various natural processes. In a pure or nearly pure state, oxygen is employed for many purposes in which increased chemical activity is required. The table on page 16 summarizes these uses.

## USES OF OXYGEN IN THE AIR

1. *Respiration*.—Animals take in oxygen to oxidize their food, and return carbon dioxide to the air. Plants do likewise to a limited extent.

2. *Burning*.—Coal, wood, gasoline, illuminating gas and all other fuels need oxygen in order to burn.

3. *Aëration*.—Bacteria in drinking water are killed by spraying the water into the air. The oxygen dissolves in the water and destroys the bacteria.

4. *Decay*.—Certain bacteria, assisted by the oxygen of the air, destroy waste animal and vegetable matter which, if allowed to accumulate, would endanger health.

## USES OF PURE OXYGEN

1. *High Temperatures*.—By means of oxy-hydrogen and oxy-acetylene torches (Fig. 8), temperatures as

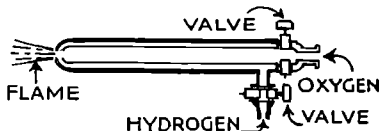


FIG. 8. Oxy-hydrogen torch.

high as 4000°C. are produced for the purpose of cutting and welding steel. Recently, mixtures of liquid oxygen and gasoline have been used to shoot experimental rockets into the upper atmosphere.

2. *Artificial Respiration*.—Sometimes ordinary air is inadequate for breathing, as in cases of pneumonia, strangulation, drowning and carbon monoxide poisoning. At such times pure, or nearly pure, oxygen is administered.

3. *High Altitudes*.—Aviators and mountain climbers carry oxygen tanks and masks to enable them to breathe in rarefied atmospheres.

4. *Emergencies*.—In fires, mine disasters, etc., oxygen helmets are used by rescue workers.

## OZONE

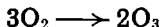
**Allotropism**.—Several elements, notably oxygen, phosphorus, sulfur, and carbon, exist in two or more forms which have different physical and chemical properties. These are known as *allotropic forms* of the element. The variation in properties is believed to be due to a difference in the amount of energy which each form possesses. In the case of an element such as sulfur, which exists in several crystalline forms (page 128), the differences may be due to the arrangement of the atoms in different space patterns.



**Allotropic Forms of Oxygen.**—Two forms of oxygen are known: (1) *ordinary oxygen*, having two atoms in a molecule ( $O_2$ ); and (2) *ozone*, having three atoms in a molecule ( $O_3$ ).

**Occurrence.**—Ozone occurs in small quantities in the upper strata of the atmosphere. Its presence in the lower atmosphere is often detected immediately after a thunderstorm.

**Preparation.**—When a high-voltage electric discharge is passed through dry oxygen, some of the gas is converted into ozone. The change is accompanied by a contraction in volume, since three volumes of oxygen yield only two volumes of ozone. The equation for the reaction is:



### Properties

1. *Physical Properties.*—Ozone is a slightly bluish gas having a strong, pungent odor. It is one and one-half times as heavy as oxygen, and is more soluble in water than oxygen.

2. *Chemical Properties.*—Because of its greater energy content, ozone is far more active chemically than oxygen. Substances like silver, rubber, etc., on which oxygen has no effect, are easily oxidized by ozone. It is an effective bleaching agent as well as a powerful germicide. It is very unstable, changing readily to oxygen, a chemical action which is accompanied by a liberation of energy.

**Uses.**—Ozone is used to bleach wood pulp, textile fibers, oils, and waxes; to purify the water supplies of large cities; and to dry varnishes, paints, and oils. It is also effective in removing foul odors from kitchens, cold storage rooms, and slaughter houses.

### QUESTIONS

1. Name the materials used in the laboratory preparation of oxygen and give the purpose of each.

2. Discuss this statement: "If an atmosphere containing approximately one-fifth, or 20%, oxygen is healthful, then an atmosphere containing five-fifths, or 100%, oxygen should be five times as healthful."

3. Define: (a) *oxidation*, (b) *combustion*, (c) *kindling temperature*, (d) *decay*, (e) *catalytic agent*.

4. Why does blowing on a burning match put out the flame, while blowing on a small fire makes it burn more vigorously?
5. If a fish is placed in water which has previously been boiled and cooled, it will soon die. Explain.
6. What is the most important chemical property of oxygen? Give reasons for your answer.
7. The coal in the hold of a ship suddenly starts to burn. (a) What is this action called? (b) State the conditions which are responsible for it.
8. (a) What is the difference between slow oxidation and rapid oxidation? (b) Give one example of each.
9. State the three principles involved in extinguishing a fire and give one practical illustration of each.
10. (a) Define *allotropic forms*. (b) What theory accounts for the different properties of oxygen and ozone?

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. When sulfur burns in air, ..... is formed.
2. The lowest temperature at which a substance begins to burn is called its .....
3. Oxygen was discovered by ..... in the year ..... when he heated .....
4. Three uses of oxygen are .....
5. An element which has a very low kindling temperature is .....
6. In order that burning may take place, the following three essential conditions must be present: .....
7. A substance which changes the speed of a chemical action without itself being changed is called a .....
8. Three essentials necessary for the decay of organic matter are .....
9. When liquid air boils, the ..... escapes first.
10. The earth's crust consists of approximately ..... per cent oxygen by weight.

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A                       | B                                   |
|-------------------------|-------------------------------------|
| ( ) Manganese dioxide   | 1. Allotropic form of oxygen        |
| ( ) Oily rags           | 2. Galvanizing                      |
| ( ) Ozone               | 3. Most abundant element            |
| ( ) Decay               | 4. Rapid oxidation                  |
| ( ) Eight-ninths oxygen | 5. Catalytic agent                  |
| ( ) Prevents rust       | 6. Potassium chlorate               |
| ( ) 4000°C.             | 7. Oxy-acetylene torch              |
| ( ) Oxygen              | 8. Commercial preparation of oxygen |
| ( ) Liquid air          | 9. Action of bacteria               |
| ( ) Explosion           | 10. Electrolysis                    |
|                         | 11. Spontaneous combustion          |
|                         | 12. Water                           |

## CHAPTER 3

### HYDROGEN

**Importance.**—Hydrogen holds the unique distinction of being the lightest element in the entire family of ninety-two. This lightness makes it the ideal gas for lifting balloons—ideal, except for its extreme combustibility, which has caused the builders of modern dirigibles to turn to another element (helium) for buoyancy with safety. The combustibility of hydrogen, however, makes it an important fuel, in itself, as well as an essential constituent of most of the gaseous and liquid fuels used today. A newly developed use of hydrogen, for obtaining extremely high temperatures, involves the conversion of the element to the atomic form and its subsequent recombining into molecules. The strong attraction of hydrogen for oxygen enables it to remove that element from certain of its compounds. Quite recently, with the magic of catalytic action, hydrogen has come into widespread use for converting cheap or waste substances, like fish oil and low grade coal, into highly valuable products.

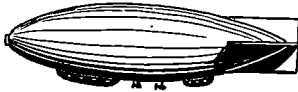
**Occurrence.**—Hydrogen is not nearly as abundant as oxygen, constituting only about 1% of the earth's crust. It differs from oxygen, also, in that it occurs on the earth, in the free state, only in minute quantities. In the combined state, however, hydrogen occurs everywhere and in great abundance.

1. In the *free state*, hydrogen occurs in natural gas, and in volcanic gases. Traces of hydrogen are found in the lower strata of the earth's atmosphere, and, to a much larger extent, at altitudes of twenty miles or more. Hydrogen is known to occur abundantly in the sun's atmosphere.

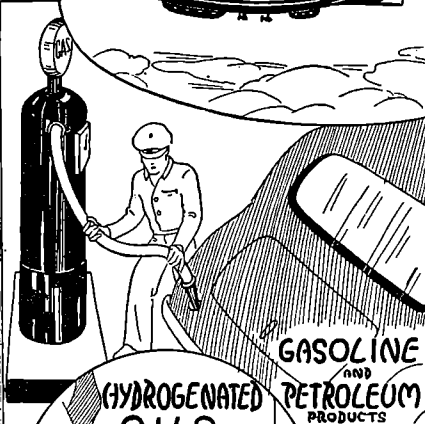
2. In the *combined state*, hydrogen occurs in water to the extent of one-ninth by weight. All acids contain hydrogen, and owe their characteristic properties to this element. Animal and vegetable substances such as butter, starch, sugar, proteins, etc., contain hydrogen in combination with carbon, oxygen, and other elements. Hydrocarbons, such as kerosene, gasoline, and other petroleum products, consist of hydrogen combined with carbon.

# HYDROGEN & its Uses

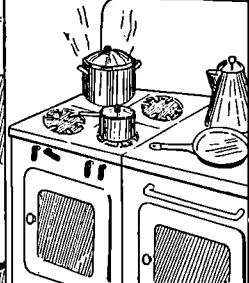
AIRSHIPS AND



BALLOONS

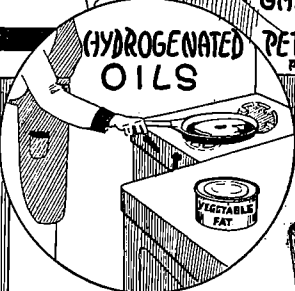


FUEL GASES



HYDROGENATED  
OILS

GASOLINE  
AND  
PETROLEUM  
PRODUCTS



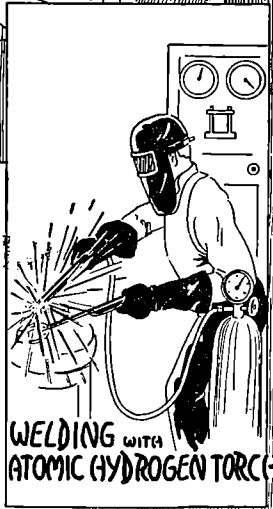
VEGETABLE  
FAT



WATER



SOME  
IMPORTANT  
HYDROGEN  
COMPOUNDS



WELDING WITH  
ATOMIC HYDROGEN TORCH

FIG. 9.

### Preparation from Water

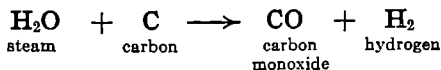
1. *By Electrolysis.*—When an electric current is passed through water to which a little sulfuric acid has been added to make it electrically conductive, hydrogen is liberated at the negative electrode (*cathode*). Pure hydrogen is prepared commercially by this process (Fig. 7).

2. *By the Action of Certain Metals.*—Various metals react with water under certain conditions, liberating hydrogen and forming either a hydroxide or an oxide of the metal. The following table summarizes these reactions.

EQUATIONS		CONDITIONS OF REACTION	SPEED OF REACTION		
2K potassium	+ 2H <sub>2</sub> O water	→ 2KOH potassium hydroxide	+ H <sub>2</sub> ↑ hydrogen	Ordinary temperature	Very rapid
2Na sodium	+ 2H <sub>2</sub> O water	→ 2NaOH sodium hydroxide	+ H <sub>2</sub> ↑ hydrogen	Ordinary temperature	Rapid
Ca calcium	+ 2H <sub>2</sub> O water	→ Ca(OH) <sub>2</sub> calcium hydroxide	+ H <sub>2</sub> ↑ hydrogen	Ordinary temperature	Slow
Mg magnesium	+ 2H <sub>2</sub> O water	→ Mg(OH) <sub>2</sub> magnesium hydroxide	+ H <sub>2</sub> ↑ hydrogen	Boiling temperature	Slow
3Fe iron	+ 4H <sub>2</sub> O water	→ Fe <sub>3</sub> O <sub>4</sub> magnetic iron oxide	+ 4H <sub>2</sub> ↑ hydrogen	Water in form of steam; iron in red-hot state	Rapid
Cu copper	+ H <sub>2</sub> O water	→ no reaction		—	—

NOTE. These reactions demonstrate the order of activity of the metals. (See *Electromotive Series*, or *Activity Series of Metals*, page 194.)

3. *By the Action of Hot Coke on Steam (Commercial Method).*—When steam is passed over white-hot coke (carbon), a mixture of hydrogen and carbon monoxide gas is formed:



When the mixture of gases is cooled to a very low temperature, the carbon monoxide changes to a solid, leaving pure hydrogen.

**Preparation from Acids (Laboratory Method).**—Such metals as zinc and iron react with certain dilute acids, notably hydrochloric acid and sulfuric acid (Fig. 10). In these reactions, the

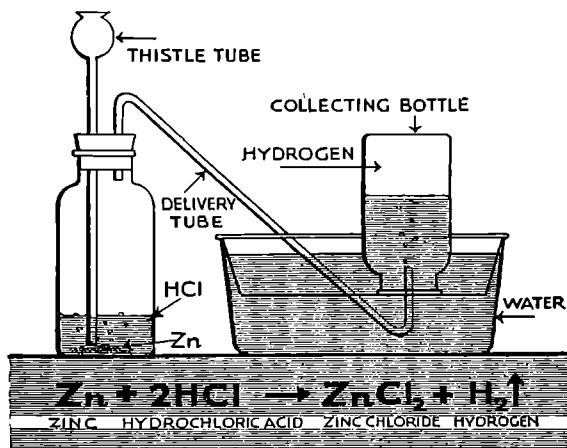
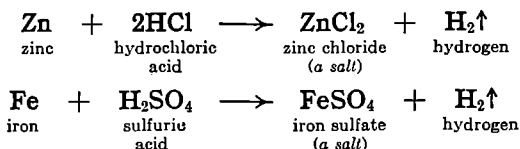


FIG. 10. Laboratory preparation of hydrogen.

metal replaces the hydrogen of the acid and liberates it as a gas, at the same time forming a new compound called a *salt*:



All metals and all acids cannot be employed to generate hydrogen. Metals like gold, silver and copper are not active enough to replace the hydrogen of acids. Nitric acid ( $\text{HNO}_3$ ) and hot concentrated sulfuric acid ( $\text{H}_2\text{SO}_4$ ) are strong oxidizing agents; hence, in the reaction of these acids with metals, any hydrogen formed is immediately oxidized to water.

**Physical Properties.**—Hydrogen is a colorless, odorless, tasteless gas, and is very slightly soluble in water. The outstanding property of hydrogen is that it is the lightest gas known, being

16 times lighter than oxygen, and 14.5 times lighter than air. Hydrogen diffuses with extreme rapidity; therefore it is difficult to keep it in containers. The metals palladium and platinum have the power to absorb hydrogen in large volumes. This absorption (*occlusion*) is accompanied by the liberation of so much heat that the metals glow. Certain types of self-lighting gas burners and cigar lighters utilize this principle.

**Chemical Properties.**—Hydrogen burns with a very hot, pale-blue flame, but does not support combustion. The equation for the burning of hydrogen is:

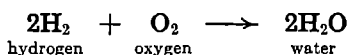


Fig. 11 illustrates the method used in the laboratory to demonstrate that water is a product of this reaction. When hydrogen is

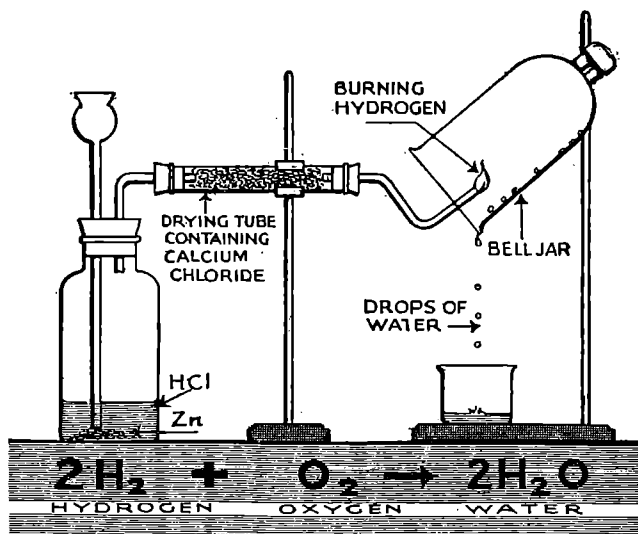
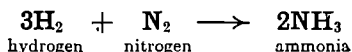
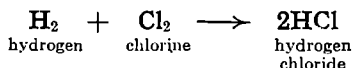


FIG. 11. Water is the only product formed when hydrogen burns in air.

mixed with air or oxygen, it explodes on the application of a flame or spark. In such an explosive mixture, the amount of hydrogen



may vary anywhere between 4% and 75% by volume. Hydrogen combines directly with other elements at high temperatures:



Hydrogen acts as a powerful reducing agent by removing oxygen from many of its compounds.

**Reduction.**—Reduction is the process of removing oxygen from a compound. A substance which can effect such a removal of oxygen is called a *reducing agent*. To prove that hydrogen is a reducing agent, we pass the dry gas over heated copper oxide, as in Fig. 12. The hydrogen combines with the oxygen present in

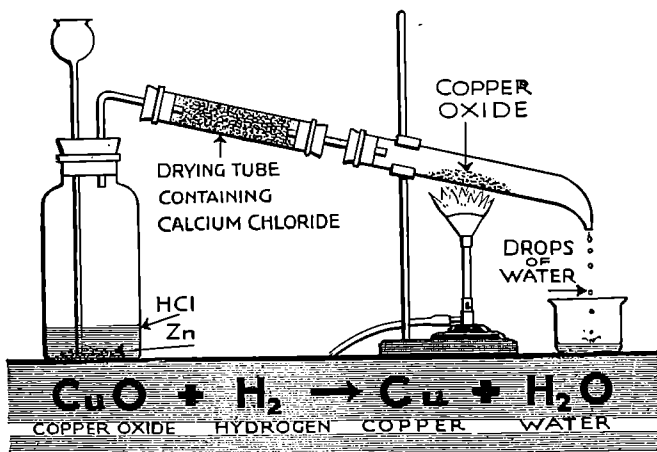
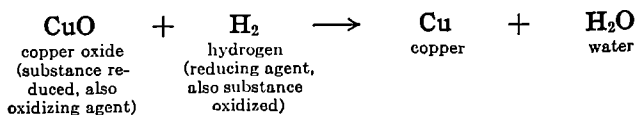


FIG. 12. Reduction of copper oxide by hydrogen.

the oxide, forming water. The residue left in the tube is metallic copper. Whenever reduction takes place, it is always accompanied by oxidation, as can be observed by examining the equation for this reaction:



Carbon, in the form of coke, is the most commonly used commercial reducing agent. It is employed chiefly to remove oxygen from iron ore and other metallic oxide ores. Since carbon is a solid, and is cheap, it is a more practical reducing agent for such purposes than hydrogen.

**Test for Hydrogen.**—Various other gases, notably carbon monoxide, are also colorless and burn with an almost invisible bluish flame. However, hydrogen may be identified by one unmistakable chemical proof—when it burns, the sole product formed is water.

### Uses of Hydrogen

1. *Balloons and Airships.*—Hydrogen, being the lightest gas known, is used for buoyancy in so-called “lighter-than-air” craft. Owing to its great combustibility, however, it has largely been superseded, at least in this country, by helium. This gas, although it is twice as heavy as hydrogen, has about 92% of its lifting power, and has at the same time the decided advantage of being non-combustible. In May, 1937, the destruction by fire of the *Hindenburg*, the world's largest airship, served as a final, tragic indictment of the use of hydrogen in commercial or military dirigibles.

2. *Welding.*—By the use of the oxy-hydrogen torch (Fig. 8), a temperature of about 2400°C. can be obtained. Recently, there has been developed an electric welding torch using monatomic hydrogen. A stream of hydrogen molecules is converted into atoms by the high temperature of an electric arc. The hydrogen atoms recombine into molecules, a process accompanied by the evolution of a great deal of heat. This heat, together with the heat produced by the burning hydrogen molecules, gives a temperature close to 5000°C., the highest obtained so far by artificial means.

3. *Fuel Gases.*—Because it produces such intense heat when burning, hydrogen is a valuable constituent of some gaseous fuels, such as water gas and producer gas.

4. *Hydrogenation.*—Many liquid oils and fats, such as cottonseed oil and cocoanut oil, are converted into solids by being made to combine with hydrogen under the influence of a catalyst (usually finely divided nickel). Most of the solid shortening materials now on the market are made in this manner. In a similar way, fish oil and other waste oils are hydrogenated to form solid fats for soap-making. Recently, a process has been developed for hydro-

generating low-grade coal and petroleum to produce valuable gas-line substitutes.

5. *Preparation of Ammonia.*—Large quantities of ammonia are now being manufactured by the direct combination of hydrogen with the nitrogen of the air.

6. *Reduction.*—Hydrogen is used as a reducing agent in the laboratory. To a limited extent, it is also used for this purpose commercially.

### QUESTIONS

1. What substances would you use to prepare hydrogen by (a) replacement, (b) reduction, (c) electrolysis? Write an equation for each method.

2. In collecting hydrogen by the displacement of air, is the bottle held mouth upward or mouth downward? Explain.

3. Why would you not expect free hydrogen to occur in nature?

4. What happens when a lighted splint is inserted into a bottle of (a) hydrogen, (b) oxygen, (c) air, (d) a mixture of hydrogen and oxygen, (e) a mixture of hydrogen and air?

5. (a) Give three properties of hydrogen. (b) For each property mentioned, state one use of hydrogen depending on this property.

6. What is *reduction*? In the reaction between copper oxide and hydrogen, name (a) the reducing agent, (b) the substance reduced, (c) the oxidizing agent, (d) the substance oxidized.

7. Why is helium superior to hydrogen for use in dirigibles?

8. (a) Why is hydrogen a good fuel? (b) Why is it not suited for illuminating purposes?

9. In order to prepare hydrogen by the reaction between a metal and sulfuric acid, would you use zinc or copper? Explain.

10. (a) Name a formerly useless substance which is now being changed to a useful product by the action of hydrogen. (b) Explain how this change is accomplished.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. The most abundant source of hydrogen is . . . . ., where it occurs in the . . . . . (free, combined) state.

2. When hydrogen burns, it forms . . . . .

3. Hydrogen and . . . . . form an explosive mixture.

4. The most important physical property of hydrogen is .....
5. (a) Zinc + hydrochloric acid  $\longrightarrow$  ..... + .....
- (b) Potassium + water  $\longrightarrow$  ..... + .....
- (c) Copperoxide + hydrogen  $\longrightarrow$  ..... + .....
6. A reducing agent removes ..... from a compound.
7. The process which always accompanies reduction is called .....
8. Gasoline consists of compounds which are made up of the two elements .....
9. Vegetable oils are converted into solid fats by a process called .....
10. Two acids which do not yield hydrogen when reacting with metals are .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A                             | B                                       |
|-------------------------------|---|
| ( ) Petroleum                 | 1. Occlusion                            |
| ( ) Steam + red-hot iron      | 2. Helium                               |
| ( ) Finely divided nickel     | 3. Atmosphere of the sun                |
| ( ) Monatomic hydrogen torch  | 4. Test for hydrogen                    |
| ( ) Potassium                 | 5. Mixture of hydrocarbons              |
| ( ) Contains free hydrogen    | 6. Commercial reducing agent            |
| ( ) Hydrogen + nitrogen       | 7. Catalyst in hydrogenation<br>of oils |
| ( ) Platinum absorbs hydrogen | 8. 4000°C.                              |
| ( ) Carbon                    | 9. Copper oxide                         |
| ( ) Burns to form water       | 10. Reacts violently with water         |
|                               | 11. Commercial preparation of ammonia   |
|                               | 12. Commercial preparation of hydrogen  |

## CHAPTER 4

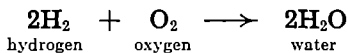
### WATER AND SOLUTION

**Importance of Water.**—Water is our most familiar, most abundant, and most useful chemical compound. It covers vast portions of the earth's surface, and is present in a host of the commonest substances. In vapor form, it determines climate and daily weather changes. As rain, snow, and other forms of atmospheric precipitation, it influences soil fertility, the habitability of regions, and indeed, the character and trend of entire civilizations. It forms a vital part of all plant and animal matter, of all foods, and of the human body. As one of the agents of erosion, water affects the very face of our planet. Without water, life on earth would be impossible. Indeed, it has been said that the chemistry of our world is a "water chemistry."

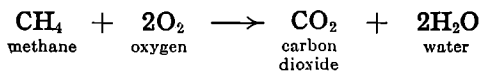
**Occurrence.**—Water occurs in nature in all three states, solid, liquid, and gas, usually free as the compound  $H_2O$ , but also loosely combined with other compounds. Bodies of water such as rivers, lakes, oceans, etc., cover three-quarters of the earth's surface. In some places the depth of this layer of water is five or six miles. Water is found in the air as vapor. It occurs also in many rocks and mineral substances. All animal and vegetable matter contains water. Lean meat, for example, is about 60% water, while cucumbers and tomatoes are about 95% water. The human body contains about 70% water.

**Formation.**—Water, being so abundant, is not prepared from other materials. However, it is a product of many chemical reactions, some of which follow:

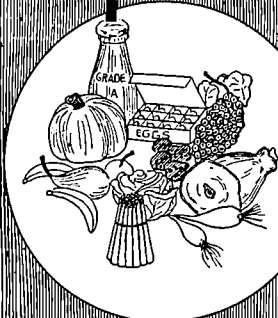
1. *Direct union of oxygen and hydrogen, by burning:*



2. *Oxidation of a compound of hydrogen:*



# Importance & USES OF Water



FOODS CONTAIN AS HIGH AS 90% WATER



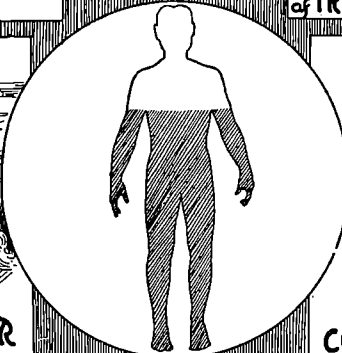
MOST IMPORTANT SOLVENT



COVERS  $\frac{3}{4}$  of the EARTH'S SURFACE - A MEDIUM of TRANSPORTATION



WATER POWER



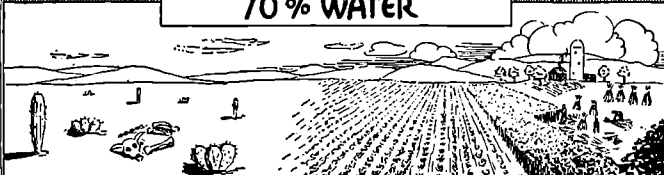
HUMAN BODY IS 70% WATER



DRINKING

WASHING

COOKING

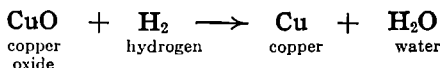
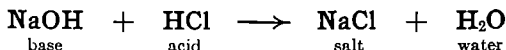


NO WATER

PLENTY of WATER

ESSENTIAL TO SOIL FERTILITY

FIG. 13.

3. *Reduction* of an oxide by hydrogen:4. *Neutralization* (page 70):**Physical Properties**

1. *General Characteristics*.—Water is a liquid at ordinary temperatures, and is tasteless and odorless. It is colorless in thin layers, but has a bluish tinge when observed through great thicknesses.

2. *Freezing and Boiling Points*.—Under a pressure of one atmosphere, water freezes at 0°C. (32°F.) and boils at 100°C. (212°F.). These temperatures are employed as the fixed points on thermometer scales.

3. *Standard for Specific Gravities*.—Water has the peculiar property of expanding when it cools from 4°C. to 0°C., and while freezing. At 4°C. it has its greatest density, namely 1 (*i.e.*, 1 cc. weighs 1 gm.). Water is used as a standard for determining the relative weights (specific gravities) of all other liquids and solids.

4. *Solvent Action*.—Water dissolves almost all substances to some extent, and may therefore be considered the universal solvent. The pleasing taste of drinking water is due to dissolved air and minerals.

5. *Conductivity*.—Water, when pure, does not conduct the electric current. The addition of even a small amount of an acid, base, or salt (*electrolytes*) causes water to become a conductor of electricity (pages 12 and 65).

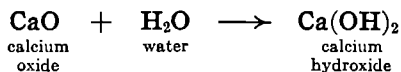
**Chemical Properties**

1. *Stability*.—Water is extremely stable. It must be heated to a temperature of 1000°C. before any decomposition takes place. At 2500°C. only about 3% of it has been converted into hydrogen and oxygen.

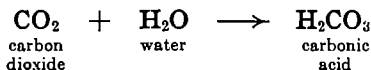
2. *Action with Metals*.—Water reacts with various active metals such as sodium, calcium, and iron, which replace either part or all of its hydrogen (page 22).

3. *Electrolysis*.—Water is decomposed by the electric current (page 12).

4. *Action with Metallic Oxides.*—Water combines with certain metallic oxides (*basic anhydrides*), forming bases:

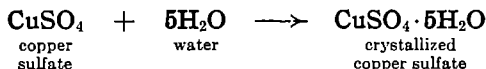


5. *Action with Non-Metallic Oxides.*—Water combines with certain non-metallic oxides (*acid anhydrides*), forming acids:



6. *Hydrolysis.*—Water reacts with certain salts, forming both an acid and a base (page 71).

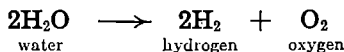
7. *Water of Crystallization.*—Water combines with some compounds when they crystallize from solution, forming *hydrates* (page 40):



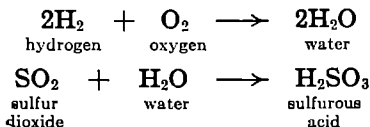
8. *Action as Catalyst.*—Many chemical reactions cannot take place unless there is some moisture present. For example, when a spark is applied to a mixture of perfectly dry hydrogen and oxygen, no action occurs. However, if there is a trace of water present, the mixture explodes readily. Similarly, phosphorus does not ignite in dry air, but burns readily if there is even a minute amount of moisture present.

**Analysis and Synthesis.**—One of the first steps in the study of a chemical compound is to learn its composition, that is, what elements it contains, and the exact amount of each. There are two methods of obtaining this information:

1. *Analysis* is the process of decomposing a compound into its constituent elements:



2. *Synthesis* is the process of forming a compound by the direct union of its elements or of simpler compounds:





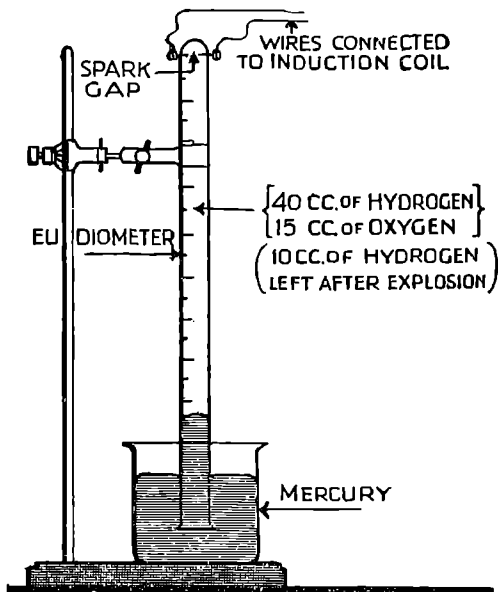


Fig. 14. A eudiometer for determining the composition of water by volume.

### SYNTHESIS OF WATER BY VOLUME (FIG. 14)

METHOD	ILLUSTRATIVE EXPERIMENT
Mix definite quantities of oxygen and hydrogen in a eudiometer.	Volume of hydrogen      40 cc. Volume of oxygen        15 cc.
Pass a spark through the mixture.	An explosion occurs and drops of water form on the inner surface of the tube. Some gas is left in the tube.
Measure and test the gas remaining in the tube.	The excess gas measures 10 cc., and proves to be hydrogen.
Calculate the volumes of hydrogen and oxygen used.	Volume of hydrogen at start   40 cc. Volume of hydrogen left        10 cc. Volume of hydrogen used       30 cc. Volume of oxygen used         15 cc.

CONCLUSION: Water is composed of two parts of hydrogen and one part of oxygen, by *volume*.

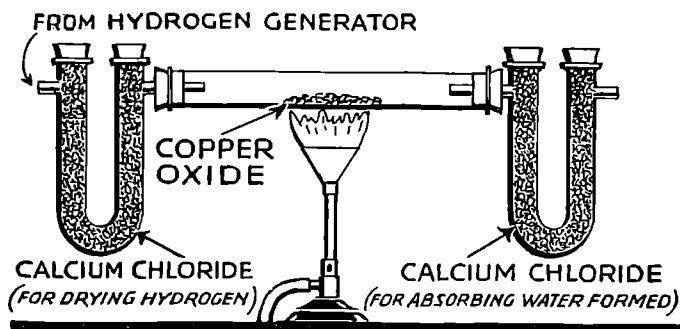


FIG. 15. Apparatus for determining the composition of water by weight.

#### SYNTHESIS OF WATER BY WEIGHT (FIG. 15)

METHOD	ILLUSTRATIVE EXPERIMENT
Pass <i>dry</i> hydrogen over heated copper oxide (CuO).	Weight of CuO tube            100 gms.
Allow the water vapor formed by the reduction of the copper oxide to be absorbed in calcium chloride (CaCl <sub>2</sub> ).	Weight of CaCl <sub>2</sub> tube        241 gms.
Calculate the gain in weight of the calcium chloride tube. This gives the weight of water formed.	Weight of CaCl <sub>2</sub> tube (end) 250 gms. Weight of CaCl <sub>2</sub> tube (start) 241 gms. Weight of water formed        9 gms.
Calculate the loss in weight of the copper oxide tube. This gives the weight of oxygen used.	Weight of CuO tube (start) 100 gms. Weight of CuO tube (end)    92 gms. Weight of oxygen used        8 gms.
Subtract the weight of oxygen used from the weight of water formed. This gives the weight of hydrogen used.	Weight of water formed        9 gms. Weight of oxygen used        8 gms. Weight of hydrogen used       1 gm.

CONCLUSION: Water is composed of one part of hydrogen and eight parts of oxygen, by *weight*.

**Analysis of Water by Volume.**—When an electric current is passed through water containing a little sulfuric acid, the water is decomposed into its component gases, in the ratio of two volumes of hydrogen to one volume of oxygen. This electrolytic process is an important commercial source of these two gases.

NOTE.—The *volumetric* analysis of a substance (always a gas) tells what its component gases are, and the *volume* of each. The *gravimetric* analysis of a substance (either gas, liquid, or solid) tells what its component elements are and the *weight* of each.

**Law of Definite Proportions.**—No matter from what source water may be obtained, the ratio of the weights of hydrogen and oxygen which it contains is always the same. This uniformity of composition is true of every chemical compound. Numerous careful experiments have shown that a given compound always contains the same elements in exactly the same proportion by weight. This fundamental truth is expressed in the Law of Definite Proportions, which is stated as follows: *Every compound has a definite composition by weight.*

**Impurities in Water.**—Water, as it condenses out of the atmosphere in the form of rain, snow, hail, dew, etc., is the purest form of the compound in nature. In contact with the air and earth, however, it soon takes up impurities such as gases, bacteria, dissolved substances, and finely divided material in suspension. Some of this foreign matter may be harmless to human beings, or in some cases even beneficial. Other impurities are harmful, and must be removed in order to make the water fit for human consumption, or for industrial and domestic uses.

**Purification of Water.**—The table below summarizes the methods commonly employed to remove foreign matter from water. Each method of treatment is designed to remove one or more types of impurity.

METHOD	IMPURITIES REMOVED
<i>Filtration (in the home).</i> Water is filtered through charcoal.	Insoluble solids. Gases.
<i>Filtration (large-scale).</i> For city water supplies, water is passed through layers of sand and gravel.	Insoluble solids.

(Continued on p. 36)

METHOD	IMPURITIES REMOVED
<i>Boiling.</i> This is the most convenient method for purifying small amounts of water at home or in camp.	Bacteria. Dissolved gases.
<i>Distillation</i> (Fig. 16). Water is vaporized by heat, and the vapor is condensed by cooling.	Bacteria. Non-volatile impurities (solids or liquids, dissolved or suspended, which have a higher boiling point than water).
<i>Chlorination.</i> About four pounds of chlorine per million gallons of water are added. This method of water purification is now used more extensively than any other.	Bacteria.
<i>Aëration.</i> Water is sprayed into the air. This oxidizes the organic matter present in the water.	Gases. Substances that give an unpleasant odor or taste to the water. Bacteria.
<i>Coagulation.</i> Alum (sometimes with lime) is usually added to purify city water.	Bacteria. Particles in colloidal suspension.
<i>Ozone.</i> This active form of oxygen is passed through water to oxidize some impurities.	Organic matter. Bacteria.
<i>Ultra-Violet Rays.</i> This powerful type of radiation destroys microscopic forms of life.	Bacteria.

## SOLUTION

## Definitions

1. A *solution* is a uniform mixture consisting of a solute and a solvent. The proportions of these ingredients may vary widely. Hence, a solution is a mixture and not a compound, because it does not conform to the Law of Definite Proportions.

2. The *solute* is the substance which goes into solution. It may be either a solid, a liquid or a gas.

3. The *solvent* is the substance which dissolves the solute. It is usually a liquid like water, although it may be either a solid, a liquid, or a gas.

4. A *suspension* is a mixture of a solid and a liquid in which the solid particles settle out on standing. Sand in water is an example.

5. *Miscible liquids* are liquids which remain mixed in all proportions, thus forming a true solution. *Examples:* alcohol and water; glycerine and water.

6. *Non-miscible liquids* are liquids which do not mix, and tend to separate on standing. Such liquids, when thoroughly shaken, form an emulsion. *Example:* oil and water.

7. An *emulsion* is a mixture of two or more non-miscible liquids in which separation takes place very slowly. Milk is an emulsion.

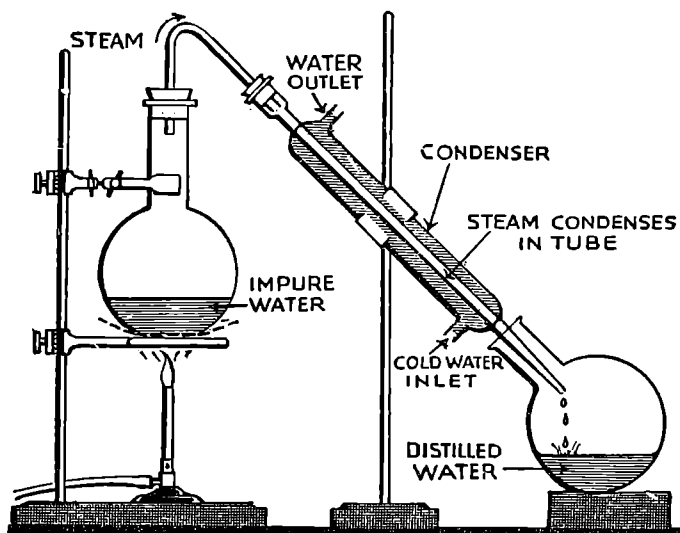


FIG. 16. Apparatus for distilling water in the laboratory.

**Kinds of Solutions.**—Solutions are classified according to the proportion of solute and solvent which they contain.

1. A *dilute solution* is one which contains a small amount of solute in proportion to the solvent.

2. A *concentrated solution* is one which contains a large amount of solute in proportion to the solvent.

3. An *unsaturated solution* is one which can dissolve more of the solute under the same conditions of temperature and pressure. For example, if a pinch of sugar is dissolved in a glass of water, the solution is unsaturated because it can readily dissolve more sugar.

4. A *saturated solution* is one which contains all the solute it can normally dissolve at a given temperature and pressure. For example, if several teaspoonsful of sugar are stirred in a glass of water, some of the sugar will dissolve, while the rest will drop to the bottom of the glass. The clear solution is now said to be saturated at the given temperature and pressure.

5. A *supersaturated solution* is one which contains more of the solute than it can hold under the given conditions of temperature and pressure. For example, sodium thiosulfate (*hypo*) is more soluble in hot water than in cold. Let us assume that at 20°C. (approximately room temperature) 30 grams of *hypo* dissolved in a given amount of water produce a saturated solution. On raising the temperature of this solution to 60°C. we find that the same amount of solvent will now dissolve 40 grams of *hypo*, producing a saturated solution. If we next carefully cool the solution back to 20°C., we observe that the excess *hypo* does not crystallize out, but remains dissolved, thus yielding a solution of *hypo* at 20°C. which has more solute in it than a saturated solution at that temperature should normally have. Such a solution is said to be *supersaturated*. This condition is an unstable one. If a minute crystal of *hypo* is added ("seeding"), or if the solution is shaken, the 10 grams of solvent which are in excess will quickly crystallize out, yielding the original saturated solution which contains 30 grams of *hypo* dissolved in the given amount of water at 20°C.

### Effect of Temperature on Solubility

1. *Solids*.—An increase in the temperature of the solvent usually increases the solubility of solids. For example, sugar is much more soluble in hot water than in cold. *Exceptions:* (a) Sodium chloride is about equally soluble in hot and cold water. (b) Calcium sulfate is less soluble in hot water than in cold.

2. *Gases*.—An increase in the temperature of the solvent decreases the solubility of gases. If a glass of drinking water, for example, is allowed to stand in a warm room, bubbles of air slowly form on the inner surface of the glass. The reason is that the air which is in solution in the cold water is less soluble as the water becomes warm, and some of it is forced out of solution.

### Effect of Pressure on Solubility

1. *Solids*.—An increase in pressure has very little effect on the solubility of solids.

2. *Gases*.—An increase in pressure increases the solubility of gases. Soda water, for example, contains dissolved carbon dioxide which is put in solution and kept there by great pressure. When the liquid is drawn from its container, the pressure is reduced and much of the gas escapes in the form of bubbles.

**Aids to Solution**.—The speed with which a solid goes into solution is increased by *pulverizing*, *stirring*, and *heating*. Pulverizing and stirring cause a solute to dissolve *faster*, but do not increase the *amount* dissolved. Heating, however, not only hastens the process of solution, but usually increases the amount of solid dissolved, as explained above.

**Solutions and Suspensions**.—The following table compares the characteristics of a solution and of a suspension:

IN A SOLUTION THE SOLUTE PARTICLES—	IN A SUSPENSION THE SUSPENDED PARTICLES—
1. Are so small that they cannot be seen, thus rendering the solution clear.	1. Are relatively coarse and can be seen, thus rendering the suspension turbid.
2. Do not settle out on standing.	2. Settle out on standing.
3. Readily pass through any filter.	3. Do not pass through an ordinary filter.
4. Are uniformly distributed through the entire solution.	4. Are not uniformly distributed through the entire suspension.
5. Are taken up by the solvent in a definite amount at a given temperature.	5. Are taken up by the suspension within no definite limits.
6. Raise the boiling point and lower the freezing point of the solvent.	6. Have no effect on the boiling point or freezing point of the liquid in which they are suspended.

**Crystals**.—A crystal is a solid mass having a regular geometrical shape. For example, salt crystals are cubes; alum crystals are octahedral (eight-sided); etc. Crystals may be formed from solution in several ways, as follows:

1. *By the Evaporation of the Solvent.*—When a solution of copper sulfate is heated, the water evaporates, leaving blue crystals of copper sulfate.

2. *By Cooling a Hot, Saturated Solution.*—The solute in this case must be one which is less soluble in cold water than in hot. When a hot, saturated solution of potassium nitrate is cooled, the excess solute crystallizes out of solution.

3. *By "Seeding" a Supersaturated Solution.*—This may be illustrated by the solution of *hypo*, described earlier in this chapter.

NOTE.—Many crystals can be formed without the use of water. Some substances, like sulfur, form crystals on being cooled from the melted state. Other substances, like iodine, form crystals by *sublimation*.

**Water of Crystallization.**—Many compounds, when crystallizing from their water solutions, combine chemically with a definite amount of water which is an essential substance in the structure of the crystal. This water is called *water of crystallization*, or *water of hydration*. A substance which contains water of crystallization is called a *hydrate*. When a crystal of a hydrate is heated, or, in some cases, merely exposed to the air, it gives off its water of crystallization and changes into a powder form. In this form it is called an *anhydrous* compound. Examples:

SALT	ANHYDROUS FORM	HYDRATE FORM
Copper sulfate ( <i>blue vitriol</i> )	$\text{CuSO}_4$	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Calcium sulfate ( <i>gypsum</i> )	$\text{CaSO}_4$	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
Sodium carbonate ( <i>washing soda</i> )	$\text{Na}_2\text{CO}_3$	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

Certain crystals contain water which is not part of the crystal, but which is mechanically enclosed. Common salt is an example. When heat is applied to such crystals, the water is rapidly converted into steam, causing the crystals to crackle and burst (*decapitation*). Many crystals (*e.g.*, diamond, quartz) contain no water.

**Deliquescence and Efflorescence.**—Certain compounds exhibit an unstable composition with regard to their water of hydration. Under various conditions of temperature and moisture, some of them absorb water from the air, while others give it up.



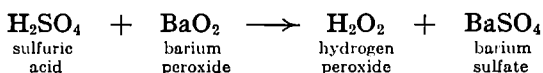
1. A *deliquescent* substance is one that absorbs enough water from the air to dissolve in it. For example, calcium chloride and sodium hydroxide become wet on exposure to the air.

2. An *efflorescent* substance is one that gives up its water of crystallization when exposed to the air. For example, crystals of washing soda ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ), on standing in the air, lose their water of crystallization, and crumble to a white powder.

### HYDROGEN PEROXIDE ( $\text{H}_2\text{O}_2$ )

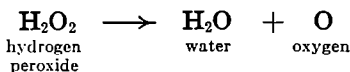
**Composition.**—Hydrogen and oxygen combine to form two different compounds, namely, water, containing 1 part of hydrogen and 8 parts of oxygen by weight; and hydrogen peroxide, containing 1 part of hydrogen and 16 parts of oxygen by weight.

**Preparation.**—Hydrogen peroxide is prepared by the action between dilute sulfuric acid and barium peroxide:



**Properties.**—Pure hydrogen peroxide is a clear, syrupy liquid about one and one-half times as heavy as water. Because it holds its excess of oxygen very loosely, it is a powerful oxidizing agent. It is so unstable that it readily explodes when exposed to heat or light. The ordinary variety that is used in the home is a water solution containing about 3% hydrogen peroxide.

**Uses.**—Hydrogen peroxide decomposes according to the following equation:



The oxygen liberated in this case is called *nascent* oxygen. When a gaseous element is liberated from a compound, the element is said to be in the *nascent state* at the moment when its atoms are set free, and before they combine to form molecules. In this condition, the gas is highly active chemically. Hence, hydrogen peroxide is used extensively as an antiseptic to check the growth of bacteria, and as a bleaching agent for hair, silk, wool, feathers, and ivory. It has the added advantage that, when it decomposes, it leaves nothing but water.

**Law of Multiple Proportions.**—In water, 8 parts by weight of oxygen are united with 1 part by weight of hydrogen. In hydrogen peroxide, 16 parts by weight of oxygen are united with 1 part by weight of hydrogen. Hence the amounts of oxygen which combine with a fixed amount of hydrogen to form these two compounds are in the ratio of 8 to 16, or 1 to 2. This relationship, and a similar relationship existing in the oxides of carbon, are represented in the following tables:

COMPOUNDS	FORMULAS	FIXED WEIGHT OF HYDROGEN	WEIGHTS OF OXYGEN	RATIO OF WEIGHTS OF OXYGEN
Water.....	H <sub>2</sub> O	1	8	1
Hydrogen peroxide.	H <sub>2</sub> O <sub>2</sub>	1	16	2

COMPOUNDS	FORMULAS	FIXED WEIGHT OF CARBON	WEIGHTS OF OXYGEN	RATIO OF WEIGHTS OF OXYGEN
Carbon monoxide...	CO	12	16	1
Carbon dioxide.....	CO <sub>2</sub>	12	32	2

The above compounds illustrate the Law of Multiple Proportions, which states that: *When two elements combine to form a series of compounds, if the weight of one element is fixed, the weights of the other element that combine with this fixed weight are in the ratio of small whole numbers to each other.*

### QUESTIONS

- (a) What is the composition of water by volume? (b) How is this determined by (1) analysis; (2) synthesis?
- A mixture of 50 cc. of hydrogen and 50 cc. of oxygen was put into a eudiometer and exploded by means of a spark. Which gas, and how much of it, was left in the eudiometer?
- From the following data for the determination of the composition of water by weight, calculate the ratio between hydrogen and oxygen in water:

Weight of CuO tube at start . . . . .	234 gms.
Weight of CuO tube at end . . . . .	218 gms.
Weight of CaCl <sub>2</sub> tube at start . . . . .	324 gms.
Weight of CaCl <sub>2</sub> tube at end . . . . .	342 gms.

4. Explain why water pipes sometimes burst in winter.
5. What chemical actions occur in nature which give off water vapor to the air?
6. Of the following substances, state which do and which do not follow the Law of Definite Proportions, giving a reason for your answer in each case: (a) salt, (b) soil, (c) air, (d) sugar solution, (e) washing soda.
7. (a) What is the purest form of water found in nature? (b) Why is it practically impossible for water to remain pure in nature?
8. (a) What is distilled water? (b) Why is distilled water used in making solutions for careful work in the laboratory?
9. A small quantity of sugar is stirred in water until the sugar disappears. What name is given to (a) the sugar? (b) the water? (c) the resulting liquid?
10. Common table salt becomes moist and does not pour in damp weather because of the presence of a small amount of magnesium chloride. What property of magnesium chloride is responsible for such behavior?
11. Explain why "soda water" does not bubble while in a sealed bottle, but does when poured into a glass.
12. Describe how you would determine whether a given solution is (a) unsaturated, (b) saturated, (c) supersaturated.
13. If the crystallized and the powder forms of sodium carbonate are the same in price for a given weight, which would you buy for greater economy? Why?
14. Would you use an efflorescent or a deliquescent substance to dry hydrogen? Explain.
15. When hydrogen peroxide is put on a fresh cut, much frothing takes place. (a) What gas is liberated? (b) What property of hydrogen peroxide does this illustrate?

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. The composition of water by weight is . . . . . part(s) of hydrogen and . . . . . part(s) of oxygen; its composition by volume is . . . . . part(s) of hydrogen and . . . . . part(s) of oxygen.

2. A dissolved solid . . . . . the boiling point and . . . . . the freezing point of water.
3. A deliquescent substance . . . . . water when exposed to the air; an efflorescent substance . . . . . water when exposed to the air.
4. When the temperature of the solvent is raised, the solubility of a solid is usually . . . . .
5. Water is filtered to remove . . . . ., is boiled to remove . . . . ., and is chlorinated to remove . . . . .
6. When kerosene oil and water are shaken vigorously, an . . . . . is formed.
7. A solute may usually be separated from the solvent by . . . . .
8. A gas newly liberated from its compound is said to be in the . . . . . state.
9. Calcium chloride is used on tennis courts to settle dust because of its property of . . . . .
10. Ammonia gas is . . . . . (more, less) soluble in cold water than in hot.
11. Drinking water can be made from sea water by the process of . . . . .
12. Decomposing a compound into its elements is called . . . . .; building up a compound from its elements is called . . . . .
13. Water combines with metallic oxides to form compounds called . . . . ., and with non-metallic oxides to form compounds called . . . . .
14. When sodium reacts with water, . . . . . is liberated.
15. Water has its greatest density at the temperature of . . . . .

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A  | B                                |
|--|----------------------------------|
| ( ) Can dissolve more solute             | 1. Gasoline and kerosene         |
| ( ) Suspension                           | 2. Analysis                      |
| ( ) Aids to solution                     | 3. Filtration                    |
| ( ) About 70% water                      | 4. To oxidize organic impurities |
| ( ) Miscible liquids                     | 5. Calcium chloride              |
| ( ) To dry hydrogen                      | 6. Synthesis                     |
| ( ) Rids water of undissolved impurities | 7. Unsaturated solution          |
| ( ) Aération                             | 8. Pulverizing and stirring      |
| ( ) Electrolysis of water                | 9. Hydrate                       |
| ( ) Crystallized copper sulfate          | 10. Hydrogen peroxide            |
|  | 11. Human body                   |
|  | 12. Muddy water                  |

## CHAPTER 5

### ATOMS, MOLECULES, FORMULAS, AND EQUATIONS

**Importance.**—In chemistry we deal with the interaction of a large variety of substances—water, air, salt, acids, metals, etc. Before we can understand the chemical conduct of these very finite materials, however, we must picture to ourselves the behavior of extremely tiny particles—the atoms and molecules which are believed to make up all the things about us. This involves imagination and theory, for no one has ever seen these ultimate constituents of matter. Yet, imagination and theory, followed by experimental proof, are the fundamental steps in the modern scientific method.

Utilizing these theoretical concepts of atoms and molecules, the chemist has devised a highly practical, graphic system of representing elements and compounds, how they are built up, and how they react with each other. In his symbols, valence, formulas, and equations, he has at his command an efficient set of tools for indicating what is present in his test tube, and what happens when he adds something else to it. Chemistry would still be the hodge-podge that it was in the Middle Ages, were it not for the marvels of chemical theory and the ingenuity of chemical shorthand.

**The Atomic Theory.**—In 1803, John Dalton, an English school teacher, devised his famous *Atomic Theory*, which makes the following assumptions:

1. All matter is composed of very small particles, called *atoms*.
2. All the atoms of the same element are exactly alike in size, shape, and weight, but differ in all respects from the atoms of every other element.
3. Atoms of most elements are able to unite with atoms of other elements.
4. In all chemical changes atoms do not break up, but act as individual units.

**Law of Conservation of Matter.**—This law has already been stated on page 5. We may now state it in a somewhat different form, as follows: *No weight is either gained or lost in a chemical action.* According to the Atomic Theory, a chemical action involves merely a rearrangement of atoms to form new combinations. During this shifting of atoms, their weights remain unchanged. Consequently, the total weight of the products should be exactly equal to the total weight of the original substances. This is experimentally true.

**Law of Definite Proportions.**—This law states, as we have already seen, that *every compound has a definite composition by weight.* According to the Atomic Theory, a chemical compound is formed by the union of a definite number of atoms of one element with a definite number of atoms of another. Since all the atoms of an element have the same weight, and since atoms do not divide in a chemical change, and since, furthermore, no atoms are gained or lost in such a change, it follows that there must exist a definite weight relation between the elements that make up a compound. This is merely another way of stating the Law of Definite Proportions.

**Law of Multiple Proportions.**—This law states, as we have already seen, that *when two elements combine to form a series of compounds, if the weight of one element is fixed, the weights of the other element that combine with this fixed weight are in the ratio of small whole numbers to each other.* This law was discovered by Dalton while investigating the two compounds carbon monoxide and carbon dioxide. It may be conveniently explained on the basis of the Atomic Theory by reference to the two compounds of hydrogen and oxygen already studied, namely, water and hydrogen peroxide.

The composition of water by weight is 1 part of hydrogen and 8 parts of oxygen, and that of hydrogen peroxide is 1 part of hydrogen and 16 parts of oxygen. The ratio between the amounts of oxygen in the two compounds is 1 : 2. According to the Atomic Theory, only *whole* atoms take part in chemical reactions, never parts of atoms; and these whole atoms of an element, such as oxygen, all have the same weight. Hence, if we assume that there are twice as many atoms of oxygen in hydrogen peroxide as in water, the number of hydrogen atoms being the same in both compounds, there must be a ratio of 1 : 2 between the *different* weights of oxygen which combine with a *fixed* weight of hydrogen.

**The Atom.**—The atom is regarded as the unit in chemical action. It is defined as the *smallest particle of an element that takes part in a chemical change*. Although our modern conception of the structure of the atom (page 55) makes it appear far different from the round, hard, impenetrable particle of matter, as conceived by Dalton and others, this simple definition still holds true.

**Atomic Weight.**—Since atoms are too small to be seen even by the most powerful microscope, no ordinary means can be used for weighing them. However, since the *absolute* weights of atoms are of little practical importance, we resort to a system of *relative* weights. The standard of comparison is oxygen, since this element forms compounds with virtually all other elements. Setting the atomic weight of oxygen at 16, the atomic weight of hydrogen, the lightest of the elements, is just a trifle greater than 1 (1.008), while the relative weights of the remaining elements range up to 238 for uranium, the heaviest of all. (See *Table of Atomic Weights*, page 327).

**Symbols.**—A chemical symbol consists of one or two letters of the English or Latin name of an element. A symbol represents one atom of the element. (In problems, it represents the atomic weight of the element.) The following table gives the symbols and the approximate atomic weights of some common elements:

ELEMENT	SYMBOL	APPROXIMATE ATOMIC WEIGHT
Hydrogen.....	H	1
Oxygen.....	O	16
Carbon.....	C	12
Sulfur.....	S	32
Calcium.....	Ca	40
Zinc.....	Zn	65
Chlorine.....	Cl	35.5
Sodium (Latin, <i>natrium</i> ).....	Na	23
Iron (Latin, <i>ferrum</i> ).....	Fe	56
Silver (Latin, <i>argentum</i> ).....	Ag	108

**Molecules.**—A molecule is the smallest particle of an element or a compound that retains the characteristics of the original substance. The molecule of an element may contain one or more atoms of the element. We shall learn later that the molecules of the common

gases oxygen, hydrogen, nitrogen, chlorine, etc. consist of two atoms. The molecule of a compound consists of a group of atoms. For example, a molecule of water contains two hydrogen atoms joined to an oxygen atom.

**Molecular Weight.**—*The molecular weight of a substance is the sum of the weights of all the atoms in its molecule.* Thus, since a molecule of water consists of two atoms of hydrogen and one atom of oxygen, the molecular weight of water is  $(2 \times 1) + 16 = 18$ .

The *gram-molecular weight* or *mole* of a substance is its molecular weight expressed in grams. Thus, a mole of water is 18 grams, while a mole of sodium chloride is 58.5 grams.

**Formulas.**—A chemical formula represents the composition of a molecule. It is made up of one or more symbols that denote the atoms which compose the molecule, together with subscript numbers that indicate how many of each atom are present. (The subscript 1 is never written.) For example, the formula for the hydrogen molecule is  $H_2$ , which shows that a molecule of hydrogen contains two atoms of hydrogen. The formula for water is  $H_2O$ , which shows that a molecule of water consists of two atoms of hydrogen and one atom of oxygen. The following table illustrates the use of these formulas:

SUBSTANCE	FORMULA	MOLECULAR WEIGHT
Hydrogen.....	$H_2$	2
Water.....	$H_2O$	18
Carbon dioxide.....	$CO_2$	44
Hydrochloric acid.....	$HCl$	36.5
Sodium chloride.....	$NaCl$	58.5
Sulfuric acid.....	$H_2SO_4$	98
Sodium hydroxide.....	$NaOH$	40

**Valence.**—Valence is a measure of the ability of an element to combine with other elements. The standard used is hydrogen, since experiments have shown that the hydrogen atom never combines with more than one atom of any other element. Hence, *the valence of an element is the number of atoms of hydrogen with which one atom of the element can combine, or which it can displace.* The valences of most elements are small numbers, ranging between one and six. If we study the formulas of four common compounds con-



taining hydrogen, we can observe clearly the significance of valence. The compounds are:

Hydrogen chloride.....	HCl
Water.....	H <sub>2</sub> O
Ammonia.....	NH <sub>3</sub> or H <sub>3</sub> N
Methane.....	CH <sub>4</sub> or H <sub>4</sub> C

From these formulas it can be seen that chlorine is *univalent* or *monovalent* (valence = 1); oxygen is *divalent* (valence = 2); nitrogen is *trivalent* (valence = 3); and carbon is *tetravalent* (valence = 4).

**Valence of Radicals.**—A radical is a group of elements which behaves as a unit in chemical actions, and has an individual valence, just like any element. For example, the ammonium radical (NH<sub>4</sub>) is *univalent*. Similarly, the sulfate radical (SO<sub>4</sub>) combines with two hydrogen atoms, forming the compound sulfuric acid (H<sub>2</sub>SO<sub>4</sub>); hence it is *divalent*.

**Positive and Negative Valences.**—It is convenient to consider valence as a sort of electrical or magnetic phenomenon. Thus metals almost always have a positive (+) valence, while non-metals usually have a negative (–) valence. Certain elements exhibit either type of valence under different conditions. In general, the formation of a compound is the result of the combining of two types of elements or radicals having opposite valence signs.

The table on page 50 gives the valences of important elements and radicals. It will be observed that copper may have a valence of 1 or 2, while iron may have a valence of 2 or 3. In such cases, the *lower* valence is represented by the ending *-ous* (cuprous, ferrous), and the higher valence by the ending *-ic* (cupric, ferric).

**Writing Formulas.**—A knowledge of the valences listed in the table (page 50) is an essential requirement for writing the great majority of formulas encountered in elementary chemistry. If we know that two elements or radicals combine to form a compound, we proceed to build up its formula by the following steps:

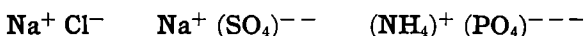
1. Represent by symbols, first the metallic part, and then the non-metallic part, of the compound:



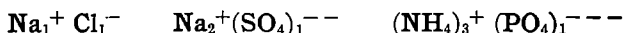
TABLE OF IMPORTANT VALENCES

VALENCE	POSITIVE		NEGATIVE	
I	hydrogen	H <sup>+</sup>	fluorine	F <sup>-</sup>
	sodium	Na <sup>+</sup>	chlorine	Cl <sup>-</sup>
	potassium	K <sup>+</sup>	bromine	Br <sup>-</sup>
	mercurous	Hg <sup>+</sup>	iodine	I <sup>-</sup>
	cuprous	Cu <sup>+</sup>	hydroxide	OH <sup>-</sup>
	silver	Ag <sup>+</sup>	nitrate	NO <sub>3</sub> <sup>-</sup>
	ammonium	NH <sub>4</sub> <sup>+</sup>	chlorate	ClO <sub>3</sub> <sup>-</sup>
II	calcium	Ca <sup>++</sup>	oxygen	O <sup>--</sup>
	magnesium	Mg <sup>++</sup>	sulfur	S <sup>--</sup>
	zinc	Zn <sup>++</sup>	sulfite	SO <sub>3</sub> <sup>--</sup>
	lead	Pb <sup>++</sup>	sulfate	SO <sub>4</sub> <sup>--</sup>
	cupric	Cu <sup>++</sup>	carbonate	CO <sub>3</sub> <sup>--</sup>
	mercuric	Hg <sup>++</sup>		
	ferrous	Fe <sup>++</sup>		
III	aluminum	Al <sup>+++</sup>	nitrogen	N <sup>---</sup>
	ferric	Fe <sup>+++</sup>	phosphorus	P <sup>---</sup>
	arsenic	As <sup>+++</sup>	phosphate	PO <sub>4</sub> <sup>---</sup>
	antimony	Sb <sup>+++</sup>		
	gold	Au <sup>+++</sup>		
IV	carbon	C <sup>++++</sup>	carbon	C <sup>----</sup>
	silicon	Si <sup>++++</sup>	silicon	Si <sup>----</sup>
	sulfur	S <sup>++++</sup>		
	tin	Sn <sup>++++</sup>		

2. Indicate the respective valences above these symbols by means of an appropriate number of plus and minus signs:



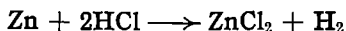
3. Write a subscript number for each element or radical equal to the valence of the other element or radical in the compound. (Radicals should be enclosed in parentheses.)



4. The subscript 1 is omitted, as are also the parentheses of radicals whose subscript is 1, and the plus and minus signs. The completed formulas are therefore:



**Equations.**—A chemical equation consists of symbols and formulas representing a chemical reaction. For example, the following equation represents the reaction between zinc and hydrochloric acid:



A chemical equation tells in shorthand language what substances react and what products are formed. The two sides of an equation are separated by an arrow, which is read “yields” or “forms.”

**Writing Equations.**—If we are given two or more substances which we know will react with each other under certain conditions, and if we know what products result from the reaction, we may write an equation to represent the chemical action in the following way:

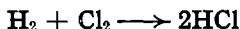
1. On the *left* of the arrow express by means of a symbol or a formula each substance taking part in the reaction.

2. On the *right* of the arrow express in a similar way each product formed.

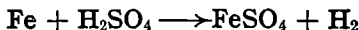
3. *Balance* the equation by writing coefficients before the symbols or formulas, if necessary, so as to give the same number of atoms of each element on both sides of the arrow.

**Aids in Balancing Equations.**—Following are several simple rules which should prove helpful in balancing equations:

1. Most gaseous elements, such as hydrogen, oxygen, nitrogen, and chlorine, exist in the free or uncombined state as diatomic molecules ( $\text{H}_2$ ,  $\text{O}_2$ ,  $\text{N}_2$ ,  $\text{Cl}_2$ ) and should be so written in equations.



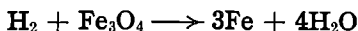
2. In general, a radical is not changed during a chemical action; hence, it may be carried over from the left to the right side of the equation without any alteration.



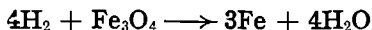
3. It is advisable to start balancing with the most complex formula in the equation, and work back to the simpler formulas. Thus, in writing the equation for the reaction between hydrogen and heated magnetic iron oxide, we first represent the reaction as follows:



It may be seen at a glance that this equation does not balance. Starting with the most complex formula, which is  $\text{Fe}_3\text{O}_4$ , we observe that there must be  $3\text{Fe}$  on the right side of the equation. Also, the four atoms of oxygen in  $\text{Fe}_3\text{O}_4$  must all be used in forming water. We have then:



Further inspection shows that the hydrogen does not balance. It is evident that we must have  $4\text{H}_2$  on the left. The correctly balanced equation is therefore:

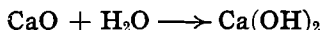
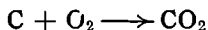


4. Never change the subscripts in any of the formulas, because that would mean an alteration in the composition of the substance, and would thus make the equation state something which is not chemically true.

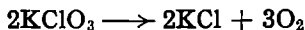
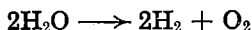
5. Check carefully to make sure that every element is represented by the same number of atoms on both sides of the equation. To facilitate the operation, put a check ( $\checkmark$ ) above each symbol as you reach a balance for the particular element involved.

**Types of Chemical Reactions.**—In general, all chemical reactions can be grouped into four main divisions.

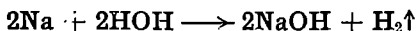
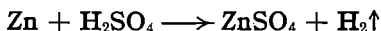
1. *Direct Combination.*—The chemical union of two or more elements or compounds to form a more complex substance is called direct combination or *synthesis*.



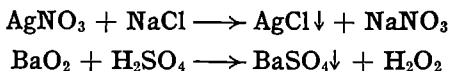
2. *Decomposition.*—The reaction in which a compound is broken up into its elements or into simpler compounds is called decomposition or *analysis*.



3. *Simple Replacement.*—The reaction in which a free element replaces another element in a compound is called simple replacement or *substitution*.



4. *Double Replacement*.—The reaction in which two compounds exchange metallic and non-metallic portions to form two new compounds is called double replacement or *metathesis*.



NOTE: An arrow pointing upward ( $\uparrow$ ) indicates the evolution of a gas; an arrow pointing downward ( $\downarrow$ ) indicates the formation of a precipitate.

### QUESTIONS

1. State the assumptions made in Dalton's Atomic Theory.
2. How does the Atomic Theory explain (a) the Law of Conservation of Matter, (b) the Law of Definite Proportions, (c) the Law of Multiple Proportions?
3. Explain why it is more useful to know the relative weight of an atom than its absolute weight.
4. (a) From what are the symbols of chemical elements derived? (b) Write the symbols of ten common elements.
5. The formula for sulfuric acid is  $\text{H}_2\text{SO}_4$ . (a) What elements are present in this compound? (b) How many atoms of each element are present in one molecule of this compound? (c) How many atoms of oxygen are present in five molecules of  $\text{H}_2\text{SO}_4$ ?
6. (a) Define *valence*. (b) If the valence of element X is 3, and that of element Y is 2, how would you determine the formula of the compound XY?
7. (a) Define *radical*. (b) Write the names and formulas of four radicals, and indicate the valence of each.
8. Write the names and formulas of the compounds formed when each metallic element or radical in the following list combines with each non-metallic element or radical:
  - (a) *Metallic*: H, Na, Ag,  $\text{NH}_4$ , Ca, Fe(-ous), Al
  - (b) *Non-metallic*: Cl,  $\text{ClO}_3$ , S,  $\text{NO}_3$ , OH,  $\text{SO}_4$ ,  $\text{PO}_4$
9. In the following formulas explain what each symbol and number means:  $\text{HgNO}_3$ ,  $\text{Al}_2\text{O}_3$ ,  $(\text{NH}_4)_2\text{SO}_4$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{CaCO}_3$ .
10. Potassium chlorate is decomposed by heat. State the steps that you would follow in balancing the equation for this reaction.
11. Write the equation for the reduction of copper oxide by hydrogen, and show that this equation also illustrates simple replacement.

12. Write a balanced equation for each of the following reactions, and indicate to what general type it belongs: (a) Oxidation of copper. (b) Reduction of magnetic oxide of iron with hydrogen. (c) Preparation of hydrogen from zinc and sulfuric acid. (d) Action between calcium hydroxide and hydrochloric acid. (e) Preparation of hydrogen peroxide. (f) Action of hydrochloric acid on ferrous sulfide. (g) Decomposition of water by electrolysis. (h) Action of potassium on water. (i) Decomposition of mercuric oxide by heat. (j) Burning hydrogen in air.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. An atom is the smallest part of an . . . . . which can take part in . . . . .

2. The Law of Multiple Proportions is illustrated by the two compounds . . . . .

3. The composition of a molecule of a substance is represented by a chemical . . . . .

4. (a) Oxygen is used as a standard for atomic weights because . . . . . (b) Hydrogen is used as a standard for valence because . . . . .

5. In general, chemical compounds are formed by the union of elements or radicals having a . . . . . valence with elements or radicals having a . . . . . valence.

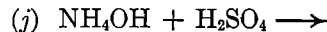
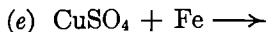
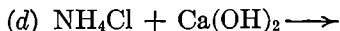
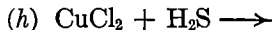
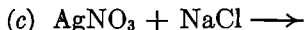
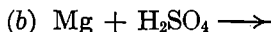
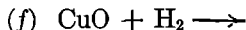
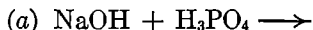
6. An expression consisting of symbols and formulas which represents a chemical reaction is called an . . . . .

7. In an equation, the number written in front of a molecule is called a . . . . .; the number written below and to the right of a symbol is called a . . . . .

8. A gaseous element involved in an equation is usually written as a . . . . . consisting of two . . . . .

9. An equation is balanced to make it conform with the Law of . . . . .

10. *Complete and balance the following equations:*



## CHAPTER 6

### THE ELECTRON THEORY

**Importance.**—Perhaps the most fundamental contribution made in the twentieth century to the sciences of physics and chemistry has been the development of the Electron Theory of the structure of matter. This new concept has cleared up doubts and misunderstandings, explained facts of chemical structure and chemical changes that had always been riddles, correlated data which had hitherto defied organization, and provided a beautifully integrated picture of matter and its behavior. It is true that the theory has its flaws, and considerable work remains to be done in clarifying certain features that are still very much of a puzzle. Nevertheless, the Electron Theory, in its present form, stands as one of the greatest triumphs of modern science, and, in the hands of the chemist, is a most valuable tool in simplifying the study of elements and compounds, and their reactions with each other.

**Atomic Structure.**—Abandoning Dalton's notion that the atom is a hard, round, solid, *indivisible* mass of matter, the modern Electron Theory presents the following revolutionary ideas:

1. The atoms of all the elements are made up of a certain number of positively charged particles called *protons*, and negatively charged particles called *electrons*. The proton is about 1835 times as heavy as the electron.

2. Each particle bears a *unit* charge of electricity.

3. The atom of any element contains exactly the same number of protons and electrons; hence, it is electrically neutral.

4. The dense, central portion of the atom, called the *nucleus*, contains all the protons and about half the electrons; hence, the nucleus is always strongly positive. The modern belief is that no free electrons exist as such within the nucleus, but that each electron is tied up with a proton to form an electrically neutral mass called a *neutron*. The nature of this particle of matter is taken up more fully in Chapter 27.

5. The remaining electrons (*planetary* electrons) are arranged about the central nucleus in one or more *orbits* or *rings*.

6. The number of planetary electrons is equal to the number of *excess* protons in the nucleus.

7. The entire chemical behavior of the atom depends solely on the number of these planetary electrons, and on the way they are arranged in space.

**Atomic Numbers.**—Henry Moseley, a brilliant English chemist, discovered a fundamental relationship among the elements which has proved to be of great value in clarifying our knowledge of atomic structure. He studied the X-ray spectra of various elements and observed a definite variation in wave lengths as he went from the lighter to the heavier elements. As a result of this research, all the elements can be arranged in a series which indicates the number of planetary electrons present in their atoms. In this series, each element differs from its neighbor by one electron outside its nucleus. The *atomic number* of an element is equal to the number of planetary electrons in its atom. Thus hydrogen, with one electron outside the nucleus, has an atomic number of 1. Uranium, with ninety-two electrons outside the nucleus, has an atomic number of 92. (For a complete list of the elements arranged according to their atomic numbers, see page 347). It is interesting to note that, since the weight of the atom is concentrated almost entirely in the nucleus, the *atomic weight* of an element is equal to the weight of its *protons and neutrons*, while the *atomic number* is equal to the number of its *excess protons* or *planetary electrons*.

**Electron Rings.**—The planetary electrons are believed to revolve in definite orbits about the nucleus, much as the planets of the solar system revolve about the sun. The ring nearest the nucleus can hold no more than two electrons. The second and third rings hold a maximum of eight each, and succeeding rings can hold increasingly greater numbers. Figure 17 illustrates the atoms of five common elements. Their electronic structure may be explained as follows:

1. *Hydrogen*, the smallest of all atoms, has a nucleus consisting of one proton around which revolves one electron. Its atomic number is 1.

2. *Helium* has a nucleus consisting of two neutrons and two excess protons. There are two planetary electrons, both in the same ring, relatively close to the nucleus. The atomic number of helium is 2.

3. *Lithium* has three excess protons in its nucleus and three



planetary electrons. Two of these, as in the case of helium, are in the ring close to the nucleus. The third electron is in a new ring some distance away. The atomic number of lithium is 3.

4. *Neon* has ten excess protons in its nucleus. The ten planetary electrons are just sufficient in number to provide two for the first ring and eight for the second, thus filling each ring to capacity. The atomic number of neon is 10.

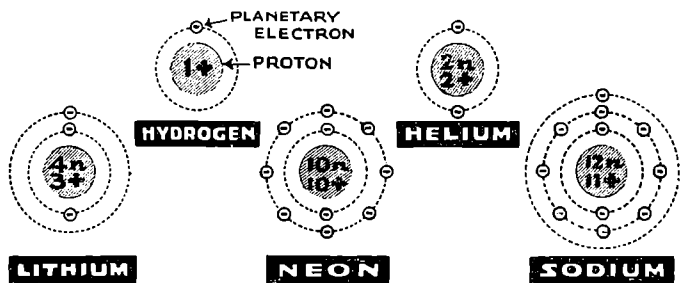


FIG. 17. The electronic structure of five common elements. (The letter *n* in the nucleus represents a *neutron*, composed of a *proton* and an *electron*.)

5. *Sodium* has eleven excess protons in its nucleus. The eleven planetary electrons are enough to fill the first two rings, as in the case of neon, with the eleventh electron starting a new orbit at a still greater distance from the nucleus. The atomic number of sodium is 11.

**Application of the Electron Theory.**—As stated earlier in this chapter, the Electron Theory has shed considerable light on the properties and behavior of elements. While it has not cleared up all the traditional problems of chemistry, the new conception of atomic structure has, in a great measure, explained such matters as (1) valence, (2) metallic, non-metallic and amphoteric elements, (3) chemical activity or affinity, (4) the mechanism of chemical union, (5) oxidation and reduction, (6) isotopes, (7) radioactivity, (8) transmutation of elements, and (9) the periodic arrangement of elements. We shall now see how each of these matters has been illuminated by the Electron Theory.

**Valence.**—In the electronic structure of the atom, the most stable arrangement seems to be one in which all the rings contain their full capacity of electrons, be it two, eight, or even higher for the more distant orbits. In all cases of an incomplete outer ring,

the atom *tends to gain or lose electrons* to form stable structures. *The valence of an element is the number of electrons which its atom must borrow or lend to form a complete outer ring.* For example, the diagram of the fluorine atom (Fig. 18) shows that there are seven electrons in the outer ring. Since eight electrons are required to make this ring stable, fluorine is inclined to borrow one electron. Consequently, the valence of fluorine is 1.

The diagram of the lithium atom (Fig. 18) shows that there is but one electron in the outer ring. By lending that electron to some other atom, the remaining ring now becomes the outer ring, with its capacity fully satisfied. Hence, the valence of lithium is 1.

The diagram of the neon atom (Fig. 17) shows that the outer ring already contains the maximum quota of eight electrons. Since this atom will neither borrow nor lend electrons, neon is said to have a valence of 0.

**Metals and Non-Metals.**—It appears that the atom tends to achieve a complete outer ring by the smallest possible change. Thus, an element having *less than half* the capacity number of electrons in its outer ring will *lend* these electrons, while an element having *more than half* will *borrow* electrons to complete the ring. For example, sodium, which is a typical metal, can lend the sole electron in its outer ring more easily than it can borrow seven to complete the ring. On lending this single electron, the sodium atom becomes charged positively. On the other hand, chlorine, a typical non-metal with seven electrons in its outer ring, can easily borrow one more electron, thus leaving the atom charged negatively. Accordingly, a metal is defined as a "lender of electrons," while a non-metal is defined as a "borrower of electrons."

Some elements contain exactly half, or nearly half, the required number of electrons in the outer ring. An example is carbon (atomic number = 6), which may either lend its four outer electrons or borrow four additional ones. Such an element is called *amphoteric*; it acts either as a metal or as a non-metal.

**Chemical Activity or Affinity.**—Carbon, with four electrons to lend or borrow, is a rather inactive element. Aluminum and nitrogen, with three electrons to lend or borrow, are but slightly more active. Calcium and sulfur, with two electrons to lend or borrow, are fairly active. Sodium and chlorine, with but one electron to lend or borrow, are extremely active. Thus, in general, the fewer the electrons an element must lend or borrow, the greater is its chemical activity.

Helium, neon, and argon are examples of elements whose outer ring is already complete. Such elements show no tendency to lend or borrow electrons, hence they are chemically *inert*.

**Chemical Union.**—When two elements combine to form a compound, the union is believed to result from the mutual lending and borrowing of electrons so that each atom is left with a complete outer ring. For example, when a lithium atom combines with a fluorine atom, the single outer electron of the former fits into the one vacant space in the outer ring of the latter, thus forming the stable compound lithium fluoride, as shown in Fig. 18. Lithium

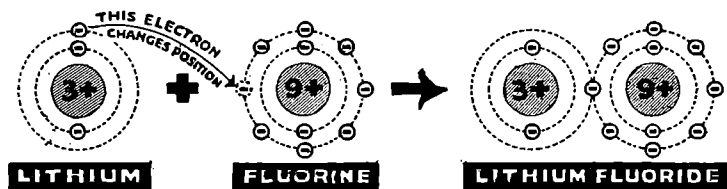


FIG. 18. The formation of lithium fluoride from its elements.

fluoride is known as a *polar compound*, which is a compound formed by the interchange of *single* electrons. In some cases, two atoms unite, not by the borrowing and lending of single electrons, but by the sharing of one or more *pairs* of electrons. Such a linkage results in the formation of a *non-polar compound*. The atoms of the gaseous elements such as oxygen, nitrogen, and fluorine combine into molecules in this manner.

**Oxidation and Reduction.**—According to the Electron Theory, oxidation involves much more than the chemical union of a substance with oxygen, while reduction is more than the mere removal of oxygen from a compound. The following equations illustrate oxidation reactions according to the new viewpoint:

	STANDARD EQUATIONS	ELECTRONIC EQUATIONS
1.	$2\text{Cu} + \text{O}_2 \longrightarrow 2\text{CuO}$	$2\text{Cu}^0 - 4e \longrightarrow 2\text{Cu}^{++}$ $2\text{O}^0 + 4e \longrightarrow 2\text{O}^{--}$
2.	$2\text{Na} + \text{Cl}_2 \longrightarrow 2\text{NaCl}$	$2\text{Na}^0 - 2e \longrightarrow 2\text{Na}^+$ $2\text{Cl}^0 + 2e \longrightarrow 2\text{Cl}^-$
3.	$2\text{FeCl}_2 + \text{Cl}_2 \longrightarrow 2\text{FeCl}_3$	$2\text{Fe}^{++} - 2e \longrightarrow 2\text{Fe}^{+++}$ $2\text{Cl}^0 + 2e \longrightarrow 2\text{Cl}^-$

In both (1) and (2), the neutral atoms of the metals lose electrons in combining with the non-metals. Such loss of electrons is regarded as oxidation. In (3), the Fe atom of  $\text{FeCl}_2$  loses one electron to the extra Cl atom with which it combines in forming  $\text{FeCl}_3$ . In other words, the valence of the Fe has been increased from 2 to 3. A broad concept of oxidation is that it involves: (a) combination with oxygen; or (b) loss of electrons; or (c) raising of positive valence.

Reduction is the reverse of oxidation, both in the restricted sense and in the broad sense. The following equations illustrate reduction reactions:

	STANDARD EQUATIONS	ELECTRONIC EQUATIONS
1.	$\text{CuO} + \text{H}_2 \longrightarrow \text{Cu} + \text{H}_2\text{O}$	$\text{Cu}^{++} + 2e \longrightarrow \text{Cu}^0$ $2\text{H}^0 - 2e \longrightarrow 2\text{H}^+$
2.	$\text{Zn} + \text{S} \longrightarrow \text{ZnS}$	$\text{Zn}^0 - 2e \longrightarrow \text{Zn}^{++}$ $\text{S}^0 + 2e \longrightarrow \text{S}^{--}$
3.	$\text{FeCl}_3 + \text{H} \longrightarrow \text{FeCl}_2 + \text{HCl}$	$\text{Fe}^{+++} + e \longrightarrow \text{Fe}^{++}$ $\text{H}^0 - e \longrightarrow \text{H}^+$

In (1), the combined Cu atom with a valence of +2 takes two electrons from the H atoms, thus changing to free Cu with a valence of 0. Each H atom, having lost an electron, changes its valence from 0 to +1. In (2), the neutral S atom gains two electrons from the Zn atom, and is thus reduced from a valence of 0 to a valence of -2. In (3), an H atom gives an electron to the trivalent Fe atom, thus converting it to divalent Fe. A broad concept of reduction is that it involves: (a) removal of oxygen; or (b) gain of electrons; or (c) lowering of positive valence.

According to the electronic interpretation, direct combination, decomposition, and simple replacement may all be called oxidation and reduction reactions. Double replacement does not involve a transfer of electrons, and therefore cannot be regarded as an oxidation-reduction reaction.

**Isotopes.**—Isotopes are two or more varieties of the same element, having the same atomic number and chemical properties, but having different atomic weights. For example, oxygen is now known to be a mixture of three distinct isotopic varieties of the element. These have atomic weights of 16, 18, and 17 respectively,

the first form being present in greatest abundance. Figure 19 shows how the Electron Theory helps to explain this unusual condition.

It will be observed that each nucleus contains a varying number of neutrons, but always the same number of excess protons. The net positive charge on the nucleus is therefore the same in each case, namely, eight, which means that each atom has eight planetary electrons. Since chemical properties depend on *planetary*

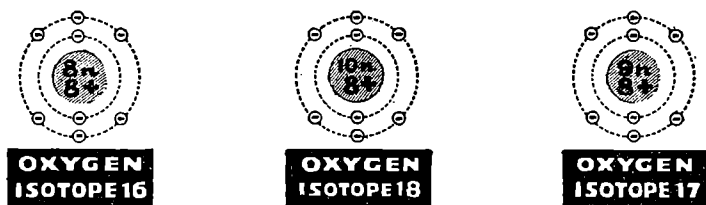


FIG. 19. Three isotopes of oxygen. Ordinary oxygen is a mixture of these three types of atoms.

*electrons*, while atomic weight depends on *nuclear protons* and *neutrons*, we see why an element may have several isotopes. Thus, an element like oxygen is really a mixture of several kinds of atoms, and its atomic weight is the average of the atomic weights of the isotopes present in the mixture. The metal tin is a mixture of eleven isotopes, while bismuth has no less than fourteen.

**Radioactivity.**—In Chapter 23 we shall take up in detail the study of radium, the strange element which undergoes spontaneous decomposition. Several other elements exhibit this property of radioactivity. The breaking up of a radioactive element is not like an ordinary chemical reaction. Whereas most chemical reactions are hastened by raising the temperature, and slowed up by lowering the temperature, the rate of decomposition of radium is not affected by heat, or by any other known human agency.

The Electron Theory leads us to believe that, whereas an ordinary chemical change involves some alteration in the *planetary electrons*, a radioactive change involves the *nucleus* of the atom. The latter is a far more deep-seated action than the former.

**Transmutation of Elements.**—During the Middle Ages, alchemists devoted their efforts to the futile task of transmutation, that is, the conversion of base metals into gold. Today we know

that transmutation actually is possible. For example, radium is undergoing steady decomposition, forming new elements, such as helium and lead. Chemists have already succeeded in splitting atoms by the use of tremendous electrical energy, as a result of which atoms of other elements are produced. Thus, the lithium nucleus has been bombarded with powerful streams of proton particles, resulting in the formation of helium atoms. Chemists now realize that the dream of transmutation is gradually being made a reality by the application of the Electron Theory.

**The Periodic Table.**—In 1869 the Russian chemist, Mendelejeff, devised a system of grouping the elements which has been of great help to the study of chemistry. If all the elements (omitting hydrogen) are arranged in the order of their atomic weights, we note a gradual change in properties from strongly metallic to strongly non-metallic, up to the eighth element. The next eight elements repeat this same gradation of properties. These sixteen elements (with their atomic weights) are arranged in columns as follows:

O	I	II	III	IV	V	VI	VII
He(4)	Li(7)	Gl(9)	B(11)	C(12)	N(14)	O(16)	F(19)
Ne(20)	Na(23)	Mg(24)	Al(27)	Si(28)	P(31)	S(32)	Cl(35.5)

We see that each of the vertical columns is made up of elements that are strikingly similar in properties, as helium and neon, lithium and sodium, and fluorine and chlorine. This fact is expressed in the Periodic Law, as follows: *The properties of the elements are periodic functions of their atomic weights.*

In the light of the newer knowledge of atomic structure, the Periodic Law has been somewhat revised. Moseley's work resulted in an arrangement of the elements in the order of their atomic numbers, which provides us with a classification that is basically sounder and much more reliable. Certain irregularities in the old table, which could not be explained, were eliminated with almost magical facility. Elements which were obviously out of their proper place in the original table automatically fell into their correct position in the new arrangement. Thus, the new Periodic Law, based on atomic numbers, gives us a picture of the elements in perfect gradation.

The Periodic Law has been of immeasurable value in the development of chemistry. It has made the study of the science vastly simpler, in that a knowledge of one element reveals a great deal about others in the same group. It has disclosed errors

in the atomic weights of certain elements, and led to the correction of those errors. Perhaps more remarkable than anything else, has been its value in suggesting the existence of certain new elements and predicting their physical and chemical properties with uncanny accuracy. Every gap in the periodic arrangement has now been filled in, and the table presents a practically complete picture of the entire family of ninety-two elements from hydrogen, the lightest, to uranium, the heaviest.

NOTE.— The most recent discoveries in the field of atomic structure are treated in Chapter 27.

### QUESTIONS

1. (a) Name the types of electrical particles that make up the atom. (b) How do they differ in electric charge and in weight?
2. (a) Draw the electronic picture of a nitrogen atom (atomic number = 7). (b) Explain the chemical properties of nitrogen on the basis of its atomic structure.
3. What valences have the elements whose atomic numbers are 3, 8, 9, 11 and 13 respectively?
4. An atom has an atomic number of 19. Is the element metallic, non-metallic or amphoteric?
5. The atomic numbers of sulfur and chlorine are 16 and 17 respectively. Why is chlorine more active than sulfur?
6. A certain atom has a nucleus made up of 23 protons and 12 electrons. (a) What is the net electrical charge in the nucleus? (b) How many planetary electrons are there in the atom? (c) What is the atomic number? (d) What is the approximate atomic weight? (e) Name the element.
7. Why are metals called *lenders*, and non-metals, *borrowers*?
8. State two ways in which an atom may obtain a complete outer ring of electrons.
9. How does the Electron Theory explain chemical union?
10. Draw an electronic diagram to show the union of sodium and chlorine.
11. Explain why silicon and carbon combine with both metals and non-metals.
12. When steam is passed over heated iron, the steam is reduced to hydrogen. Explain the reduction on the basis of the Electron Theory.
13. By reference to the Electron Theory, prove that the burning of hydrogen to form water is an oxidation-reduction reaction.
14. (a) What are isotopes? (b) How does the Electron Theory explain the existence of isotopes?
15. (a) State the Periodic Law. (b) In what three ways has this law been of great value in advancing our knowledge of chemistry?

## COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. The nucleus of an atom contains ..... and ....., with the ..... always in excess.
2. The valence of an element depends on the number of planetary electrons ..... or ..... during a chemical action.
3. An element which can either lend or borrow electrons is called .....
4. An element whose outer ring is complete is called .....
5. Fluorine is a very active non-metal because .....
6. Potassium is a very active metal because .....
7. The number of planetary electrons in an atom is equal to the ..... of the element.
8. Calcium, whose atomic number is 20, is a ..... (more, less) active metal than sodium because .....
9. According to the newer concept, reduction is either (a) ....., or (b) ....., or (c) .....
10. In the action between  $\text{FeCl}_2$  and  $\text{Cl}$ , the ..... (increase, decrease) of the valence of iron is called .....

## MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

A	B
( ) Planetary electrons	1. An active metal
( ) Zero valence	2. In the nucleus
( ) Atomic number of 11	3. Transmutation
( ) Sodium chloride	4. Determine chemical properties
( ) Neutron	5. Reduction
( ) Helium from lithium	6. An active non-metal
( ) Atomic number of 6	7. Isotopes
( ) Due to protons	8. One proton and one planetary electron
( ) Hydrogen atom	9. A polar compound
( ) Atomic number of 17	10. Atomic weight
	11. Argon
	12. An amphoteric element



## CHAPTER 7

### THE IONIZATION THEORY

**Purpose.**—The Ionization Theory, often called the *Theory of Electrolytic Dissociation*, was devised by Svante Arrhenius, a Swedish chemist, to explain the puzzling characteristics exhibited by many substances in solution. Some of these peculiarities are:

1. Certain substances, when dissolved in water, conduct the electric current, while others do not.

2. Substances whose solutions conduct the electric current are decomposed by it (*electrolysis*), while substances whose solutions do not conduct the electric current are unaffected.

3. Solutions of various acids and bases with the same molecular concentrations (having the same number of dissolved molecules per unit volume of solution) do not possess the same chemical activity.

4. Some compounds, when dissolved in water, lower the freezing point and elevate the boiling point of the water to a greater degree than other compounds of the same molecular concentration.

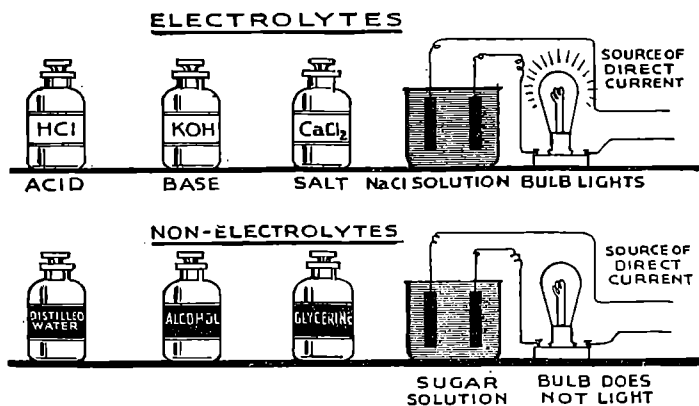


FIG. 20. Electrolytes and non-electrolytes.

**Conductivity of Solutions.**—In order to demonstrate the electrical conductivity of various substances in solution, the apparatus

shown in Fig. 20 is employed. An ordinary electric light bulb is placed in series with two platinum electrodes and a source of current. The electrodes are immersed in the solution to be tested. The bulb lights if the solution permits the current to pass through, and does not light if the solution does not permit the current to flow. A study of the electrical conductivity of thousands of chemical compounds in solution reveals that they may be divided into two general groups, electrolytes, and non-electrolytes.

1. An *electrolyte* is a substance which, in solution, conducts the electric current, and is decomposed by it. All acids, bases, and salts are electrolytes.

2. A *non-electrolyte* is a substance which, in solution, does not conduct the electric current, and is not changed by it. Distilled water, sugar, glycerine, and alcohol are examples of non-electrolytes.

**The Ionization Theory.**—The chief points of the Ionization Theory are as follows:

1. Acids, bases, and salts, when dissolved in water, dissociate into *ions*.

2. Dissociation takes place during the process of solution and is brought about by the action of the solvent on the electrolyte.

3. Ions are atoms or groups of atoms carrying electric charges. These ions are produced when the atoms or groups of atoms gain or lose electrons in the process of dissociation.

4. Metallic ions carry *positive* charges, because of a *loss* of electrons, while non-metallic ions carry *negative* charges, because of a *gain* of electrons.

5. The total number of positive charges in the solution is equal to the total number of negative charges, hence the solution, as a whole, remains electrically neutral.

6. Ions are able to wander freely throughout the solution.

7. When an electric current is passed through the solution, the positive ions (*cations*) move toward the negative pole (*cathode*), while the negative ions (*anions*) move toward the positive pole (*anode*). At the cathode, the positive ions *take up* electrons, while at the anode the negative ions *lose* electrons.

**Illustration of Dissociation.**—According to the Ionization Theory, sodium chloride, when dissolved in water, is broken up into positively charged sodium ions ( $\text{Na}^+$ ) and negatively charged chlorine ions ( $\text{Cl}^-$ ). The ionic equation is:



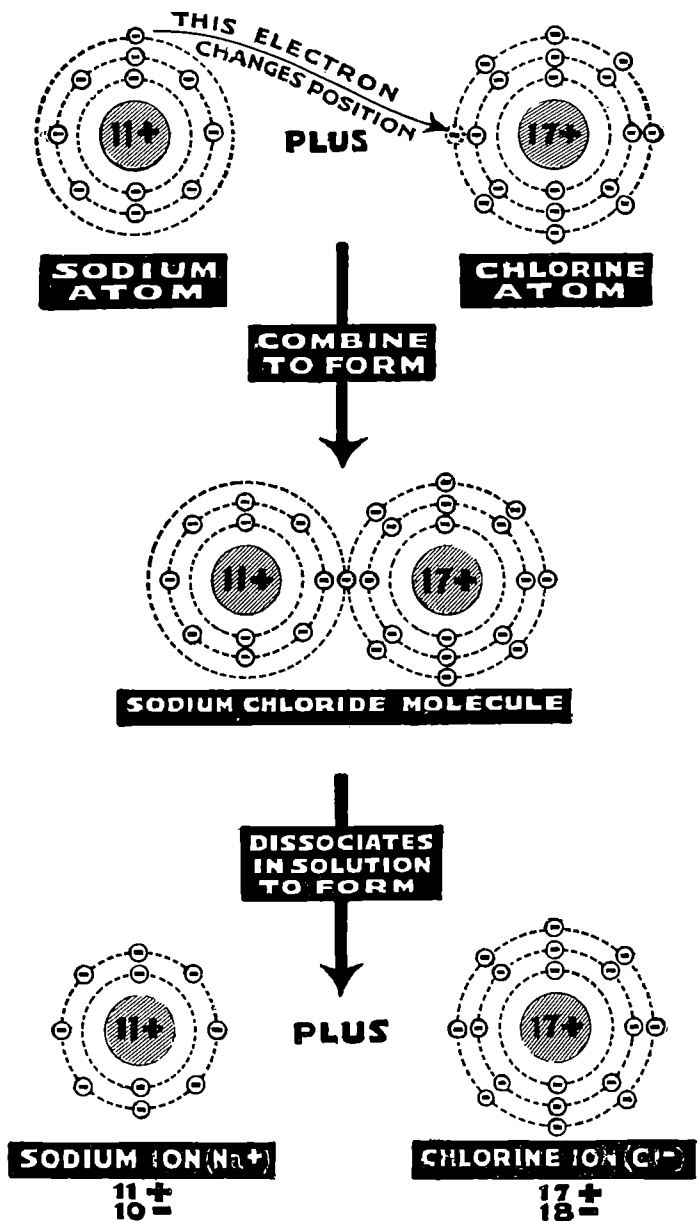
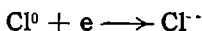
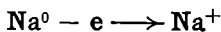


FIG. 21. Electronic picture of the formation of sodium chloride from its elements, and its dissociation in solution.

The Electron Theory gives us a helpful picture of the way in which the elements sodium and chlorine combine to form sodium chloride, and of how this compound dissociates into ions in solution (Fig. 21). The sodium atom, having one electron in its outer ring, unites with the chlorine atom, having one vacant space in its outer ring. By a transfer (*lending*) of the electron in question, both atoms achieve complete outer rings, thus forming the stable and neutral compound sodium chloride. When dissolved in water, the sodium chloride molecule is split into two parts: (1) a sodium ion, which is the original sodium atom shorn of its electron; and (2) a chlorine ion, which is the original chlorine atom supplemented by this transferred electron. The electronic equations involved are:



**Difference between an Atom and an Ion.**—The atom and the ion differ from each other only in that one is neutral, while the other carries an electric charge. So vital is this difference, however, that the atom and the ion are poles apart in their chemical and physical properties. This is illustrated by the following comparison.

SODIUM ATOM ( $\text{Na}^0$ )	SODIUM ION ( $\text{Na}^+$ )
Has incomplete outer ring.	Has complete outer ring.
Has <i>eleven</i> excess positive charges in nucleus as against <i>eleven</i> planetary electrons, thus making the atom electrically neutral.	Has <i>eleven</i> excess positive charges in nucleus, as against <i>ten</i> planetary electrons, thus giving the ion a charge of +1.
Chemically active.	Chemically inactive.
Reacts violently with water, producing hydrogen and sodium hydroxide.	Has no effect on water.
Poisonous.	Non-poisonous.
Has a metallic luster.	Colorless.

**Electrolysis.**—Electrolysis is the process of decomposing a compound in solution by means of the electric current. When the cur-

rent begins to flow through the solution, the action proceeds according to the following steps:

1. The ions move toward the electrode of opposite electrical charge.

2. There the positive ions gain electrons and the negative ions lose electrons, thus becoming electrically neutral atoms or radicals.

3. A less active element such as copper or hydrogen is deposited on the electrode or is evolved as a gas.

4. A more active element such as sodium or chlorine, and a neutral radical such as the sulfate radical, react with water near the electrodes, forming new substances.

**Illustrations of Electrolysis.**—Figs. 22 and 23 explain the electrolytic decomposition of two common substances.

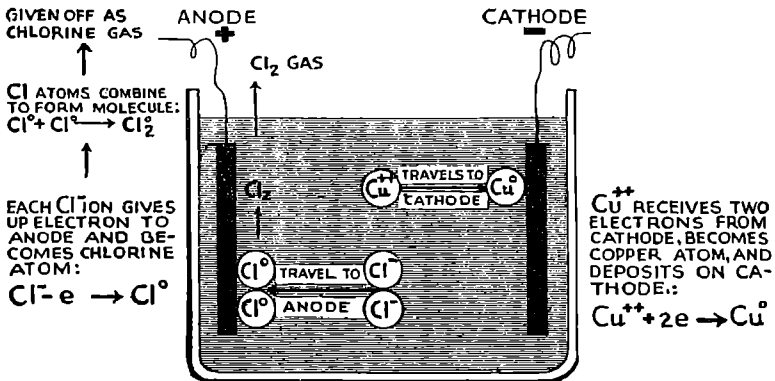


FIG. 22. Electrolysis of copper chloride (CuCl<sub>2</sub>).

**Properties of Acids and Bases.**—We know from experiment that all acids taste sour, turn blue litmus red, and neutralize bases. We know, further, that all acids dissociate in water solution to yield hydrogen ions. Therefore, we conclude that *the properties of acids are really the properties of hydrogen ions*. By a similar process of reasoning we conclude that *the properties of bases are really the properties of hydroxyl ions*. On the basis of the Ionization Theory we may redefine an acid and a base as follows:

An *acid* is a hydrogen compound whose water solution contains hydrogen (H<sup>+</sup>) ions in sufficient number to turn litmus from blue to red.

A *base* is a hydroxide compound whose water solution contains hydroxyl ( $\text{OH}^-$ ) ions in sufficient number to turn litmus from red to blue.

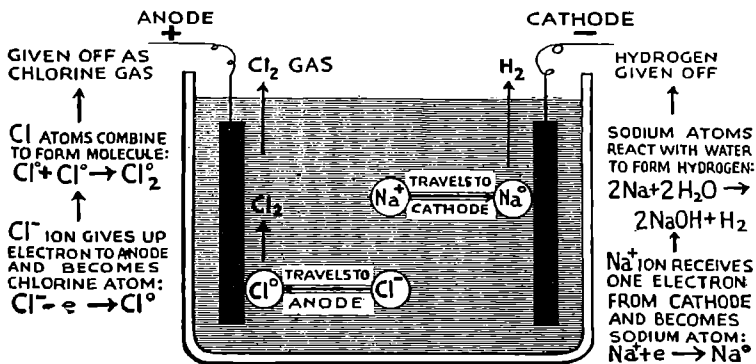
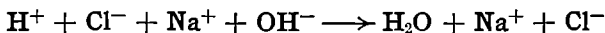


FIG. 23. Electrolysis of sodium chloride ( $\text{NaCl}$ ).

**Chemical Activity of Acids and Bases.**—Since the peculiar properties of all acids result from the hydrogen ions which they produce in solution, it can readily be seen that the strength or chemical activity of an acid is determined by the number of such ions present per unit volume of solution. Thus hydrochloric, nitric, and sulfuric acids, the three acids most commonly used commercially and in the laboratory, are known as the *strong* acids because they are almost completely ionized in dilute water solution. On the other hand, carbonic and acetic acids are known as *weak* acids because they are only slightly ionized in water solution. Applying the same reasoning to bases and the extent to which they produce hydroxyl ions in solution, we see that sodium hydroxide and potassium hydroxide are *strong* bases because they ionize almost completely, while magnesium hydroxide and ammonium hydroxide are *weak* bases because they are only slightly ionized in solution.

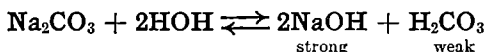
**Neutralization.**—When dilute solutions of sodium hydroxide and hydrochloric acid are mixed, there are four ions present in the mixture, namely,  $\text{Na}^+$ ,  $\text{OH}^-$ ,  $\text{H}^+$ , and  $\text{Cl}^-$ . The  $\text{Na}^+$  ions are constantly combining with the  $\text{Cl}^-$  ions to form molecules of  $\text{NaCl}$ , but these dissociate back into the ions as fast as they are formed. On the other hand, the attraction between the  $\text{H}^+$  ions and the  $\text{OH}^-$

ions results in the neutralization of their charges and the formation of water molecules. Only *two water molecules in a billion* dissociate back into the original ions. Thus, by removal of the ions upon which the properties of acids and bases depend, these properties themselves are removed. The ionic equation for this neutralization reaction is:



According to the Ionization Theory, *neutralization is the union of the hydrogen ions of an acid with the hydroxyl ions of a base, resulting in the formation of undissociated molecules of water.*

**Hydrolysis.**—Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ), formed from a weak acid ( $\text{H}_2\text{CO}_3$ ) and a strong base ( $\text{NaOH}$ ), gives a basic reaction when dissolved in water. This results from a chemical action between the salt and the water according to the equation:



Since carbonic acid is only slightly dissociated, while sodium hydroxide is highly dissociated, there is an excess of  $\text{OH}^-$  ions in

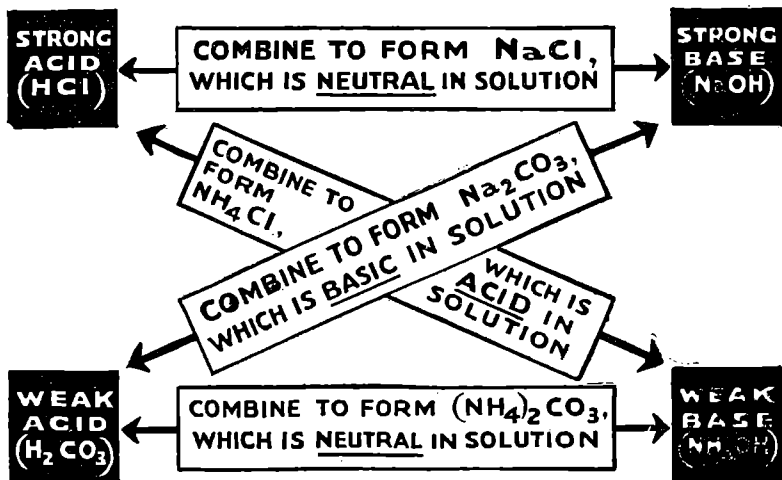
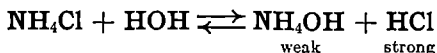


FIG. 24. Hydrolysis. How four types of salts are derived, and how they react in water solution.

the solution, and therefore it gives a basic reaction. Similarly, ammonium chloride ( $\text{NH}_4\text{Cl}$ ), formed from a weak base ( $\text{NH}_4\text{OH}$ ) and a strong acid ( $\text{HCl}$ ) dissolves in water to give an acid reaction, because the  $\text{H}^+$  ions predominate in the solution:



*Hydrolysis is the action between a salt and water resulting in the formation of an acid and a base, one of which is stronger than the other.* Hydrolysis is the opposite of neutralization. Fig. 24 shows schematically the four types of salts derived from strong and weak acids and bases, and indicates how the solutions of these salts react to litmus.

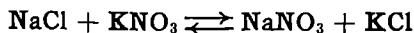
**Abnormal Boiling and Freezing Points.**—A dissolved substance raises the boiling point and lowers the freezing point of a liquid. A given number of molecules of a non-electrolyte, such as sugar or glycerine, dissolved in a given quantity of water, will raise the boiling point of the water a fixed amount, which is the same no matter which non-electrolyte is used. If an equal number of molecules of sodium chloride, an electrolyte, is dissolved in the given quantity of water, the increase in boiling point is *double* that produced by dissolving sugar; for sodium sulfate, the increase is nearly *threefold*.

The Ionization Theory offers a ready explanation for this peculiar behavior. We know that the raising of the boiling point is proportional to the *number* and not the *kind* of dissolved particles. Since sodium chloride molecules dissociate into  $\text{Na}^+$  ions and  $\text{Cl}^-$  ions, a given number of such molecules yield twice as many particles as the same number of sugar molecules. Moreover, since sodium sulfate dissociates into three ions ( $2\text{Na}^+$  and  $\text{SO}_4^{--}$ ), this salt would have about three times the effect of sugar in raising the boiling point of water. A similar process of reasoning explains the abnormal lowering of the freezing point of water by the presence of dissolved electrolytes.

**Reversible Reactions.**—A reversible reaction is one in which the products react so as to form the original substances. If solutions of sodium chloride ( $\text{NaCl}$ ) and potassium nitrate ( $\text{KNO}_3$ ) are mixed, the products formed are sodium nitrate ( $\text{NaNO}_3$ ) and potassium chloride ( $\text{KCl}$ ). These constantly react in such a way



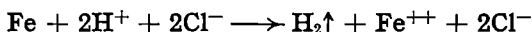
as to yield the original salts. The reaction is expressed by the following equation:



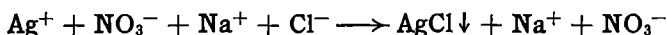
When the speed of the reaction going to the *right* is exactly equal to the speed of the reaction going to the *left*, then we have reached the point of *equilibrium*.

**Reactions Going to Completion.**—A reaction is said to go to completion when it continues in one direction until at least one of the reacting substances is completely used up. In general this is accomplished if one product is removed from the field of action as rapidly as it is formed. There are three ways by which a product is thus removed:

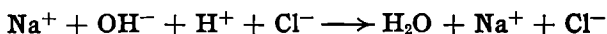
1. By *volatility*, that is, by the formation of a gas. In the preparation of hydrogen by the action between a metal, such as iron, and an acid, such as hydrochloric acid, the reaction goes to completion owing to the liberation of the hydrogen, which thus leaves the field of action. The ionic equation is:



2. By *insolubility*, that is, by the formation of an insoluble substance. In the action between two dissolved substances, insoluble products are often formed which involve the removal of ions from solution. Since an insoluble substance cannot ionize, it is, in effect, removed from the field of action. This is illustrated by mixing solutions of silver nitrate ( $\text{AgNO}_3$ ) and sodium chloride ( $\text{NaCl}$ ). The ions combine as indicated in the following ionic equation, forming insoluble silver chloride ( $\text{AgCl}$ ):



3. By *non-ionization*, that is, by the formation of a product which does not dissociate, or dissociates to only a slight extent. The best example of this type of reaction is neutralization. In such an action the  $\text{H}^+$  ions of the acid combine with the  $\text{OH}^-$  ions of the base, forming molecules of water which do not ionize appreciably. Hence water is removed from the field of action, causing the neutralization to go to completion. The ionic equation is:



## QUESTIONS

1. (a) Distinguish between an electrolyte and a non-electrolyte. (b) Name two electrolytes and two non-electrolytes.

2. According to the Ionization Theory, explain what happens when (a) potassium sulfate is dissolved in water, (b) an electric current is passed through the resulting solution.

3. State three differences between a chlorine atom and a chlorine ion.

4. In terms of the Ionization Theory, define (a) acid, (b) base, (c) neutralization.

5. Explain why sulfuric acid, rather than hydrochloric acid, is used in the electrolysis of water.

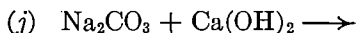
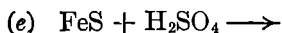
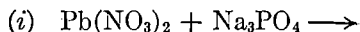
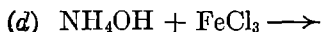
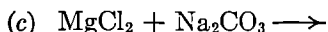
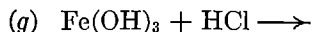
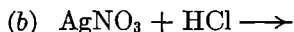
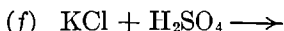
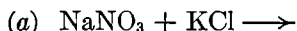
6. Write the ionic equations for the following: (a) potassium hydroxide reacts with sulfuric acid; (b) ammonium hydroxide reacts with nitric acid.

7. (a) Why is silver nitrate in water solution decomposed by the electric current, while a sugar solution is not? (b) Why do not nitric acid and acetic acid conduct electricity equally well?

8. (a) Does a solution of ferric sulfate give an acid or a basic reaction when tested with litmus? (b) Explain this behavior with the aid of an equation.

9. (a) State three conditions under which chemical actions go to completion. (b) In all three cases, what is the one action which really brings about the desired result?

10. Consult the *General Solubility Rules* (page 318) and state which of the following reactions go to completion. Complete the equations of those reactions that go to completion.



## COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. Ions are atoms or radicals that have either ..... or ..... electrons.
2. The process by which an electrolyte in solution is decomposed by the electric current is called .....
3. An electrolyte in solution raises the boiling point and lowers the freezing point of water to a ..... (greater, lesser) degree than .....
4. When an electric current is passed through a solution of sodium hydroxide, the  $\text{Na}^+$  ion is attracted to the ....., while the  $\text{OH}^-$  ion is attracted to the .....
5. In the electrolysis of a sodium chloride solution, metallic sodium is not deposited on the cathode because .....
6. When a copper ion changes to a copper atom, it ..... (gains, loses) ..... (one, two, three) electrons.
7. Sodium hydroxide is stronger than calcium hydroxide because .....
8. Hydrolysis is the action between ..... and ....., forming ..... and ....., one of which is ..... than the other.
9. A solution of aluminum chloride changes the color of litmus from ..... to ..... because the salt is derived from a ..... acid and a ..... base.
10. When a reversible reaction is going at the same rate in both directions, it is said to have reached a point of .....

## MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

- | A                                | B   |
|----------------------------------|---|
| ( ) Hydroxyl ions                | 1. A non-electrolyte                            |
| ( ) Neutralization               | 2. A precipitate is formed                      |
| ( ) By gaining electrons         | 3. A metallic ion                               |
| ( ) Sulfuric acid                | 4. In all bases                                 |
| ( ) Attracts positive ions       | 5. Arrhenius                                    |
| ( ) $\text{AgNO}_3 + \text{HCl}$ | 6. Weakly ionized                               |
| ( ) Glycerine                    | 7. $\text{Cu}^{++} \longrightarrow \text{Cu}^0$ |
| ( ) $\text{NH}_4^+$              | 8. A strong electrolyte                         |
| ( ) Turns litmus red             | 9. $\text{H}^+$ ions + $\text{OH}^-$ ions       |
| ( ) Carbonic acid                | 10. Zinc chloride solution                      |
|                                  | 11. $\text{Zn} + \text{HCl}$                    |
|                                  | 12. Cathode                                     |

## CHAPTER 8

### OXIDES, ACIDS, BASES, AND SALTS

**Introduction.**—If each of the ninety-two elements were to unite with one or more other elements in all possible combinations, the total number of compounds so formed would be indeed staggering. The number of definite compounds actually known to chemists is now over half a million, with the list growing year by year. The task of studying this stupendous array of substances would be an impossible one, were it not for the fact that they may be conveniently grouped into a very few general classes. Thus, to know the properties of one representative member of the class is to have a fairly good acquaintance with thousands of other compounds that belong to the same class.

Practically all the compounds treated in elementary chemistry fall into one of four groups, namely, *oxides*, *acids*, *bases*, and *salts*. Each group is distinguished from the others by peculiarities of composition, preparation, chemical properties, and nomenclature.

#### OXIDES

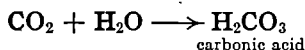
**Definition.**—An *oxide* is a compound consisting of oxygen and another element. Oxides, being composed of only two elements, are known as binary compounds. Many compounds containing oxygen are not oxides, since they consist of three or more elements. An example is calcium carbonate ( $\text{CaCO}_3$ ).

**Important Oxides.**—Water, or *hydrogen oxide* ( $\text{H}_2\text{O}$ ), is the most abundant oxide found on earth, and is man's most useful compound. Carbon dioxide ( $\text{CO}_2$ ) is one of the gases present in the atmosphere. Carbon monoxide ( $\text{CO}$ ) is a constituent of most fuel gases used commercially. Calcium oxide ( $\text{CaO}$ ) is the quicklime used in building construction. Ordinary iron rust is a complex oxide of iron.

**Anhydrides.**—Oxygen may combine with either metals or non-metals. Those oxides which are soluble in water combine

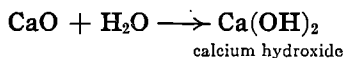
with the water to form either a base or an acid. Such oxides are called *anhydrides*.

1. An *acid anhydride* is a non-metallic oxide that unites with water to form an acid:



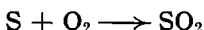
Carbon dioxide is thus the anhydride of carbonic acid. It is sometimes called *carbonic anhydride*.

2. A *basic anhydride* is a metallic oxide that unites with water to form a base:



**Preparation of Oxides.**—Oxides may be prepared in two ways:

1. *Direct Union of an Element with Oxygen.*—When sulfur burns in air, the product formed is sulfur dioxide:



2. *Decomposition of Certain Compounds.*—When limestone (calcium carbonate) is heated, we obtain quicklime (calcium oxide):



**Naming Oxides.**—The name of an oxide consists of the element with which the oxygen is united, followed by the word *oxide*. In some cases, an element forms a series of oxides of varying composition. These are distinguished from each other by the prefixes *mon-*, *di-*, *tri-*, etc., which designate the number of atoms of oxygen present in a molecule of the oxide. *Examples:* carbon *monoxide* (CO), carbon *dioxide* (CO<sub>2</sub>), phosphorus *trioxide* (P<sub>2</sub>O<sub>5</sub>), phosphorus *pentoxide* (P<sub>2</sub>O<sub>5</sub>).

## ACIDS

**Definition.**—*An acid is a hydrogen compound whose hydrogen may be replaced by a metal, and whose water solution changes the color of litmus from blue to red.* (For the ionic definition of an acid, see page 69.) Many compounds, such as sugar and alcohol, contain hydrogen, but are not acids, because they do not conform to the two conditions stated in the definition above. Combined with the hydrogen in an acid, is a non-metal element, or radical, such as Cl, S, NO<sub>3</sub>, SO<sub>4</sub>, and PO<sub>4</sub>.

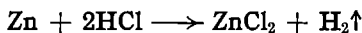
**Important Acids.**—The three most common acids used in the laboratory and in commercial work are *hydrochloric acid* (HCl), *sulfuric acid* (H<sub>2</sub>SO<sub>4</sub>), and *nitric acid* (HNO<sub>3</sub>). These are called the *strong* or *active* acids. Many other acids are known which are *weak* and relatively *inactive*. Thus, "soda water" contains *carbonic acid* (H<sub>2</sub>CO<sub>3</sub>), vinegar contains *acetic acid* (H·C<sub>2</sub>H<sub>3</sub>O<sub>2</sub>), sour milk contains *lactic acid* (H·C<sub>3</sub>H<sub>5</sub>O<sub>3</sub>), and lemon juice contains *citric acid* (H<sub>3</sub>·C<sub>6</sub>H<sub>5</sub>O<sub>7</sub>).

**Properties of Acids.**—Acids exhibit their characteristic properties only when in water solution. The explanation for this was provided in the previous chapter on the Ionization Theory. The four distinguishing properties common to all acids are:

1. *Taste.*—Acids taste *sour*. The familiar sour taste of lemons, grapes, vinegar, rhubarb, etc., is due to the acids which they contain.

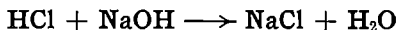
2. *Action on Indicators.*—Certain complex substances, known as *indicators*, undergo a color change when in contact with an acid. The best known indicator is *litmus*, which is extracted from a species of lichen plant. In the presence of an acid, litmus turns from blue to red. *Methyl orange*, another common indicator, turns from yellow to orange in the presence of an acid.

3. *Action with Metals.*—All acids contain hydrogen which is replaceable by a metal:



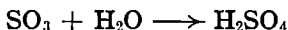
In this illustration, the zinc displaces the hydrogen, which is driven out as a gas, and forms a new compound, zinc chloride, which is a *salt*. Not all metals can replace hydrogen from acids. Copper, silver, and gold are examples of metals that are unable to do so.

4. *Action with Bases.*—Acids react with bases (which are metallic hydroxides), forming salts and water. This reaction is called *neutralization*:

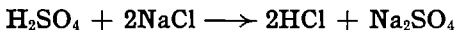


**Preparation of Acids.**—There are two common methods of preparing acids.

1. *Action of an Acid Anhydride with Water*



2. *Action of Sulfuric Acid with a Salt of the Desired Acid*



## BASES

**Definition.**—A base is a compound of a metallic element or radical with one or more hydroxyl (OH) groups. The term is usually applied to a water solution of a metallic hydroxide.

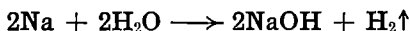
**Important Bases.**—Sodium hydroxide (NaOH) is a typical base used in the laboratory and in commercial work. Ammonium hydroxide (NH<sub>4</sub>OH) is ordinary household ammonia. Calcium hydroxide [Ca(OH)<sub>2</sub>] is dissolved in water to form limewater. Magnesium hydroxide [Mg(OH)<sub>2</sub>], when mixed with water to form a suspension, is the well-known milk of magnesia.

**Properties of Bases.**—Like acids, bases exhibit their characteristic properties only when in water solution. The distinguishing properties common to all bases are:

1. Bases turn red litmus blue.
2. In general, bases have a bitter taste, and their solutions have a soapy or slippery feel.
3. Bases neutralize acids, forming salts and water.

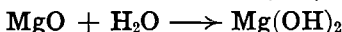
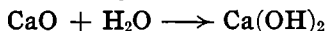
**Preparation of Bases.**—There are four common methods of preparing bases.

1. Action of an Active Metal on Water

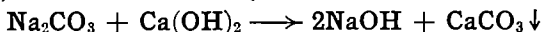


This reaction is the basis for the commercial preparation of sodium hydroxide by the electrolysis of brine.

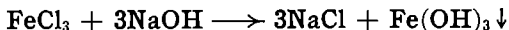
2. Action of a Basic Anhydride with Water



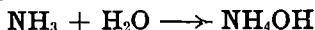
3. Action between a Base and a Salt.—In the soda-lime process of making sodium hydroxide, solutions of sodium carbonate and calcium hydroxide are mixed and boiled, forming insoluble calcium carbonate, which is removed by filtration:



This method may also be used to prepare *insoluble* bases:



4. Action between a Gas and Water.—The only common example of this method is the formation of ammonium hydroxide by dissolving ammonia gas in water:



## SALTS

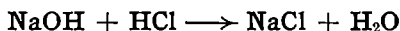
**Definition.**—A salt is a compound consisting of a metal or a metallic radical combined with a non-metal or an acid radical. The term salt, as commonly used, refers to sodium chloride, but in chemistry salt is the general name of a class of compounds which resemble sodium chloride.

**Important Salts.**—The great majority of inorganic compounds are salts. *Sodium chloride* (NaCl), *calcium carbonate* (CaCO<sub>3</sub>), and *ammonium chloride* (NH<sub>4</sub>Cl) are common salts used in the laboratory. In ammonium chloride the radical NH<sub>4</sub> acts like a metal. The earth's crust is composed largely of salts. Thus, limestone is mainly calcium carbonate, and clay is an impure aluminum silicate.

**Properties of Salts.**—Salts vary widely in characteristics. As a general rule, salts are white, crystalline solids with a "salty" taste. They are usually soluble in water, and neutral to litmus. Among the many exceptions to the above, are copper sulfate, which is blue; calcium carbonate (marble), which is insoluble in water; and sodium carbonate, which, in solution, reacts basic to litmus, owing to hydrolysis.

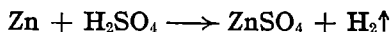
**Preparation of Salts.**—Salts are prepared by a variety of methods.

1. *Neutralization.*—Neutralization is the reaction between an acid and a base, forming a salt and water:

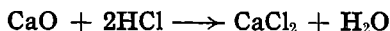


In making NaCl by this method, a small quantity of sodium hydroxide solution is placed in a dish with a piece of blue litmus paper. Hydrochloric acid is then added drop by drop until the litmus paper just turns red. If this neutral solution is now evaporated to dryness, a white deposit of sodium chloride remains, and may be identified by its salty taste.

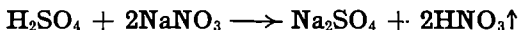
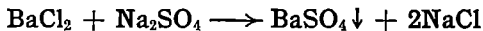
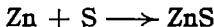
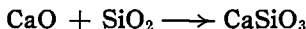
2. *Action of an Acid on a Metal*



3. *Action of an Acid on the Oxide of a Metal*





4. *Action of an Acid on the Salt of a More Volatile Acid*5. *Double Replacement Resulting in the Formation of an Insoluble Product*6. *Direct Union of the Elements*7. *Union of a Metallic Oxide and a Non-Metallic Oxide***Types of Salts**

1. An *acid salt* is one in which only a part of the hydrogen of the acid has been replaced by a metal. Thus, sodium acid carbonate, or sodium bicarbonate ( $\text{NaHCO}_3$ ), and sodium acid phosphate ( $\text{Na}_2\text{HPO}_4$ ) are acid salts.

2. A *basic salt* is one which contains one or more hydroxyl (OH) radicals. Thus, bismuth subnitrate [ $\text{Bi}(\text{OH})_2\text{NO}_3$ ] is a basic salt used in medicine.

3. A *normal salt* is one which contains only a metal or a metallic radical, combined with a non-metal or an acid radical. It contains neither hydrogen replaceable by a metal, nor a hydroxyl radical. Thus, calcium carbonate ( $\text{CaCO}_3$ ) and sodium chloride ( $\text{NaCl}$ ) are normal salts.

4. A *double salt* is one which contains two metals combined with one acid radical. Thus common alum is a double sulfate of potassium and aluminum, having the formula  $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$  or  $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ .

**NAMING ACIDS, BASES, AND SALTS**

**Acids.**—A *binary acid* is one composed of only *two* elements, namely, hydrogen and a non-metal. Such an acid is named by using the stem, or full name, of the non-metal with the prefix *hydro-* and the ending *-ic*. *Examples:* hydrochloric acid ( $\text{HCl}$ ); hydrosulfuric acid ( $\text{H}_2\text{S}$ ).

A *ternary acid* is one composed of *three* elements, namely, hydrogen, a non-metal, and (usually) oxygen. In most cases the three elements are able to combine in various proportions to produce several acids. The name of the *most common* acid of the

series consists of the stem of the non-metal with the ending *-ic*. The acid containing *one less atom of oxygen* has the ending *-ous*. The acid containing *still less oxygen* has the prefix *hypo-* and the ending *-ous*. The acid containing *one more atom of oxygen* than the *-ic* acid has the prefix *per-* and the ending *-ic*. The following table of the acids of chlorine and sulfur illustrates the method of nomenclature:

ACIDS OF CHLORINE	FORMULAS	ACIDS OF SULFUR	FORMULAS
Hydrochloric acid	HCl	Hydrosulfuric acid	H <sub>2</sub> S
Hypochlorous acid	HClO	Hyposulfurous acid	H <sub>2</sub> SO <sub>2</sub>
Chlorous acid	HClO <sub>2</sub>	Sulfurous acid	H <sub>2</sub> SO <sub>3</sub>
Chloric acid	HClO <sub>3</sub>	Sulfuric acid	H <sub>2</sub> SO <sub>4</sub>
Perchloric acid	HClO <sub>4</sub>	Persulfuric acid	H <sub>2</sub> SO <sub>5</sub>

**Bases.**—Bases are named by placing the name of the metal before the word "hydroxide." *Examples:* sodium hydroxide (NaOH); calcium hydroxide [Ca(OH)<sub>2</sub>]. Strong bases such as sodium hydroxide and potassium hydroxide are often called *alkalies*.

**Salts.**—The name of a salt consists of two parts: the name of the metal, and the name derived from the acid.

The name of a *binary salt* (derived from a binary acid) consists of the name of the metal and the name of the non-metal, with the ending *-ide*. *Examples:* sodium chloride is a salt of hydrochloric acid; calcium sulfide is a salt of hydrosulfuric acid.

The name of a *ternary salt* (derived from a ternary acid) consists of the name of the metal and the name of the non-metal, with the ending *-ate* or *-ite*. The ending *-ate* is used if the name of the acid ends in *-ic*; the ending *-ite* is used if the name of the acid ends in *-ous*. If the name of the ternary acid has a prefix, such as *hypo-* or *per-*, that prefix is retained in the name of the salt. The following table illustrates the method of naming the salts of the chlorine acids:

ACIDS OF CHLORINE	FORMULAS	SODIUM SALTS OF THESE ACIDS	FORMULAS
<i>Hydrochloric acid</i>	HCl	Sodium chloride	NaCl
<i>Hypochlorous acid</i>	HClO	Sodium hypochlorite	NaClO
<i>Chlorous acid</i>	HClO <sub>2</sub>	Sodium chlorite	NaClO <sub>2</sub>
<i>Chloric acid</i>	HClO <sub>3</sub>	Sodium chlorate	NaClO <sub>3</sub>
<i>Perchloric acid</i>	HClO <sub>4</sub>	Sodium perchlorate	NaClO <sub>4</sub>

## QUESTIONS

1. Define: (a) acid; (b) base; (c) salt. Give the characteristics of each.
2. Classify the following as oxides, acids, bases, or salts:  $\text{FeCl}_3$ ,  $\text{HNO}_3$ ,  $\text{LiF}$ ,  $\text{H}_2\text{S}$ ,  $\text{HClO}$ ,  $\text{AlPO}_4$ ,  $\text{Fe}_2\text{O}_3$ ,  $\text{K}_2\text{CO}_3$ ,  $\text{Na}_2\text{SO}_3$ ,  $\text{Zn}(\text{OH})_2$ .
3. (a) What is the anhydride of sulfuric acid? (b) Carbon dioxide is the anhydride of what acid?
4. (a) Write the formulas of a common acid and of a common base. (b) Write the equation for the reaction which occurs when solutions of these two compounds are mixed.
5. Chlorine, bromine, iodine, fluorine, and sulfur form acids which do not contain oxygen. (a) Give the names and formulas of these acids. (b) Give the names and formulas of their potassium, zinc, and aluminum salts.
6. Give the names and formulas of the six salts formed by the action of the three most active acids on sodium and magnesium.
7. Give the name and formula of the acid from which each of the following salts is derived: sodium carbonate, calcium phosphate, lead bromide, ferric nitrate, potassium sulfite, ammonium hypochlorite, copper iodide, sodium persulfate, barium nitrite, calcium sulfide.
8. In an experiment on neutralization, how can one tell whether to add more acid or more base to complete neutralization?
9. Give the names of the following salts and explain the derivation of each name:  $\text{KBr}$ ,  $\text{PbSO}_4$ ,  $\text{CaSO}_3$ ,  $\text{NaClO}$ ,  $\text{K}_3\text{PO}_4$ ,  $\text{K}_3\text{PO}_3$ ,  $\text{MgI}_2$ ,  $(\text{NH}_4)_2\text{S}$ ,  $\text{AgCl}$ ,  $\text{CaCO}_3$ .
10. In preparing each of the following salts by neutralization, name the acid and the base required, and write the equation for the reaction involved: (a) sodium nitrate ( $\text{NaNO}_3$ ); (b) sodium sulfate ( $\text{Na}_2\text{SO}_4$ ); (c) calcium sulfate ( $\text{CaSO}_4$ ); (d) sodium phosphate ( $\text{Na}_3\text{PO}_4$ ); (e) aluminum chloride ( $\text{AlCl}_3$ ).

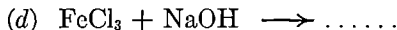
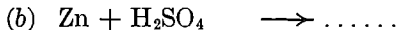
## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. An acid is a compound containing . . . . . which is replaceable by a . . . . .
2. A base is a compound consisting of a . . . . . and a . . . . .
3. Neutralization is the action between an . . . . . and a . . . . ., resulting in the formation of a . . . . . and . . . . .
4. A salt is formed by the union of the . . . . . of an acid with the . . . . . of a base.

5. Copper chloride may be prepared by the action of ..... on copper oxide. The other product formed is .....

6. Complete and balance the following equations:



7. Household lye is the chemical substance .....

8. The names and formulas of the sodium salts of three acids of sulfur are .....

9. An acid consisting of hydrogen, a non-metal, and oxygen is called a ..... acid.

10. Because calcium oxide combines with ..... to form a base, it is called a .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

A	B
( ) Vinegar	1. The most abundant oxide
( ) Has a soapy feel	2. $\text{NaHSO}_4$
( ) $\text{H}_2\text{O}$	3. Neutralization
( ) A normal salt	4. Phosphorus pentoxide
( ) Ammonia water	5. Contains acetic acid
( ) $\text{H}_2\text{S}$	6. A basic anhydride
( ) $\text{SO}_3$	7. Sodium hydroxide
( ) An acid salt	8. Contains citric acid
( ) $\text{K}_2\text{O}$	9. A binary acid
( ) Acid + base	10. Turns litmus blue
	11. $\text{MgSO}_4$
	12. An acid anhydride

## CHAPTER 9

### THE HALOGENS .

**Importance.**—The halogen family consists of four elements: fluorine, chlorine, bromine, and iodine. The name “halogen,” which means *salt-former*, is applied to these four elements because they combine with metals to form compounds that resemble common salt. The halogen group is a striking example of a *chemical family*, the members of which show a very regular and definite gradation of both physical and chemical properties. The table below summarizes the principal characteristics of these four elements, which we shall consider later in greater detail:

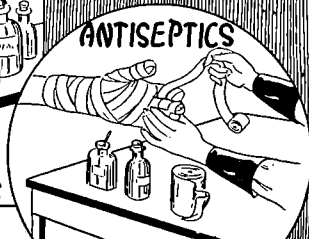
PROPERTIES	FLUORINE	CHLORINE	BROMINE	IODINE
Atomic weight	19	35.5	80	127
Physical state	Gas	Gas	Liquid	Solid
Color	Pale-yellow	Greenish-yellow	Dark-red	Purplish-black
Solubility in water	Very soluble (decomposes water violently)	Fairly soluble	Less soluble	Least soluble
Chemical activity	Most active	Less active	Still less active	Least active
Replacement	Replaces Cl, Br, I	Replaces Br, I	Replaces I	Replaces none
Reaction with hydrogen	Rapid in cold and dark	Rapid in strong sunlight	Only when heated	Slow even if heated
Formula of hydrogen acid	HF	HCl	HBr	HI
Stability of hydrogen acid	Most stable	Less stable	Still less stable	Least stable

# The Halogens & their Uses

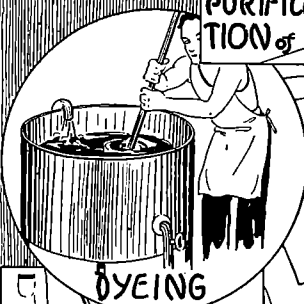
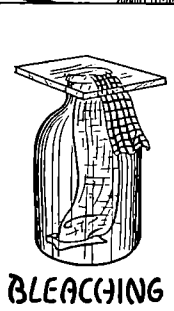
SALT



ANTISEPTICS



DRUGS AND MEDICINES



ETHYL GASOLINE

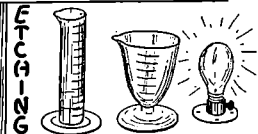
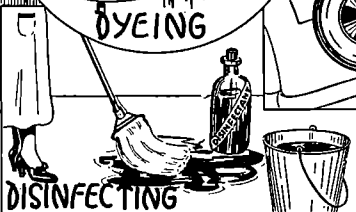


FIG. 25.

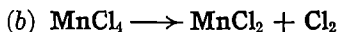
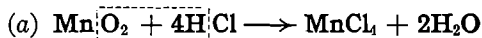
Both as elements and in their compounds, the halogens are of great importance to man. Chlorine, the chief member of the family, is a valuable bleaching and disinfecting agent. In modern warfare, chlorine, as a constituent of poison gas, destroys life; while, in the form of antiseptic solutions, it preserves life. Sodium chloride, the most abundant halogen compound, occurs in vast quantities in the oceans of the earth. Hydrochloric acid, the most important halogen compound made by man, is used extensively in chemical industries. The other three halogen elements, together with their compounds, are employed widely in the preparation of drugs, dyes, medicines, antiseptics, bleaching agents, and photographic chemicals.

### CHLORINE

**Occurrence.**—Chlorine is far too active to exist free in nature. In the combined state, it is found principally as sodium chloride and other metallic chlorides in the ocean, salt lakes, salt wells, and beds of solid salt.

#### Preparation

1. *Oxidation of Hydrochloric Acid by Manganese Dioxide (Laboratory Method).*—About one-half a test tubeful of manganese dioxide is placed in a flask (Fig. 26) and about a test tubeful of concentrated hydrochloric acid is added through the thistle tube. The mixture is warmed gently. Chlorine is liberated and is collected by downward displacement of air. The reaction may be regarded as proceeding in two stages:



By combining the two, we obtain the equation:



Chlorine, unlike oxygen and hydrogen, cannot be collected over water, because it is soluble in water. Since chlorine is much heavier than air, it is collected by the downward displacement of air, as shown in the diagram. Frequently, a mixture of salt, sulfuric acid, and manganese dioxide is used in the preparation of chlorine. The reaction of salt and sulfuric acid yields hydro-

chloric acid, which is in turn oxidized by the manganese dioxide, as above, yielding chlorine.

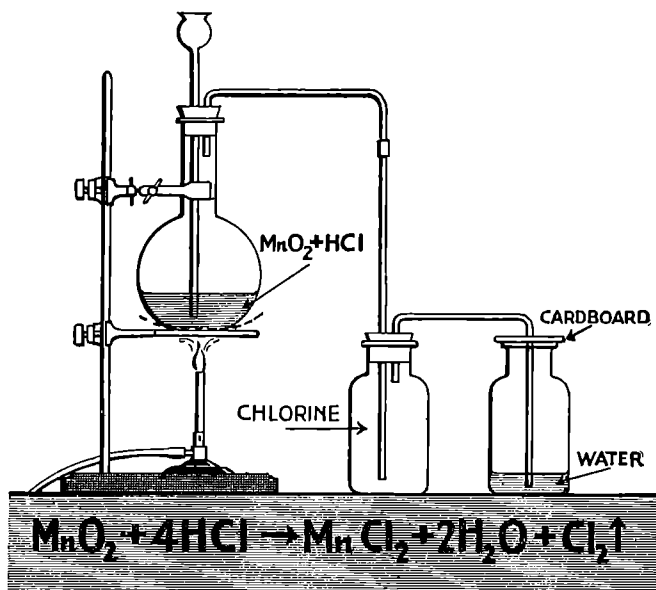
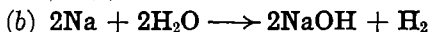
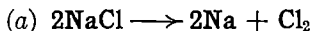
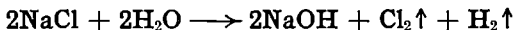


FIG. 26. Laboratory preparation of chlorine.

2. *Electrolysis of Sodium Chloride Solution (Commercial Method).*—An electric current is passed through a concentrated salt solution. Free chlorine is liberated at the anode. Sodium is liberated at the cathode, but reacts immediately with the water, forming sodium hydroxide and liberating hydrogen. This process also serves as a commercial method for making the important base, sodium hydroxide. The two equations are:



The combined equation is:



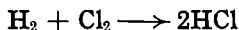
**Physical Properties.**—Chlorine is a greenish-yellow gas with a strong, irritating odor. When inhaled, it attacks the delicate membranes of the nose, throat and lungs. Chlorine is extremely



poisonous. A deep breath of the pure gas may cause death, although it is used in minute quantities in the treatment of colds. Chlorine is about  $2\frac{1}{2}$  times as heavy as air. It is somewhat soluble in water, one volume of water dissolving about two volumes of gas to form chlorine water, a yellowish liquid. Chlorine is easily liquefied by pressure alone at room temperature.

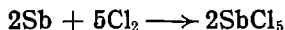
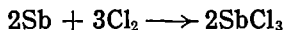
**Chemical Properties.**—Most of the chemical properties of chlorine depend on its extreme activity. It combines directly with many elements, forming chlorides. In most cases, the reaction is so vigorous that heat and light are produced. Such reactions are called *combustions*. Therefore chlorine may be said to support combustion.

1. *Action with Hydrogen.*—When a jet of burning hydrogen is thrust into a bottle of chlorine, it continues to burn, forming hydrogen chloride:

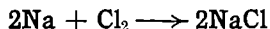


A mixture of hydrogen and chlorine does not react when kept in the *dark*. In *diffused sunlight*, the two gases unite slowly. In *direct sunlight*, they unite with explosive violence.

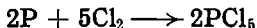
2. *Action with Metals.*—When a little antimony is sprinkled in a bottle of chlorine, the metal catches fire, forming antimony trichloride and antimony pentachloride:



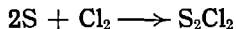
When a bit of sodium is heated and put into chlorine, the elements unite with a brilliant flame, forming sodium chloride:



3. *Action with Non-Metals.*—Phosphorus unites vigorously with chlorine, forming phosphorus pentachloride:

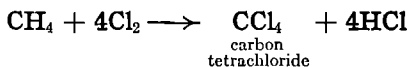
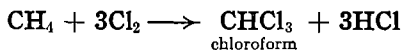


Sulfur burns in a limited supply of chlorine, forming sulfur monochloride:

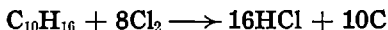


4. *Action with Hydrocarbons.*—Chlorine reacts with hydrocarbons (compounds of hydrogen and carbon), replacing part or all of the hydrogen. With methane ( $\text{CH}_4$ ), four different products

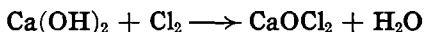
may be obtained, depending on the number of hydrogen atoms replaced by chlorine. Two examples are:



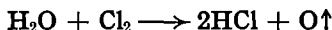
If a piece of filter paper is saturated with turpentine ( $\text{C}_{10}\text{H}_{16}$ ) and thrust into a bottle of warm chlorine, the hydrogen and chlorine unite with a flash of flame, and a cloud of black smoke (carbon) is evolved:



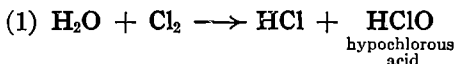
5. *Action with Slaked Lime.*—Chlorine reacts with slightly moist slaked lime to form bleaching powder (sometimes called “chloride of lime”):



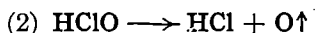
6. *Action with Water.*—Chlorine reacts with water, liberating nascent (atomic) oxygen:



**Bleaching Action of Chlorine.**—When chlorine dissolves in water, the reaction may be represented in two stages:



Hypochlorous acid is a very unstable compound, easily decomposed by sunlight into hydrochloric acid and nascent oxygen:



If we put a piece of dry colored calico in a bottle of dry chlorine, no bleaching results. If, however, either the calico or the chlorine is moist, the color of the calico disappears. The presence of water is therefore essential in order to bleach with chlorine. The bleaching effect results from the action of nascent oxygen on the dye, a colorless compound being formed.

**Uses.**—Most of the uses of chlorine depend on its powerful oxidizing properties, either in its free state, or in its unstable compounds, which decompose, liberating atomic oxygen. Chlorine is used in the following ways:

1. *As a Bleaching Agent.*—This is the chief use of chlorine. Nearly 90% of all the chlorine produced is employed to bleach cotton and wood pulp. Silk and wool are bleached by other methods, because chlorine destroys their fibers.

2. *In Purifying Water.*—Chlorine is an efficient destroyer of disease-producing bacteria in water. About four pounds of chlorine per million gallons of water are sufficient to render city water supplies and swimming pools safe.

3. *As an Antiseptic.*—Dakin's solution and Javelle water are essentially weak solutions of sodium hypochlorite ( $\text{NaClO}$ ). This compound decomposes slowly, liberating nascent oxygen, which destroys and stops the growth of toxic bacteria, particularly in open wounds.

4. *As a Disinfectant.*—Bleaching powder or "chloride of lime" ( $\text{CaOCl}_2$ ) is often used in the household to oxidize waste organic materials and to destroy bacteria. It also serves as a convenient bleaching agent.

5. *In Manufacturing Chlorine Compounds.*—Chloroform ( $\text{CHCl}_3$ ), carbon tetrachloride ( $\text{CCl}_4$ ), and sulfur monochloride ( $\text{S}_2\text{Cl}_2$ ) are among the many valuable compounds made by the use of chlorine.

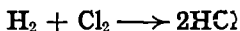
6. *In Warfare.*—Free chlorine was the first poison gas used in warfare. It is a constituent of mustard gas, phosgene, and chlorpicrin, all well-known toxic gases employed in the World War.

### HYDROCHLORIC ACID ( $\text{HCl}$ )

**Occurrence.**—Hydrochloric acid is a water solution of hydrogen chloride, which is the most important halogen compound artificially made. It occurs as free hydrogen chloride in the vapors erupted by certain volcanoes. It is also found as hydrochloric acid in the gastric juice of the stomach in a concentration of about 0.3 per cent.

#### Preparation

1. *Direct Union of Hydrogen and Chlorine.*—This method is now employed on a large scale in the commercial manufacture of pure hydrochloric acid:



2. *Action of Sulfuric Acid on Sodium Chloride (Laboratory and Commercial Method).*—A mixture of concentrated sulfuric acid

and sodium chloride is heated. Hydrogen chloride is readily evolved and is collected by downward displacement of air (Fig. 27). If the gas is allowed to flow into some water in a bottle, as shown in the diagram, it dissolves in the water, forming hydrochloric acid. The reaction which takes place in the preparation of hydrogen chloride may proceed in either of two ways, depending on the temperature:

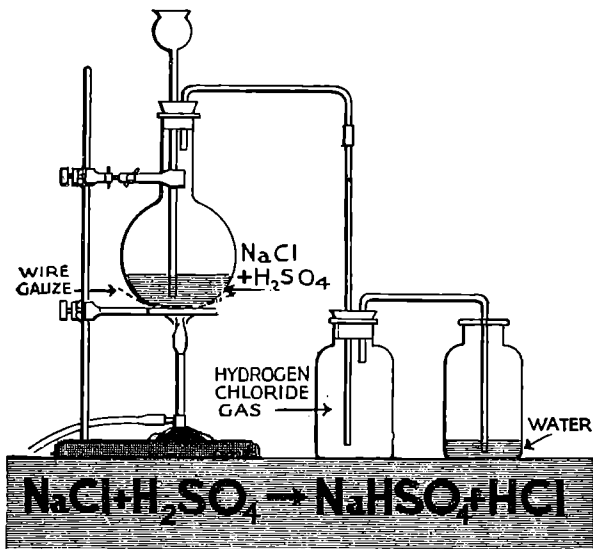
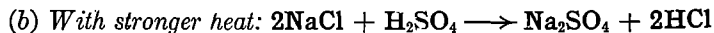


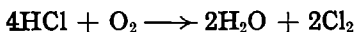
FIG. 27. Laboratory preparation of hydrogen chloride.

This preparation illustrates the general method of making acids, namely, the action of sulfuric acid on the salt of the desired acid. In preparing hydrochloric acid, sodium chloride is used because it is the cheapest source of chlorine. Sulfuric acid is used because: (a) it has the highest boiling point of the most common acids (338°C.), and therefore remains behind with the other substances in the reaction, while the hydrogen chloride is driven off by gentle heating; (b) it is a source of the hydrogen needed to

make the acid; (c) it is a dehydrating agent, removing the water from the gas before it is driven off; and (d) it is cheap.

**Physical Properties of Hydrogen Chloride.**—Hydrogen chloride is a colorless gas having a sharp, choking odor. It is heavier than air and is easily liquefied and solidified. The gas is extremely soluble in water, one volume of water dissolving about 500 volumes of the gas (Fig. 28). Hydrogen chloride absorbs and condenses the moisture of the air, causing fumes which consist of minute droplets of hydrochloric acid.

**Chemical Properties of Hydrogen Chloride.**—Hydrogen chloride is a stable compound, which neither burns nor supports combustion. The dry gas does not possess acid properties. It can be oxidized to form water and chlorine:

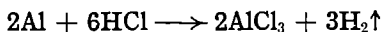
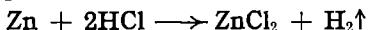


**Physical Properties of Hydrochloric Acid.**

—Chemically pure hydrochloric acid is a colorless solution of pure hydrogen chloride in distilled water. The commercial variety, known as *muratic acid*, contains 30% hydrogen chloride by weight, and is yellow in color, owing to impurities. The concentrated acid, which is poisonous, fumes strongly in moist air.

**Chemical Properties of Hydrochloric Acid.**

—Hydrochloric acid is a typical acid, and exhibits all the characteristics of that group of compounds. It turns blue litmus red, has a sour taste, and reacts with many metals, forming chlorides and liberating hydrogen:



Hydrochloric acid neutralizes bases, forming salts and water:

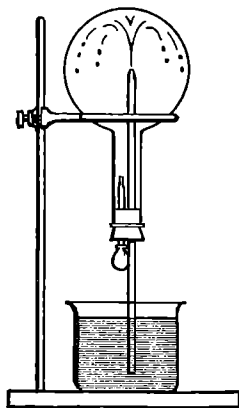
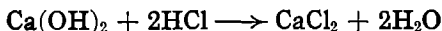
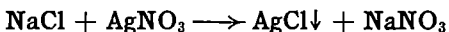


FIG. 28. The hydrogen chloride fountain demonstrates the extreme solubility of the gas. The flask is filled with dry hydrogen chloride and set up as shown above. The medicine dropper, which contains some water, is pinched, thus forcing out a few drops of water. This dissolves so much of the gas that a partial vacuum is created. The outside air pressure then forces the water up through the tube in the form of a fountain.

**Uses.**—Hydrochloric acid has a variety of uses, both in the laboratory and commercially. Its chief industrial use is in cleaning (“pickling”) metals before they are tinned or galvanized. It is employed in making glucose, glue, gelatine, dyes, chlorides, and a variety of other substances. It is also used in preparing *aqua regia*, a mixture of hydrochloric and nitric acids, which has the power of dissolving gold and platinum.

**Chlorides.**—A chloride is a salt of hydrochloric acid. It contains a metal, or a metallic radical, combined with chlorine, as zinc chloride ( $\text{ZnCl}_2$ ) and ammonium chloride ( $\text{NH}_4\text{Cl}$ ). Ordinary table salt is sodium chloride ( $\text{NaCl}$ ). All the common chlorides are soluble in water, with the exception of silver chloride, mercurous chloride, and lead chloride.

**Test for a Chloride.**—A distilled water solution of silver nitrate ( $\text{AgNO}_3$ ) is added to the solution to be tested. If, for example, sodium chloride is present in the solution, the following reaction takes place:



Silver chloride settles as a white precipitate which remains unchanged on the addition of dilute nitric acid. This test is so sensitive that it often shows the presence of chlorides in ordinary drinking water.

## BROMINE

**Occurrence.**—Bromine is never found free in nature. In the combined form, it occurs in a variety of compounds. Large deposits of crude sodium chloride, containing  $\text{NaBr}$  and  $\text{MgBr}_2$ , are found scattered in various parts of the United States. Vast beds of magnesium bromide are located at Stassfurt, Germany. Ocean water is known to contain considerable quantities of bromine salts. Very recently a large plant was erected at Cape Fear, N. C. to extract bromine from the sea on a commercial scale. This plant draws up over 100,000 tons of sea-water each day, from which it extracts about  $7\frac{1}{2}$  tons of bromine.

### Preparation

1. *Laboratory Method.*—Bromine is prepared in a manner similar to that of chlorine, namely, by the oxidation of its hydrogen compound, manganese dioxide being used as the oxidizing agent. A small quantity of sodium bromide is mixed with an equal amount of manganese dioxide and put into a dry test tube. A little concen-

trated sulfuric acid is added, and the mixture is heated gently. Bromine vapor is evolved, which may be cooled and condensed directly to the liquid state, or passed into water as shown in Fig. 29. Some of the bromine dissolves, while the rest collects as a reddish liquid below the water. The equation is:

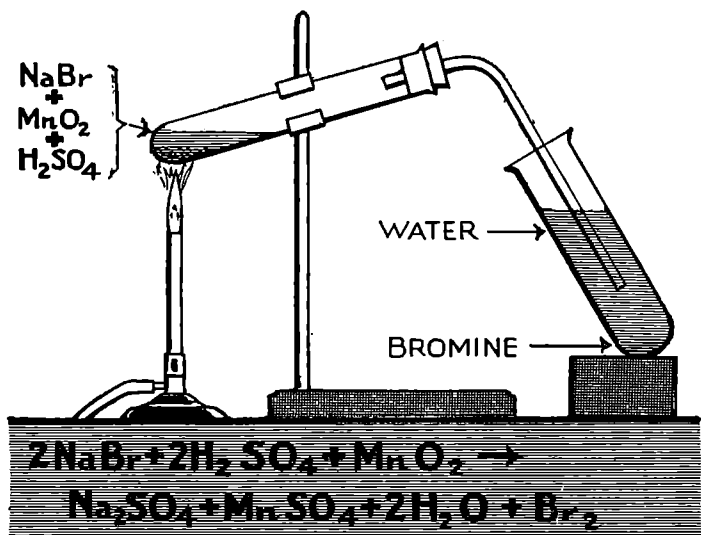
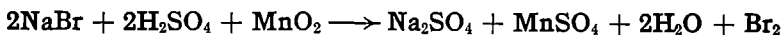
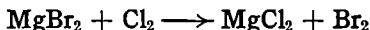


FIG. 29. Laboratory preparation of bromine.

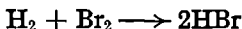
2. *Commercial Method.*—Since chlorine is more active than bromine, it replaces bromine from its salts. A solution of crude magnesium bromide is treated with chlorine, and the liberated bromine is distilled with steam. The equation is:



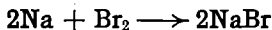
**Physical Properties.**—Bromine is a dark-red liquid about three times as heavy as water. It has a disagreeable odor (the word "bromine" means *stench*). The liquid vaporizes readily, so that a partly filled bottle always contains bright red bromine gas in the upper portion. It is somewhat soluble in water, forming reddish bromine water. It is readily soluble in certain organic solvents, such as carbon disulfide, yielding an orange-red solution. Bromine

is poisonous in both the liquid and the gaseous form. The liquid is very corrosive in its action on the skin.

**Chemical Properties.**—Bromine is less active than chlorine. It combines with hydrogen, forming hydrogen bromide:



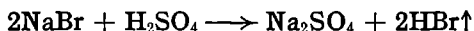
It also combines with many elements, usually metals, forming bromides:



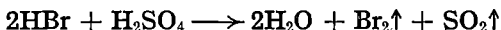
In water solution, bromine is a mild oxidizing and bleaching agent.

**Uses.**—Bromine is employed in making potassium bromide (KBr), which is used as a medicine; and in the preparation of silver bromide (AgBr), which is used in photography. A variety of drugs and dyes contain bromine in combination with other elements. Ethylene dibromide, used in making "ethyl" or "anti-knock" gasoline, is an important compound of bromine. The element is also present in "tear gas," used by police in dispersing mobs. In the laboratory, bromine is used as a bleaching and oxidizing agent.

**Hydrogen Bromide.**—Hydrogen bromide, like hydrogen chloride, may be prepared by the action of concentrated sulfuric acid on the corresponding sodium salt:



However, hydrogen bromide is less stable than hydrogen chloride and some of it is oxidized by the hot, concentrated sulfuric acid, yielding bromine and sulfur dioxide:



If phosphoric acid is used in place of sulfuric acid, the hydrogen bromide is not oxidized. Hydrogen bromide resembles hydrogen chloride in that it is a colorless gas, very soluble in water, and fuming in moist air. When dissolved in water, it forms hydrobromic acid, which exhibits all the properties of a strong acid.

## IODINE

**Occurrence.**—Iodine occurs in sea water and in seaweed as *iodides* (NaI, KI, and MgI<sub>2</sub>), and in Chile saltpeter deposits as *iodates* (NaIO<sub>3</sub>).



**Preparation**

1. *Laboratory Method.*—The same reaction is used to prepare iodine as is used for both chlorine and bromine. A mixture of equal parts of potassium iodide and manganese dioxide is placed in a dish, and some concentrated sulfuric acid is added (Fig. 30).

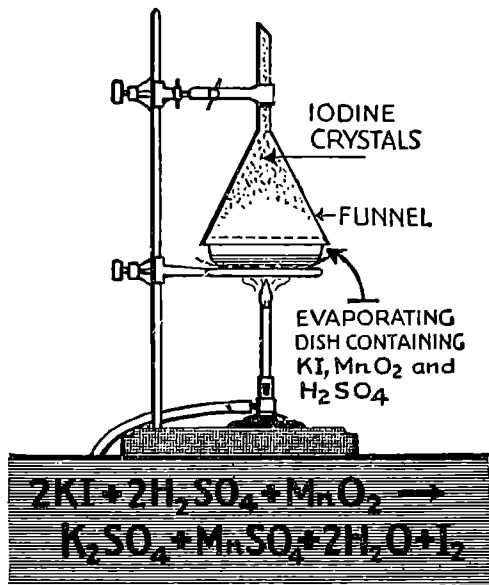


FIG. 30. Laboratory preparation of iodine.

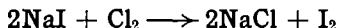
The mixture is heated gently. The iodine is evolved as a violet vapor which condenses as purplish-black crystals on the inner surface of a glass funnel fastened over the dish:



This method of obtaining crystals of pure iodine is called *sublimation*, which is the process of converting a substance from the solid state directly to the vapor state and then back to the solid state, without its passing through the intermediate liquid state. Very few solids behave in this way.

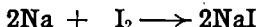
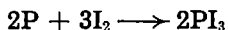
2. *Commercial Method.*—Seaweed (kelp) is burned, and the ashes, which contain the iodides of sodium and potassium, are

treated with water to extract these compounds. The iodine is then obtained by displacement with chlorine:

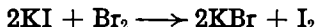
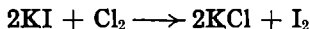


**Physical Properties.**—Pure iodine is a crystalline steel-gray solid; in air, however, this color turns quickly to purplish black. It is slightly soluble in water, more soluble in a water solution of potassium iodide, and very soluble in alcohol. The alcohol solution is called *tincture of iodine*. Iodine is also soluble in carbon tetrachloride, chloroform, and carbon disulfide, yielding a violet solution in each case. Iodine vaporizes readily, liberating fumes of a striking violet color. It undergoes sublimation, as explained above, in the laboratory preparation of the element.

**Chemical Properties.**—Iodine combines with many elements, forming iodides:



Iodine is the least active of the halogens, and is therefore replaced by the others:

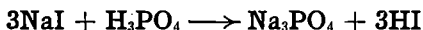


When iodine is mixed with starch paste, a blue-black color is produced. This is used as a test for either starch or free iodine.

**Uses.**—The chief commercial use of iodine is in the making of the valuable antiseptic, tincture of iodine. Iodoform ( $\text{CHI}_3$ ) is another important antiseptic. Iodine is used to produce various iodides that are employed in medicine and photography, and to make dyes and drugs.

Iodine is also an important constituent of the human body. The element is found in thyroxin, a complex substance secreted by the thyroid gland. A deficiency of this iodine compound in the body results in definite disorders, such as goitre and cretinism. To overcome these conditions, iodine, in the form of sodium iodide, is often added to the diet. Synthetic thyroxin also is now widely used in treating disorders of the thyroid gland.

**Hydrogen Iodide.**—The preparation of hydrogen iodide is similar to that of hydrogen chloride and hydrogen bromide. However phosphoric acid is used instead of sulfuric acid:

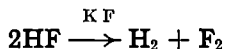


Sulfuric acid, cannot be employed to prepare hydrogen iodide because this latter compound is oxidized even more readily than hydrogen bromide, yielding iodine, sulfur dioxide, and also some hydrogen sulfide. Hydrogen iodide is a colorless gas that fumes in moist air. It dissolves readily in water, forming hydriodic acid, which is a strong acid, although less stable than the corresponding compounds of the other halogen elements.

## FLUORINE

**Occurrence.**—Fluorine, the most active of all the elements, occurs abundantly in the earth as calcium fluoride or *fluorspar* ( $\text{CaF}_2$ ), and as *cryolite* ( $\text{Na}_3\text{AlF}_6$ ).

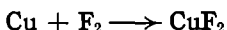
**Preparation.**—Fluorine is prepared by passing an electric current through a solution of potassium fluoride in liquid hydrogen fluoride:



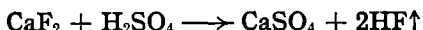
Because the element attacks metals and glass, the electrolysis must be carried out in a graphite cell with a carbon anode.

**Physical Properties.**—Fluorine is a pale-yellow gas, somewhat heavier than air, and with a pungent odor. It fumes in moist air, and is extremely poisonous.

**Chemical Properties.**—Fluorine is the most active of the halogens. It decomposes water at ordinary temperatures. It combines vigorously with most elements, forming fluorides:



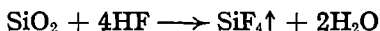
**Hydrogen Fluoride.**—Hydrogen fluoride, a colorless gas, is the most stable and most active of the hydrogen halides. It is made by the action of sulfuric acid with fluorspar:



Both the pure compound and its water solution (hydrofluoric acid) attack glass, and must therefore be stored in paraffin bottles.

**Etching of Glass.**—Hydrofluoric acid is the only acid that readily attacks silicon dioxide ( $\text{SiO}_2$ ) and various compounds of silicon, such as quartz and glass. Upon this property, depends the chief use of hydrofluoric acid. In order to etch glass, the object is first coated with a layer of paraffin. The design is then scratched

through the paraffin with a sharply pointed instrument, thus exposing the glass. The prepared surface is subjected for a short time to the action of either hydrofluoric acid solution or hydrogen fluoride vapors. The paraffin is then removed, leaving the design cut into the glass. The reaction is:



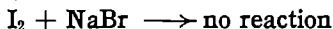
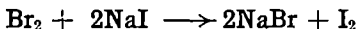
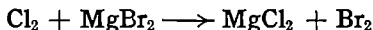
Silicon tetrafluoride ( $\text{SiF}_4$ ) is a gas, and passes off.

### REPLACEMENT POWER OF THE HALOGENS

**Order of Activity.**—The table on page 85, summarizing the properties of the four halogen elements, indicates a definite gradation in these properties. The order of activity is:



This is also the order of replacement, which means that any element will liberate the elements to the right of it from their compounds. Thus:



**Tests for Bromides and Iodides.**—We make use of this order of replacement, together with the solubility of the halogens in certain organic solvents, such as chloroform, carbon disulfide, and carbon tetrachloride, as a basis for identifying halogen salts.

1. *Test for a Bromide:* Dissolve the substance being tested in water in a test tube. Add a little carbon tetrachloride, then a few drops of chlorine water. Shake the tube vigorously. If a bromide is present, the chlorine displaces the bromine, which dissolves in the carbon tetrachloride. This solution, *brownish-red* in color, soon settles to the bottom.

2. *Test for an Iodide.* Treat the substance as above. If an iodide is present, the chlorine displaces the iodine, and the carbon tetrachloride assumes a *violet* color.

## QUESTIONS

1. What two chemical actions are involved in the preparation of chlorine by heating a mixture of sodium chloride, manganese dioxide, and sulfuric acid?
2. Account for the black smoke formed when a burning candle is put into a bottle of chlorine.
3. Explain why cloth must be wet in order to be bleached by chlorine.
4. Write the equations for the preparation of  $\text{SbCl}_3$  and  $\text{PCl}_3$  by direct union.
5. (a) Write the word equation for the general method of preparing the halogen acids. (b) What are the disadvantages of this method?
6. Why does a solution of hydrogen iodide turn dark on standing?
7. Why is hydrofluoric acid sold in wax or wax-lined containers?
8. Explain why the halogen elements do not occur free in nature.
9. When the fumes from a bottle of concentrated hydrochloric acid come in contact with ammonia gas ( $\text{NH}_3$ ), a white smoke is produced. Explain.
10. Iodine crystals have accidentally been mixed with sand. Explain how you would recover the iodine in a pure state.
11. An unmarked bottle is known to contain either sodium iodide or sodium bromide. Describe the test you would use to identify the salt.
12. State in which of the following a reaction would take place:
 

(a) $\text{NaCl} + \text{Br}_2$	(d) $\text{KCl} + \text{I}_2$
(b) $\text{KBr} + \text{Cl}_2$	(e) $\text{NaI} + \text{Br}_2$
(c) $\text{KI} + \text{Cl}_2$	(f) $\text{NaF} + \text{Cl}_2$
13. Explain how you would etch the graduations on an unmarked glass measuring-cylinder.
14. What human dietary need is supplied by sea-foods, such as clams, oysters, and lobsters?
15. Compare the order of activity of the halogens with the order of stability of the halogen acids.

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. In the electrolysis of sodium chloride solution, . . . . . is liberated at the anode, and . . . . . at the cathode.
2. Two important sources of iodine are . . . . .
3. Hydrochloric acid is a solution of . . . . . in water.

4. To test for a chloride, ..... is added to the solution. If a ..... is formed, which is insoluble in ....., the substance is a chloride.

5. Hydrochloric acid is found in the human body in the ..... Its function is .....

6. Two fabrics which should not be bleached by chlorine are .....

7. The newest commercial source of bromine is ..... The bromine is liberated from bromides by the action of .....

8. The chief commercial use of iodine is in the making of .....

9. When chlorine water is added to a solution of potassium iodide, free ..... is liberated. This is soluble in carbon tetrachloride, yielding a ..... colored solution.

10. Two reasons for using sulfuric acid to prepare hydrochloric acid are .....

11. The bleaching power of chlorine is due to its action with ..... and the liberation of .....

12. The commercial name of hydrochloric acid is .....

13. The halogen acid which is most easily oxidized is ..... When this compound is oxidized by  $H_2SO_4$ , 3 products formed are .....

14. When iodine crystals sublime, they change from the ..... state to the ..... state, and then from the ..... state to the ..... state. In both changes, the ..... state is skipped.

15. The halogen compound used to etch glass is ..... This halogen compound is made by the reaction between ..... and .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

A	B'
( ) Sodium + chlorine	1. Used in the test for an iodide.
( ) Hydrogen chloride	2. Oysters
( ) Chlorine water	3. A reddish liquid
( ) Silver bromide	4. Silver chloride
( ) A complex iodine compound	5. Sublimation
( ) Hypochlorous acid	6. Fluorine
( ) The most active element	7. Used in photography
( ) To purify iodine	8. Combustion
( ) An insoluble compound	9. Order of activity
( ) Bromine	10. Extremely soluble in water
	11. Thyroxin
	12. An unstable compound

## CHAPTER 10

### THE ATMOSPHERE

**The Ocean of Air.**—We live at the bottom of a vast ocean of air which extends to a height of about 200 miles. Man's exploration of this huge expanse has been confined to extremely narrow limits. He has ascended about nine miles in an airplane, and about thirteen miles in a balloon, while the greatest height ever reached by a sounding balloon is only about twenty-four miles. The density of our gaseous envelope is greatest near the earth's surface, and gradually becomes thinner as the altitude increases. About half of its weight is within four miles of the earth, although minute traces of air are believed to exist even at a distance of 2000 miles. The weight of the atmosphere produces a pressure of nearly fifteen pounds on every square inch of surface at sea level; the weight of air resting upon every square foot is more than one ton.

**Composition of the Air.**—With the exception of water vapor, the percentage of which may vary between wide limits, the constituents of the air are relatively constant in amount. The average composition of *dry* air is as follows:

GAS	PERCENTAGE BY VOLUME	PERCENTAGE BY WEIGHT
Nitrogen . . . . .	78.00	75.40
Oxygen . . . . .	21.00	23.20
Carbon dioxide . . . . .	.04	.05
Argon . . . . .	.96	1.30
Neon	} . . . . . Traces	} . . . . . Traces
Helium		
Krypton		
Xenon		
Hydrogen		
Ozone		

NOTE.—In addition, the air contains dust particles, bacteria, and minute quantities of certain gases like H<sub>2</sub>S, SO<sub>2</sub>, etc., which are evolved by volcanoes and by various manufacturing processes.

**Air Is a Mixture.**—There are several indisputable evidences to indicate that the constituents of the air are not chemically united, but are merely mixed. These observed facts are:

1. *The composition of air varies.* Although the percentages of oxygen and nitrogen in the air are fairly constant, accurate analyses of different samples of air always show that there are slight variations in composition. Hence air is not a compound, since it does not conform to the Law of Definite Proportions.

2. *The properties of air are those of its constituents.* Air behaves like an artificially prepared mixture of nitrogen and oxygen. It exhibits none of the properties of the compounds of nitrogen and oxygen (see Chapter 11), which are choking or poisonous gases, and altogether incapable of supporting life.

3. *There are no energy changes when the constituents of air are mixed.* The formation of a chemical compound is always accompanied by an evolution or an absorption of heat or light. However, when air is synthesized from its constituent gases, there are no evidences of any such energy changes.

4. *Liquid air has no definite boiling point.* When air is liquefied and then allowed to boil, it does not evaporate at a definite temperature, as does any pure compound. Instead, the nitrogen evaporates first, and the other constituents boil off at higher temperatures.

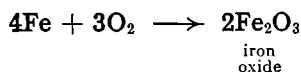
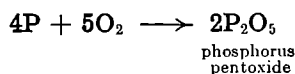
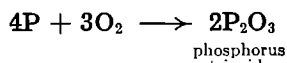
5. *Air that has been dissolved in water contains nearly twice as much oxygen as does ordinary air.* If air were a compound, the proportions of its constituent gases would be the same, whether dissolved in water or not.

6. *Air may be separated into its components by physical means.* When air is forced through a long, unglazed porcelain tube, the nitrogen is found to pass through the porcelain walls of the tube much faster than the oxygen. On the other hand, the oxygen escapes through the rubber of an automobile tire much more rapidly than the nitrogen. If air were a compound, the constituents (combined into molecules) would pass through the porcelain or rubber with the same velocity.

**Percentage of Oxygen in the Air.**—To determine the percentage of oxygen in the air, we cause a measured volume of air, enclosed over water, to remain for several hours in contact with a substance, such as white phosphorus or iron filings, that will combine with



the oxygen and remove it from the air (Fig. 31). The reactions which take place are:



The oxides of phosphorus, both white solids, are slowly absorbed by the water. The iron oxide, reddish-brown in color, adheres

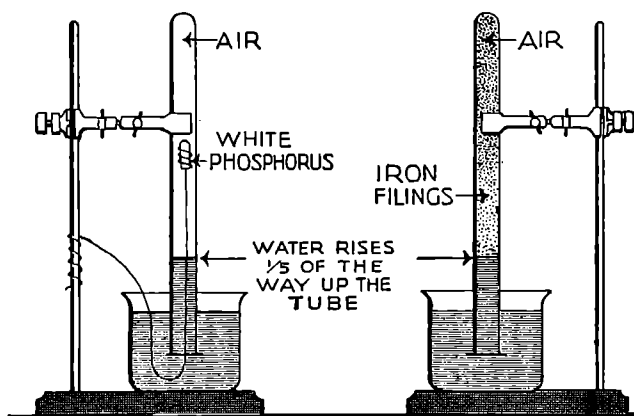


FIG. 31. Two ways to determine the percentage of oxygen in the air.

to the inner surface of the tube. When the action has ceased, the rise in water indicates the volume of the oxygen removed. Following are sample figures for such an experiment:

Volume of air at the start. . . . .	100 cc.
Volume of air at the end. . . . .	79 cc.
Volume of oxygen removed. . . . .	21 cc.
Therefore 21% of the air is oxygen.	

**The Four Essential Constituents of the Air.**—The table on the following page summarizes certain important facts relating to the primary substances that are present in the air.

GAS	PROOF OF PRESENCE	USES
Oxygen	When hydrogen burns in air, <i>water</i> is formed: $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$	For respiration. To support combustion. To aid decay of organic matter.
Nitrogen	When the gas left after the removal of oxygen from an enclosed volume of air is passed over heated magnesium, <i>magnesium nitride</i> is formed: $3\text{Mg} + \text{N}_2 \longrightarrow \text{Mg}_3\text{N}_2$	To dilute the oxygen in the air, and thus retard burning and oxidation. As a plant food.
Carbon dioxide	When clear limewater is exposed to the air, a white crust of <i>calcium carbonate</i> is formed: $\text{Ca}(\text{OH})_2 + \text{CO}_2 \longrightarrow \text{CaCO}_3 + \text{H}_2\text{O}$	As a plant food.
Water vapor	When calcium chloride, a deliquescent substance, is exposed to the air, it <i>absorbs moisture</i> and soon becomes wet.	To prevent excessive evaporation. To produce rain, snow, etc. To aid plant growth.

NOTE.—The dust present in the air is also useful, for each dust particle serves as a nucleus upon which condensation of water vapor takes place. This results in the formation of a raindrop. Dust particles also diffuse or scatter the sunlight, and so bring about the even illumination which we call daylight.

**Occurrence of Nitrogen.**—Nitrogen, the most abundant constituent of the air, is also found combined in a number of compounds. The principal ones are sodium nitrate (Chile saltpeter) and other mineral nitrates which result from the life processes of animal and vegetable organisms. Nitrogen is found abundantly in proteins and other nitrogenous organic compounds present in meat, egg white, peas, beans, clover, etc.

#### Preparation of Nitrogen

1. *Oxidizing Phosphorus in Air.*—The oxygen in the air is removed either by burning phosphorus or by allowing this substance to oxidize slowly, as described above. The gas remaining in the tube after the action has ceased is almost pure nitrogen.

2. *Passing Air over Hot Copper.*—Fine-meshed copper gauze is strongly heated in a hard glass tube and air is passed over it

(Fig. 32). The copper combines with the oxygen of the air, and the nearly pure nitrogen is collected by displacement of water:

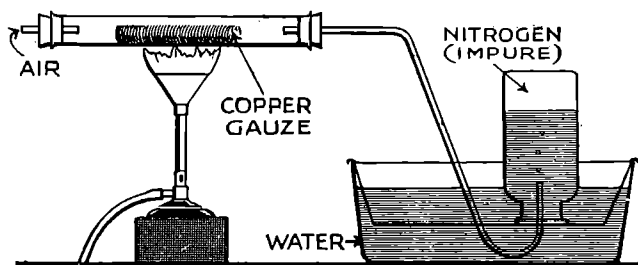
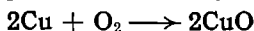
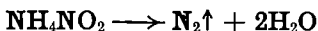
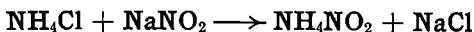


FIG. 32. Laboratory preparation of nitrogen.

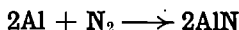
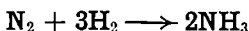
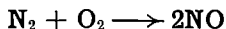
3. *Liquefying Air (Commercial Method).*—Air is reduced to a liquid by extremely high pressure and low temperature. When the pressure is released, nitrogen vaporizes first because of its low boiling point.

4. *Decomposing Ammonium Nitrite (Laboratory Method).*—A mixture of ammonium chloride and sodium nitrite is heated cautiously. The first product formed is ammonium nitrite, which is unstable, and breaks up into nitrogen and water:



**Physical Properties of Nitrogen.**—Nitrogen is a colorless, tasteless, odorless gas, somewhat lighter than air, and slightly soluble in water. It is more difficult to liquefy than oxygen.

**Chemical Properties of Nitrogen.**—Chemically, nitrogen is a rather inactive element. It does not burn or support combustion. At high temperatures, it may be made to combine with oxygen to form nitric oxide, with hydrogen to form ammonia, and with certain metals to form nitrides:



**Uses of Nitrogen.**—The free nitrogen in the air has the highly important function of diluting the oxygen present, thus partly neutralizing the chemical activity of the oxygen. In this manner,

all forms of oxidation, including burning and decay, are slowed down considerably. Free nitrogen is used commercially to fill electric light bulbs. This permits the filament to attain a higher temperature (without disintegrating) than would be possible in a completely evacuated bulb. By means of various "fixation" processes (described in the next chapter), the free nitrogen of the air is converted into a variety of products which enter into the manufacture of fertilizers, explosives, dyes, and drugs.

**The Nitrogen Cycle.**—Nitrogen is needed for plant growth. A few plants, such as beans, peas, clover and alfalfa, can make direct use of the inactive nitrogen of the air, with the aid of nitrogen-fixing bacteria, which are attached to their roots. These micro-

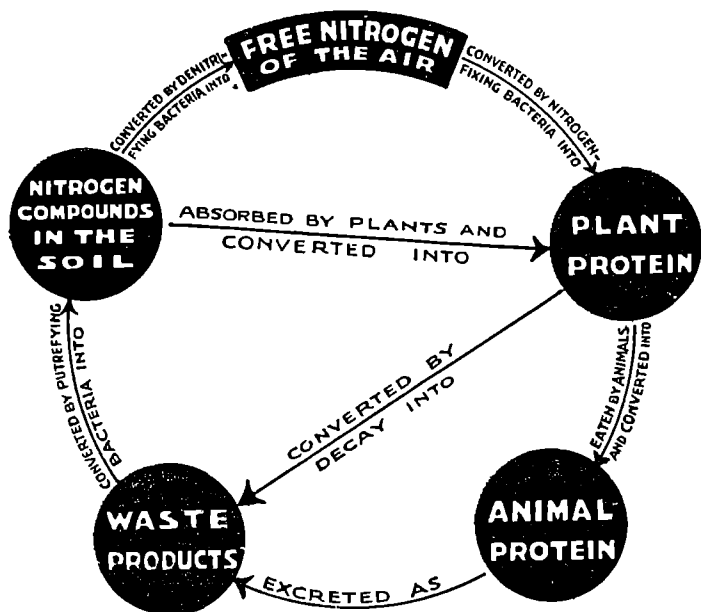


FIG. 33. The nitrogen cycle.

organisms form useful nitrogen compounds, some of which are absorbed by the plants, the rest remaining in the soil to enrich it. This removal of nitrogen from the air by plants is partly balanced by the return of nitrogen to the air during the decay of the unstable proteins in plant and animal matter. Thus, nitrogen follows a complete cycle of transformations in nature, as shown in Fig. 33.

**The Carbon Dioxide Cycle.**—Carbon dioxide is constantly being added to the atmosphere, and, at the same time is constantly

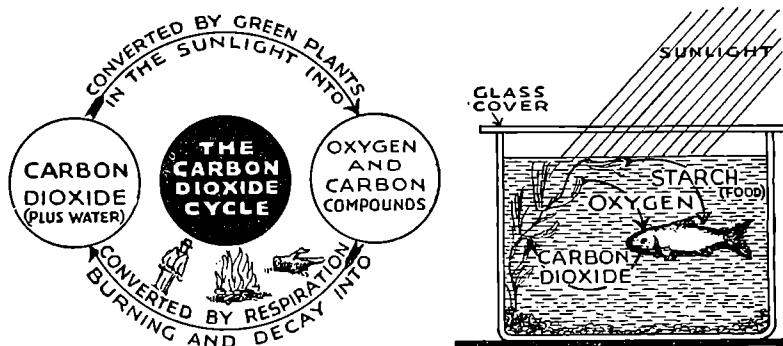


FIG. 34. The balanced aquarium is a completely closed system in which plant and animal exchange substances necessary for the life of each. Without sunlight, however, the cycle would soon stop. Why?

being withdrawn. The processes which are involved in this exchange may be summarized as follows:

CARBON DIOXIDE IS EVOLVED BY:

1. *Respiration.* Animals inhale oxygen, which combines with the carbon of their food. Carbon dioxide, the product of this oxidation, is exhaled. Green plants also oxidize their food substances, but give off carbon dioxide only at night.
2. *Burning.* Most fuels contain carbon, which is oxidized to carbon dioxide.
3. *Decay.* When organic matter decays, most of the carbon in it is converted into carbon dioxide.
4. *Volcanic Action.* Large quantities of carbon dioxide escape into the air from volcanoes and crevices in the earth's crust.

CARBON DIOXIDE IS ABSORBED BY:

1. *Plants.* All green plants, in the sunlight, absorb carbon dioxide from the air. Aided by the catalytic action of *chlorophyll* (the green coloring matter in the leaf), they convert this gas into starch and other plant food, returning a portion of oxygen to the air.
2. *Rocks.* Under the influence of certain microorganisms present in the ocean, large amounts of carbon dioxide are absorbed in the formation of chalk and limestone beds.

Nature maintains a proper balance between these two sets of opposing tendencies, so that the percentage of carbon dioxide in

the air does not change appreciably. The exchange which takes place between plants and animals through the medium of the air is known as the carbon dioxide cycle. This cycle is well illustrated in the balanced aquarium (Fig. 34).

**Water Vapor in the Air.**—Water vapor goes into the air as a result of *evaporation* from the surface of every body of water, *burning* of fuels containing hydrogen, *oxidation* of hydrogen compounds in the bodies of animals, and *decay* of organic matter.

**Humidity.**—The amount of water vapor which may be present in the air varies with the temperature. Warm air is able to hold more moisture than cold air and will give up its excess moisture when it is cooled.

1. *Absolute Humidity* is the amount of moisture in the air per unit volume (*e.g.* four grains of moisture per cubic foot of air).

2. *Relative Humidity* is the amount of water vapor which the air holds compared with what it could hold at the given temperature if it were saturated. For example, at 30°C., one cubic meter of air can hold thirty grams of water, at which point it is said to be *saturated*. The table below indicates the change in relative humidity as the amount of water vapor present in this volume of air is changed:

AMOUNT OF WATER VAPOR IN 1 CUBIC METER OF AIR AT 30°C.	RELATIVE HUMIDITY
30 grams	100%
20 grams	67%
15 grams	50%
0 grams	0%

**Inert Gases of the Air.**—*Helium, neon, argon, krypton* and *xenon* constitute a group of rare gases present in the atmosphere. They are totally inactive chemically, and no compounds are known which contain these elements. An inspection of their atomic structure (page 57) reveals the reason for their extreme inertness. The atoms of these elements are stable because they all have outer rings of electrons which are complete. Since these atoms have no tendency, therefore, either to borrow or to lend electrons, their valence is zero, and they are completely inert.

The essential facts pertaining to the rare gases are presented in the following table:

GAS	OCCURRENCE AND PREPARATION	PROPERTIES AND USES
<b>Helium</b>	In the sun. In the atmosphere (trace). In certain natural gases (about 1%), from which it is now extracted on a commercial scale.	The lightest gas, next to hydrogen. Non-inflammable. Used in balloons and dirigibles.
<b>Neon</b>	In the atmosphere (trace). Prepared by evaporation of liquid air.	Glowes with an orange-red light under the influence of an electric current. Used in advertising signs and aviation beacons.
<b>Argon</b>	In the atmosphere (0.94%). Prepared by evaporation of liquid air.	Used in electric light bulbs to prevent the vaporization or decomposition of the filament.

NOTE.—Krypton and xenon are of little practical importance.

**Liquid Air.**—Air is converted into a liquid by subjecting it to an extremely high pressure and low temperature. It is a pale-blue liquid, almost as heavy as water. When liquid air is allowed to evaporate, the nitrogen escapes first, because of its lower boiling point, while the other gases follow in the order of their boiling temperatures. This is the basis for the commercial preparation of oxygen, nitrogen and the inert gases. Because liquid air has a strong tendency to evaporate, it is stored in special double-walled containers, called Dewar flasks, which resemble the thermos bottle in construction and principle.

## QUESTIONS

1. Why is the atmosphere much denser at sea level than at the top of a high mountain?
2. Since carbon dioxide is the heaviest gas present in the air, why does it not collect as a layer close to the ground?
3. Give three proofs that air is a mixture, rather than a compound.
4. Explain why mineral compounds of nitrogen are comparatively rare, while free nitrogen is abundant.
5. (a) State the conditions under which nitrogen will combine with oxygen and hydrogen. (b) Write the equation for each reaction.
6. Account for the presence of traces of ozone in the atmosphere.
7. Describe how you would remove (a) carbon dioxide and (b) water vapor from a sample of air.
8. Name the inert gases of the atmosphere, and state three properties which are common to all of them.
9. Why does a pitcher of ice water "sweat" when standing in a warm room?
10. When a tea kettle filled with liquid air is placed on a cake of ice, the liquid boils violently. Explain.

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. The amount of ..... in the air varies between wide limits.
2. Nitrogen is an important element in a class of foods known as .....
3. In the balanced aquarium an exchange of gases occurs. The fish gives ..... to the plant, and, in return, receives ..... from it.
4. The most abundant of the totally inert gases present in the air is .....
5. Pure nitrogen is obtained when a mixture of ..... and ..... is heated.
6. The chief commercial source of oxygen, nitrogen, argon and neon is .....



7. Two functions of nitrogen in the air are . . . . .
8. In the nitrogen cycle, nitrogen-fixing bacteria convert . . . . . into . . . . .
9. The total weight of water present in a unit volume of air is called the . . . . .
10. Two conditions required to convert a gas into a liquid are . . . . .

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A                                | B                                    |
|----------------------------------|--------------------------------------|
| ( ) About 14 lb. per square inch | 1. Organic compound of nitrogen      |
| ( ) The nucleus of a raindrop    | 2. To test for moisture in the air   |
| ( ) Protein                      | 3. Limewater                         |
| ( ) Calcium chloride             | 4. Liquid air                        |
| ( ) Aids starch-making in plants | 5. Inert gases in the air            |
| ( ) To test for carbon dioxide   | 6. Chief source of water vapor       |
| ( ) Evaporation                  | 7. Relative humidity is 100%         |
| ( ) Nitrogen + hydrogen          | 8. Atmospheric pressure              |
| ( ) Air is saturated             | 9. Dust particle                     |
| ( ) Valence of zero              | 10. Chlorophyll                      |
|                                  | 11. Weight of water vapor in the air |
|                                  | 12. To form ammonia                  |

## CHAPTER 11

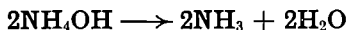
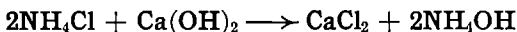
### COMPOUNDS OF NITROGEN

**Importance.**—As a free gas constituting 78% of the atmosphere, nitrogen serves the essential function of diluting the vigorously active oxygen. It is in its compounds, however, that nitrogen demonstrates its supreme usefulness to man. There is an intimate chemical relationship between the complex proteins that form a vital part of our foods, and the ammonia that is employed as a refrigerant to preserve those foods. Nitroglycerine and T.N.T., the explosives, are first cousins to calcium nitrate and ammonium sulfate, the fertilizers, as well as to rayon, celluloid, and a variety of similar substances. Anaesthetics, photographic chemicals, dyes, and drugs are only a few of the host of products derived from nitrogen. The story of nitrogen and its compounds is a story of the triumph of chemical research during the last three decades.

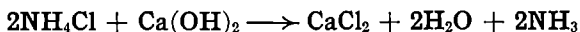
#### AMMONIA (NH<sub>3</sub>)

**Occurrence.**—Ammonia, a gaseous compound of nitrogen and hydrogen, is present in small quantities in the air. It results from the decomposition of animal and vegetable matter, as evidenced by the strong odor of ammonia often noticeable near piles of decaying organic refuse.

**Laboratory Preparation.**—Ammonia is prepared in the laboratory by gently heating a mixture of ammonium chloride and calcium hydroxide. The interaction of these two compounds produces ammonium hydroxide, which is unstable, decomposing into ammonia and water. The equations are:



Combining these two equations we have:



The ammonia is collected by upward displacement of air (Fig 35)

NOTE.—The radical  $\text{NH}_4$  is called the *ammonium radical*. It acts as a metal whose valence is 1, as indicated in the formulas  $\text{NH}_4\text{Cl}$ ,  $\text{NH}_4\text{OH}$ , etc. The ammonium group exists only in combination with an element or with another radical—never by itself. In the electrolysis of its compounds, it is attracted to the cathode, where it decomposes into  $\text{NH}_3$  and  $\text{H}$ .

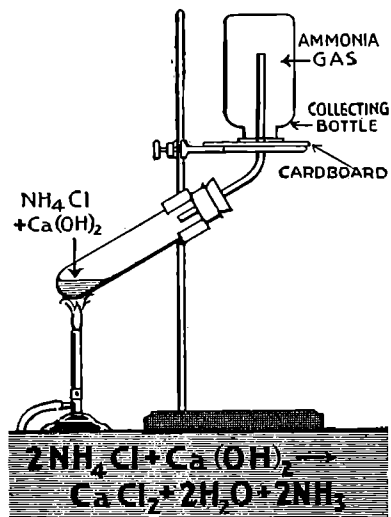
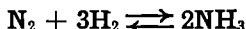


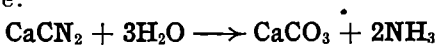
FIG. 35. Laboratory preparation of ammonia.

to combine according to the following equation:



Since this reaction is reversible (page 72), the conditions must be such that a fairly large yield of ammonia can be secured. As high as 40% of the theoretical maximum yield of ammonia is now obtained under the following conditions: (a) the use of iron as a catalyst; (b) a temperature of  $500^\circ\text{C}$ .; (c) a pressure of 200 atmospheres (3000 lb. per square inch); (d) rapid removal of the ammonia formed.

3. *By the Cyanamide Process.*—(See *Nitrogen Fixation*, page 123.) Calcium cyanamide ( $\text{CaCN}_2$ ) is made to react with steam under pressure:



### Commercial Preparation.—

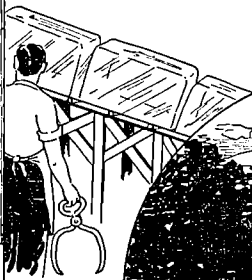
Vast quantities of ammonia are required for a variety of industrial uses. It is manufactured chiefly in three ways:

1. *From Soft Coal.*—When soft coal is heated out of contact with air, ammonia is one of the gaseous products obtained. This process is called *destructive distillation* (page 146). The ammonia is separated from the other products by being dissolved in water.

2. *By Direct Union of Its Elements (Haber Process).*—Nitrogen obtained from liquid air, and hydrogen obtained from the electrolysis of water or from water gas, are made

# NITROGEN COMPOUNDS & THEIR USES

AMMONIA IN ICE-MAKING



NITROUS OXIDE IN DENTISTRY

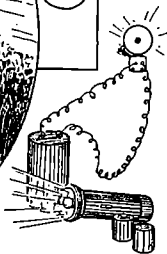


NITRATES USED AS FERTILIZERS

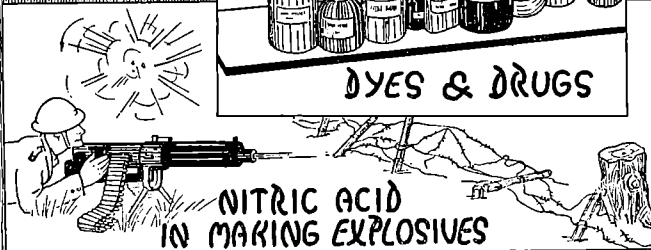


NITROGEN COMPOUNDS (PROTEINS) IN FOODS

AMMONIUM CHLORIDE IN DRY CELLS



DYES & DRUGS

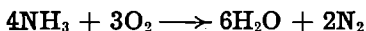


NITRIC ACID IN MAKING EXPLOSIVES

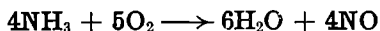
FIG. 36.

**Physical Properties.**—Ammonia is a colorless gas having a strong, pungent odor. It is only about one-half as heavy as air. It is easily liquefied, and has a high heat of vaporization, both of these properties making the gas useful as a refrigerant. One of the outstanding characteristics of ammonia is its extreme solubility in water. Under ordinary conditions of temperature and pressure, one volume of water can dissolve 700 volumes of ammonia. This property may be strikingly demonstrated by the “ammonia fountain,” an apparatus similar to the “hydrogen chloride fountain” described on page 93.

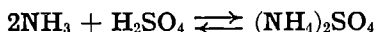
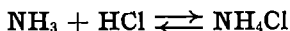
**Chemical Properties.**—Pure, *dry* ammonia is inactive. It burns in pure oxygen:



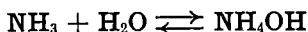
When heated with air in contact with platinum it is oxidized:



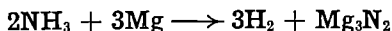
Ammonia does not support combustion. It reacts directly with acids, forming ammonium salts:



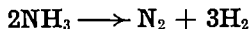
Ammonia unites with water, forming ammonium hydroxide, a weak and unstable base:



Ammonia reacts with active metals at high temperatures:



Ammonia is decomposed by strong heat:



## Uses

1. *Ice-Making and Refrigeration.*—Fig. 37 illustrates the use of ammonia as a refrigerant. The gas is first compressed by means of a pump. The heat generated by this compression is removed by a spray of cold water, which brings about the liquefaction of the gas. The liquid ammonia then passes through an expansion valve where the pressure is suddenly decreased, causing a sharp drop in temperature. The cooled ammonia gas travels through

coils of pipe, which are immersed in brine and surround cans of pure water. The brine is thus cooled below the freezing point of the water, which is converted into cakes of ice. The expanded ammonia returns to the compression pump, where it is recompressed and used over and over. For refrigeration, the cold brine is pumped through coils of pipes in the storage rooms.

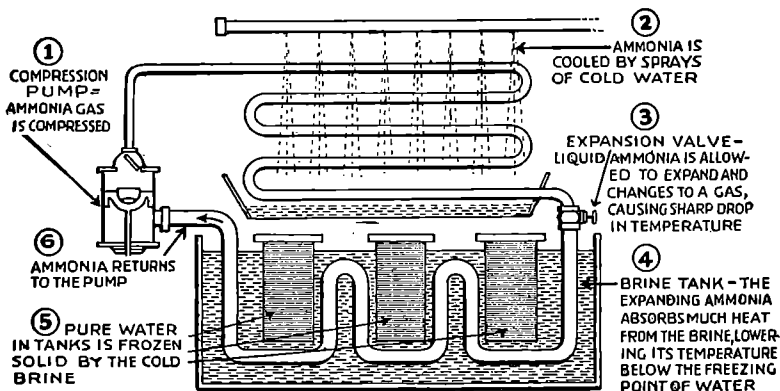


FIG. 37. The construction of a modern ice plant.

2. *Making Ammonia Water.*—Ammonia dissolves in water, thus forming common “household ammonia.” It is used to dissolve grease and remove dirt.

3. *Making Ammonium Salts.*—Ammonia is the starting point for making many ammonium salts. Some of these are ammonium chloride (*sal ammoniac*,  $\text{NH}_4\text{Cl}$ ), used in making dry cells; ammonium sulfate,  $(\text{NH}_4)_2\text{SO}_4$ , used as a fertilizer; and ammonium nitrate,  $\text{NH}_4\text{NO}_3$ , used to make explosives.

4. *Making Sodium Carbonate.*—Large quantities of this important chemical are made by the Solvay Process (page 204).

5. *Making Nitric Acid.*—Ammonia is oxidized to produce nitric oxide, which is converted into nitric acid (page 124). This is the chief commercial use of ammonia.

**Test for the Ammonium Radical.**—Heat the salt to be tested with a base, such as  $\text{NaOH}$ . If an odor of ammonia is produced and the escaping gas turns moist red litmus paper blue, the presence of an ammonium salt is indicated:

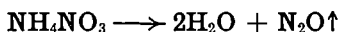


## OXIDES OF NITROGEN

Nitrogen combines with oxygen to form five distinct compounds, listed in the table below. This series of oxides illustrates (1) the variable valence of nitrogen, and (2) the Law of Multiple Proportions.

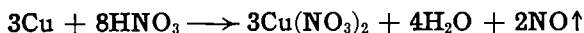
VALENCE OF NITROGEN	OXIDE	FORMULA
1	Nitrous oxide	$N_2O$
2	Nitric oxide	$NO(N_2O_2)$
3	Nitrogen trioxide	$N_2O_3$
4	Nitrogen peroxide	$NO_2(N_2O_4)$
5	Nitrogen pentoxide	$N_2O_5$

**Nitrous Oxide.**—Nitrous oxide ( $N_2O$ ) is prepared by carefully heating ammonium nitrate, and is collected by the displacement of water. The equation is:



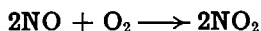
Nitrous oxide is a colorless gas having a pleasant odor and a slightly sweet taste. It is somewhat heavier than air, and is sparingly soluble in water. The gas does not burn, but it supports combustion. A glowing splint bursts into flame when placed in nitrous oxide, as it does in oxygen. These two gases may be distinguished from each other, however, because barely ignited sulfur will continue to burn in oxygen, but will not do so in nitrous oxide. Mixed with about 25% of oxygen, nitrous oxide is used as an anaesthetic in dentistry and for minor surgical operations. It is commonly known as "laughing gas."

**Nitric Oxide.**—Nitric oxide ( $NO$ ) is prepared in the laboratory by the action of dilute nitric acid on copper. The gas is collected by the displacement of water. The equation is:

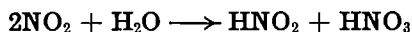


Nitric oxide is a colorless gas, slightly heavier than air, insoluble in water, and poisonous. It neither burns nor supports combustion. On contact with the air, it combines with oxygen, forming brown nitrogen peroxide.

**Nitrogen Peroxide.**—Nitrogen peroxide ( $\text{NO}_2$ ) is prepared by exposing nitric oxide to the air:

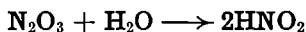


This substance is a brown gas with a disagreeable odor, and is poisonous. It is much heavier than air and is soluble in water, reacting with it to form a mixture of nitrous acid and nitric acid:



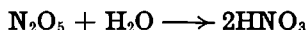
Nitrogen peroxide, along with nitric oxide, is used in the manufacture of nitric and sulfuric acids.

**Nitrogen Trioxide.**—Nitrogen trioxide ( $\text{N}_2\text{O}_3$ ) is an unstable reddish-brown gas. It unites with water, forming nitrous acid, and is therefore called *nitrous anhydride*:



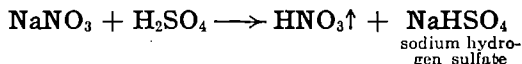
Nitrous acid is used in the manufacture of dyes. It is so unstable that it must be made just before it is needed.

**Nitrogen Pentoxide.**—Nitrogen pentoxide ( $\text{N}_2\text{O}_5$ ) is a white solid, and is also very unstable. It unites with water, forming nitric acid, and is therefore called *nitric anhydride*:

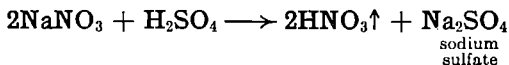


### NITRIC ACID ( $\text{HNO}_3$ )

**Preparation.**—Nitric acid, the most common acid of nitrogen, is prepared both commercially and in the laboratory by the action of sulfuric acid on a nitrate. In the laboratory preparation, a small quantity of sodium nitrate is mixed with concentrated sulfuric acid in a glass retort, and heated gently (Fig. 38). Nitric acid is formed, and is driven off as a gas, since its boiling point is much lower than that of sulfuric acid:



If a higher temperature is used, the reaction is:





**Physical Properties.**—Pure nitric acid is a colorless liquid, boiling at  $86^{\circ}\text{C}$ . It is miscible with water in all proportions. Commercial nitric acid is a water solution containing 68% of pure acid. It is often colored yellowish-brown because of dissolved nitrogen peroxide. Nitric acid is highly corrosive to the skin.

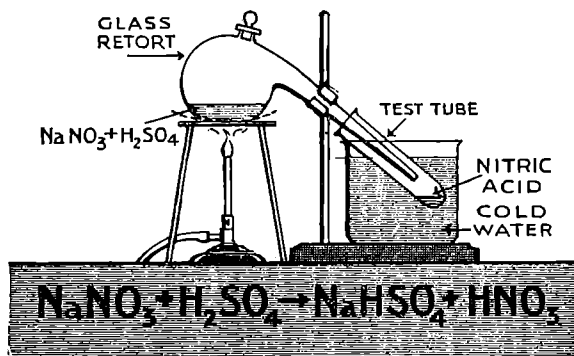
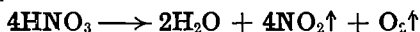


FIG. 38. Laboratory preparation of nitric acid.

### Chemical Properties

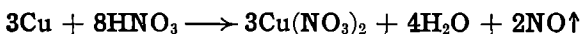
1. *Stability.*—In the presence of sunlight, or when heated, nitric acid is decomposed, yielding free oxygen, which may be identified by means of a glowing splint. The equation for the decomposition is:



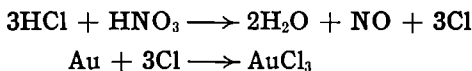
2. *Acid Properties.*—Nitric acid is a strong acid, being highly ionized, even with a small amount of water.

3. *Oxidizing Power.*—The instability of nitric acid, and the high percentage of oxygen which it contains, account for its strong oxidizing action. This action may be shown in several ways: (a) Wool is set on fire when placed in hot nitric acid vapors. (b) Hot charcoal is oxidized to carbon dioxide by concentrated nitric acid.

4. *Action with Metals.*—Nitric acid reacts with metals, forming hydrogen as the initial product. However, the hydrogen is immediately oxidized by the acid, forming water. The other compounds formed in the reaction depend on the particular metal being used, and on the concentration of the nitric acid. The complete equation for the action of cold dilute acid on copper is:



5. *Action in Aqua Regia.*—Aqua regia is a mixture of three parts of hydrochloric acid and one part of nitric acid. The nitric acid oxidizes the hydrogen of the hydrochloric acid, yielding nascent chlorine. The chlorine thus formed reacts with gold and platinum, yielding chlorides:



These metals, not affected by any one common acid, are thus dissolved by aqua regia.

**Uses.**—Nitric acid is a highly important chemical in many industries, about 635,000 tons being produced yearly in this country. Its chief use is in the making of explosives. Among the principal products of this type are nitroglycerine, dynamite, guncotton and trinitrotoluene (T.N.T.). Nitric acid is used to make fertilizers, artificial silk, drugs, dyes, and celluloid. It is essential in the production of nitrates, such as silver nitrate, which is used in the manufacture of photographic film. Another important use of nitric acid is to separate gold from silver, since it acts on and dissolves silver, but has no effect on gold. Large amounts of the acid are used to make aqua regia.

**Nitrates.**—The salts of nitric acid are called *nitrates*. All common nitrates are soluble in water. They are decomposed by heat, usually forming a metallic oxide, nitrogen peroxide, and oxygen:



Two important nitrates are sodium nitrate (*Chile saltpeter*,  $\text{NaNO}_3$ ), used as a fertilizer, and for making nitric acid and explosives; and potassium nitrate (*saltpeter*,  $\text{KNO}_3$ ), used as a preservative for meat, and in the manufacture of black gunpowder and fireworks.

**Test for a Nitrate.**—To identify a nitrate, add to a solution of the substance being tested a freshly prepared solution of ferrous sulfate. Then add carefully some concentrated sulfuric acid, allowing the acid to run slowly down the side of the test tube and collect at the bottom. The formation of a *brown ring* where the liquids meet indicates the presence of nitric acid or a nitrate.

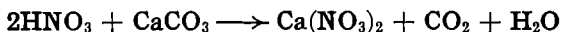
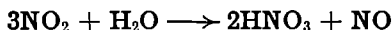
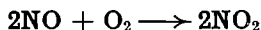
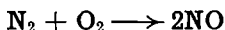
## FIXATION OF NITROGEN

**Importance.**—The fixation of nitrogen is the process of converting the free nitrogen of the air into useful compounds such as nitric acid, nitrates and ammonia. Nitrogen is an important element for the growth of plants, yet it cannot be used in its free atmospheric state by the vast majority of them. Therefore, fertilizers, chiefly nitrates, must be added regularly to soil used for farming. In addition, immense quantities of nitrogen compounds are required for the manufacture of explosives, dyes, drugs, and a variety of other products. However, the supply of natural nitrates, such as Chile saltpeter, is limited. Hence, the vast ocean of air has become man's chief reservoir for the nitrogen compounds needed both in peace and in war.

**Nitrogen Fixation by Nature.**—Certain natural processes are constantly taking place which convert some of the atmospheric nitrogen into compounds. These are:

1. *Action of Bacteria.*—Nitrogen-fixing bacteria are found in nodules on the roots of the leguminous plants such as beans, peas, clover, and alfalfa. These bacteria are able to change nitrogen directly to nitrates, which serve as food for plants (see *Nitrogen Cycle*, page 108).

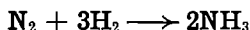
2. *Action of Lightning.*—During a thunderstorm, small amounts of nitrogen and oxygen in the air are made to combine, forming nitric oxide, which is immediately oxidized to nitrogen peroxide. This, in turn, dissolves in rain water, forming nitric acid. On reaching the ground, the acid combines with mineral substances, forming nitrates. The reactions involved are:



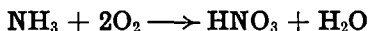
**Nitrogen Fixation by Man.**—Natural fixation processes are slow and uncertain; hence chemists have been energetically at work in recent years developing methods which will insure an adequate supply of nitrogen compounds. At present more than half the world's requirements of these compounds are filled by artificial fixation processes. These are:

1. *The Electric Arc Process.*—Air is passed through a huge electric arc, which acts on the nitrogen and oxygen in a manner similar to that of lightning, described above. The nitric oxide formed is converted into nitrogen peroxide, which combines with water in an absorption tower, forming dilute nitric acid. This is converted by limestone into calcium nitrate, which is used as a fertilizer. The reactions involved are the same as given above. This process is carried out successfully in Norway, where cheap hydro-electric power is available.

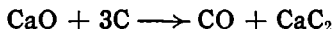
2. *The Synthetic Ammonia Process.*—By this method (the most successful one for nitrogen fixation) nitrogen and hydrogen are combined under special conditions to form ammonia (see *Haber Process*, page 115).



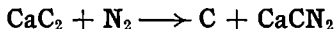
The ammonia is oxidized to nitric oxide, platinum being used as a catalyst. The nitric oxide is next changed to nitrogen peroxide, which is dissolved in water, forming nitric acid. The following equation summarizes these reactions:



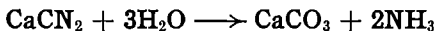
3. *The Cyanamide Process.*—Calcium carbide ( $\text{CaC}_2$ ) is made by heating lime ( $\text{CaO}$ ) and coke in an electric furnace:



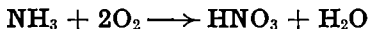
Calcium cyanamide ( $\text{CaCN}_2$ ) is then made by passing atmospheric nitrogen over white-hot calcium carbide:



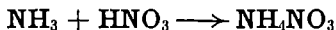
Steam under pressure is then passed through the hot calcium cyanamide, resulting in the formation of ammonia:



The ammonia is then oxidized to nitric acid, as described in the synthetic ammonia process above:



The nitric acid may be combined with ammonia to form ammonium nitrate, which enters into the manufacture of explosives:



NOTE.—The government plant at Muscle Shoals, Alabama, employs the cyanamide process for the fixation of nitrogen.

## QUESTIONS

1. Explain the difference between *ammonia* and *ammonium*.
2. What evidence is there that protoplasm is a nitrogenous compound?
3. Describe an experiment to demonstrate the extreme solubility of ammonia.
4. Compare the action of ammonia gas on red litmus paper when the paper is (a) dry, (b) wet. Explain.
5. Describe two chemical tests to prove that a given substance is ammonium nitrate.
6. (a) Why is sodium nitrate, rather than potassium nitrate, used in the commercial preparation of nitric acid? (b) Write the equation for the reaction.
7. Given six bottles, each containing one of the following gases: (a) oxygen, (b) nitrogen, (c) ammonia, (d) nitrous oxide, (e) nitric oxide, and (f) nitrogen peroxide. Explain how you would identify each gas.
8. (a) Name and state the law which is illustrated by the five oxides of nitrogen. (b) Explain how the law applies to these compounds.
9. Give two illustrations to show that nitric acid is a good oxidizing agent.
10. (a) Why is hydrogen not liberated in the reaction between copper and nitric acid? (b) Write the equation for this reaction.
11. Account for the abnormally great demand for nitric acid during the World War.
12. Write the equations for the reactions involved in the fixation of nitrogen by (a) the arc method; (b) the cyanamide method.

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Ammonia is prepared in the laboratory by heating a mixture of ..... and .....
2. Ammonia is obtained on a commercial scale as a by-product in the ..... of .....
3. Household ammonia, known chemically as ....., is used as a cleaning agent because of its ability to .....

4. Commercial nitric acid is colored . . . . . because of the presence of dissolved . . . . .
5. A mixture of nitric acid and hydrochloric acid is called . . . . . It is able to dissolve gold because of the action of . . . . . generated by the two acids.
6. Nitrous oxide is frequently called . . . . . It is employed as . . . . .
7. The arc process for making . . . . . is employed mostly in . . . . ., where . . . . . is cheap.
8. The most abundant mineral nitrate is . . . . ., commonly called . . . . .
9. Most of the uses of nitric acid depend on the fact that it is a powerful . . . . .
10. The anhydride of nitric acid is . . . . .

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

#### A

- ( ) Nitrogen + hydrogen
- ( ) Platinum as a catalyst
- ( ) Liquid ammonia
- ( ) Nitrogen trioxide
- ( ) Supports combustion
- ( ) Nitrogen fixation at Muscle Shoals
- ( ) A use of nitric acid
- ( ) To test for nitrates
- ( ) Produces nitric oxide
- ( ) Nitrogen-fixing bacteria

#### B

1. Chile saltpeter
2. Refrigerant
3. Anhydride of nitrous acid
4. To separate gold from silver
5. Fresh ferrous sulfate
6. Commercial oxidation of  $\text{NH}_3$
7. Haber Process
8. Leguminous plants
9. Lightning
10. Gold
11. Calcium cyanamide process
12. Nitrous oxide

## CHAPTER 12

### SULFUR AND SULFUR COMPOUNDS

#### SULFUR

**Importance.**—Sulfur is an element of great importance in industry. About  $1\frac{1}{2}$  million tons of it are consumed each year in the United States alone. A considerable portion of this veritable mountain of sulfur goes into the manufacture of sulfuric acid, a chemical which plays a prominent part in the preparation of a host of other materials. Sulfur dioxide, another important compound of sulfur, is used for disinfecting, bleaching, and refrigeration. In addition, sulfur is employed in the process of vulcanizing rubber, and as an ingredient of chemical sprays for trees and shrubs. It is also used for making carbon disulfide, matches, dyes, fireworks, and explosives.

**Occurrence.**—Sulfur occurs widely distributed in both the free and the combined states.

1. *Free or native* sulfur is found extensively in Texas, Louisiana, and Sicily.

2. *Combined* sulfur occurs as sulfides, such as galena ( $\text{PbS}$ ); as sulfates, such as gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ) and barite ( $\text{BaSO}_4$ ); as sulfur dioxide ( $\text{SO}_2$ ) and hydrogen sulfide ( $\text{H}_2\text{S}$ ) in volcanic gases; and in animal and vegetable substances, such as egg yolk and mustard.

**Frasch Method of Extraction.**—In Texas and Louisiana sulfur is found about 500 feet below the surface, under layers of shifting quicksands and clay, which make it necessary to employ a special mining technique (Fig. 39). Three concentric pipes (1, 3, and 6 inches in diameter, respectively) are sunk to the sulfur level, and superheated water at a temperature of  $170^\circ\text{C}$ . is forced down through the outermost pipe, causing the sulfur to melt. Compressed air is then pumped down through the innermost pipe, mixing with the sulfur to make it foamy and light. This mixture of air, water, and melted sulfur is forced up the middle pipe in a frothy

stream. The molten sulfur is then run into large wooden bins, where it separates from the air and water, and turns solid. This sulfur is 99½% pure, and, for most industrial uses, requires no further treatment.

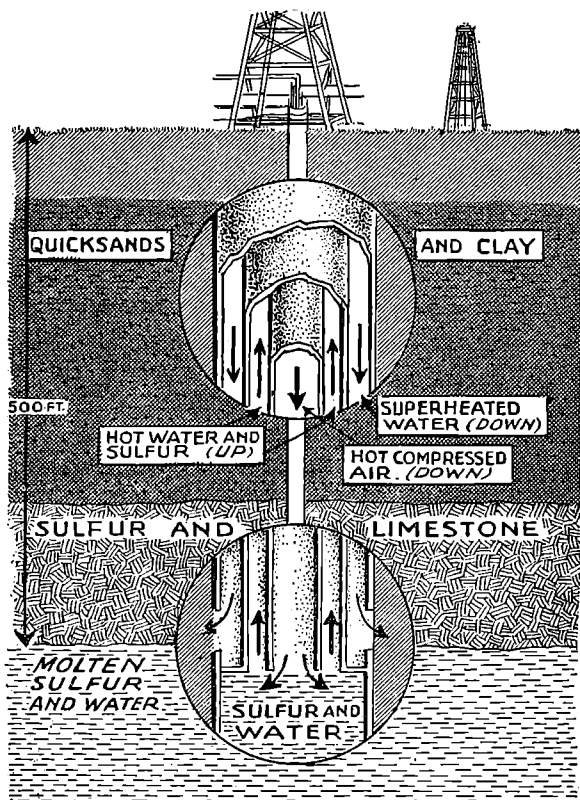


FIG. 39. Frasch method of extracting sulfur.

**Allotropic Forms.**—Sulfur occurs in three allotropic forms (Fig. 40): *rhombic*, *prismatic* (or *monoclinic*), and *amorphous* (or *plastic*). Because of differences in atomic arrangement and energy content, these allotropic forms exhibit widely varying properties, as shown in the table on the following page.



## COMPARISON OF PHYSICAL PROPERTIES

RHOMBIC	PRISMATIC	AMORPHOUS
Yellow Double-pyramid crystals Specific gravity, 2.06 Soluble in carbon disulfide Stable below 96°C.	Transparent Needle-like crystals  Specific gravity, 1.96 Soluble in carbon disulfide Stable between 96° and 114°C. Changes to rhombic below 96°C.	Amber Plastic, non-crystalline  Specific gravity, 1.95 Insoluble in carbon disulfide Stable above 114°C.  Changes to rhombic when cooled

## Preparation of Allotropic Forms

1. *Rhombic*.—Some sulfur is dissolved in a little carbon disulfide, and the solution is poured into a shallow dish and allowed to stand. The solvent evaporates slowly, leaving a residue consisting of perfect crystals of rhombic sulfur.

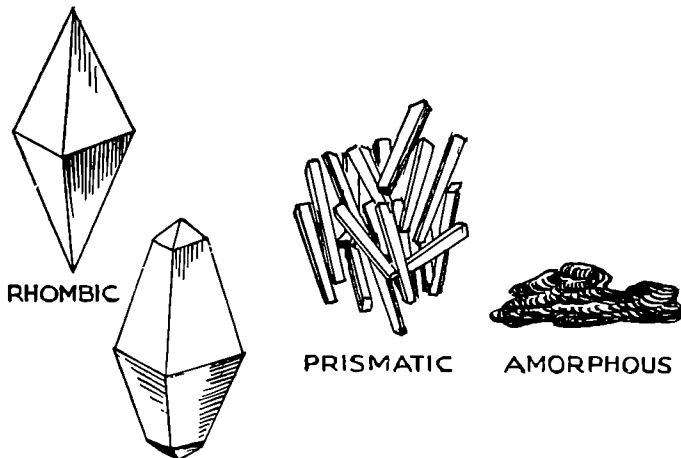


FIG. 40. The three allotropic forms of sulfur.

2. *Prismatic*.—Some sulfur is carefully heated in a test tube until it barely melts. The liquid is then poured into a dish and allowed to cool slowly. When the surface begins to harden, the

# SULFUR & its USES

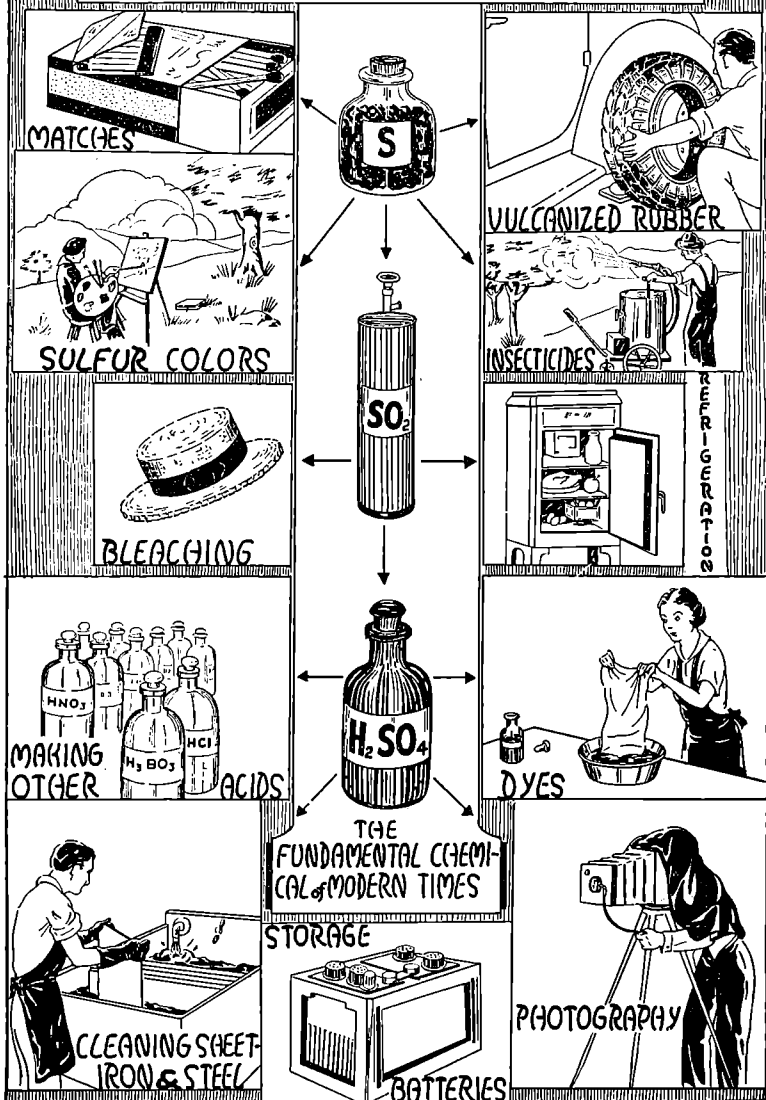


FIG. 41.

sulfur which is still liquid is poured off. The solidified portion consists of needle-like crystals of prismatic sulfur.

3. *Amorphous*.—Some sulfur is heated in a test tube until it boils. The liquid is then cooled suddenly by pouring it into cold water. A rubbery, non-crystalline product is obtained.

### Properties

1. *Physical*.—Ordinary sulfur is a pale-yellow, brittle, crystalline solid, about twice as heavy as water. It has no taste and very little odor. It is insoluble in most acids, but dissolves readily in carbon disulfide. Sulfur melts at  $114.5^{\circ}\text{C}$ ., forming a thin, straw-colored liquid. If the heating is continued, the liquid becomes darker and thicker, until a jelly-like, viscous mass is formed at about  $230^{\circ}\text{C}$ . Above this temperature, the liquid again becomes thin, although remaining dark in color. At  $445^{\circ}\text{C}$ ., it boils and produces a yellow vapor.

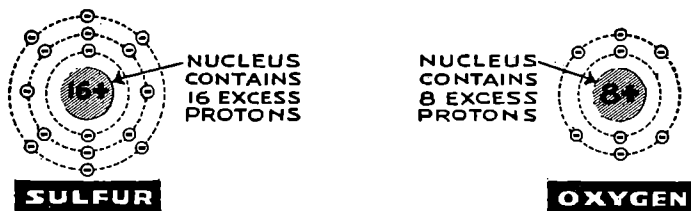
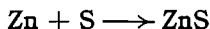
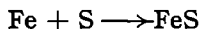
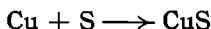
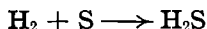


FIG. 42. Comparison of the atoms of sulfur and oxygen. Both are non-metals with a valence of  $-2$ .

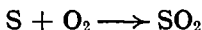
2. *Chemical*.—Sulfur has an atomic structure which is much like that of oxygen (Fig. 42), and, in its chemical properties, resembles that element to a marked degree. Thus, sulfur combines with most metals, forming sulfides which correspond to the oxides of the metals:



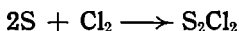
Sulfur vapor supports combustion like oxygen:



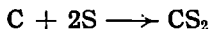
Sulfur exhibits *amphoteric* properties by combining not only with metals as above, but with non-metals like oxygen, chlorine, and carbon. Sulfur burns in air with a pale-blue flame.



Sulfur combines with chlorine, forming sulfur chloride, which is used in vulcanizing rubber:



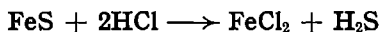
Sulfur vapor reacts with highly heated carbon, forming carbon disulfide:



### HYDROGEN SULFIDE ( $\text{H}_2\text{S}$ )

**Occurrence.**—Hydrogen sulfide is found in the gases escaping from volcanoes. It occurs, also, in the water of "sulfur" springs. It is given off by decaying organic matter, which contains sulfur, as is evidenced by the disagreeable odor of rotten eggs.

**Preparation.**—Hydrogen sulfide is usually prepared in the laboratory by the action of dilute hydrochloric or sulfuric acid on ferrous sulfide:

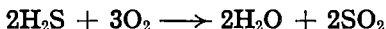


The gas is collected by displacement of air.

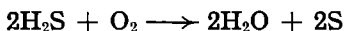
#### Properties

1. *Physical.*—Hydrogen sulfide is a colorless gas having the extremely offensive odor of rotten eggs. It is heavier than air and slightly soluble in water. The gas is poisonous, and even when a small quantity of it is inhaled it produces headache and nausea.

2. *Chemical.*—Hydrogen sulfide does not support combustion, but burns readily. If there is a sufficient supply of air, the combustion is complete, water and sulfur dioxide being formed:

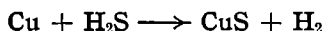
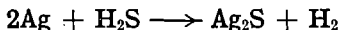


If the air supply is limited, as when the gas burns in a bottle or in a test tube, the products are water and free sulfur:

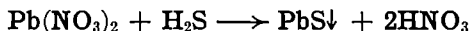


Hydrogen sulfide is a good reducing agent. It dissolves sparingly in water, forming the weak hydrosulfuric acid.

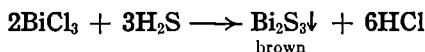
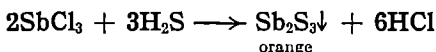
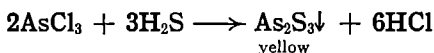
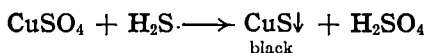
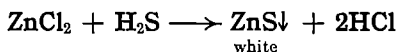
Hydrogen sulfide reacts with many metals, forming sulfides:



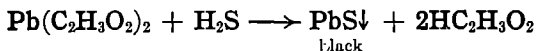
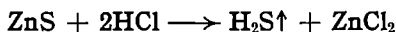
These sulfides are the principal substances present in the black tarnish which frequently forms on the surface of silver and copper articles. Hydrogen sulfide also reacts with many metallic salts in solution, forming insoluble sulfides:



**Uses.**—The principal use of hydrogen sulfide is in the precipitation of metallic sulfides from solutions of salts. When the gas is passed through such solutions, insoluble sulfides of characteristic color are precipitated, which may be employed in the laboratory to identify unknown metal ions.



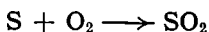
**Test for a Sulfide.**—To identify a substance as a sulfide, add dilute hydrochloric acid, and hold near the mouth of the tube a strip of paper that has been dipped in lead acetate solution. The escaping gas ( $\text{H}_2\text{S}$ ) is recognized by its odor and by the fact that it turns the paper black, owing to the formation of black lead sulfide.



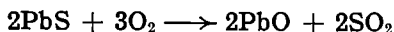
### SULFUR DIOXIDE ( $\text{SO}_2$ )

**Occurrence.**—Sulfur dioxide is found to a limited extent in volcanic gases, and in the vapors from certain mineral springs. It is evolved also by burning coal, which usually contains small quantities of sulfur, either free or in the combined form.

**Commercial Preparation.**—Sulfur dioxide is prepared commercially by (1) burning sulfur:



and (2) heating a metallic sulfide in air (roasting):



**Laboratory Preparation.**—Sulfur dioxide may be prepared on a small scale by the action of an acid on either a sulfite, a bisulfite, or

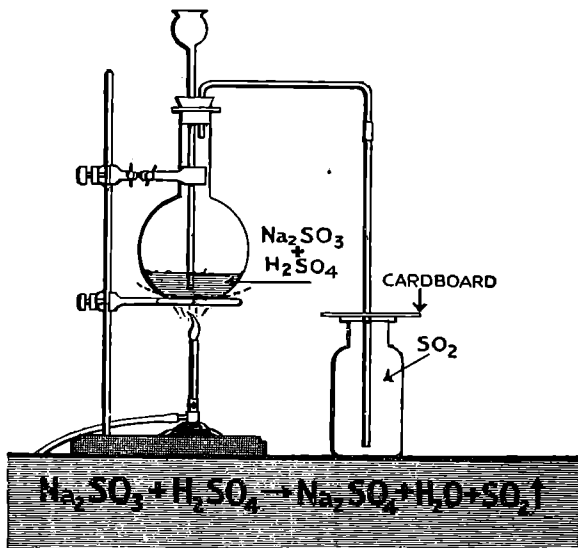
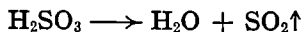
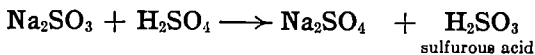
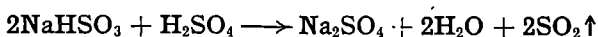


FIG. 43. Laboratory preparation of sulfur dioxide.

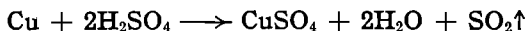
a metal like copper. In the first, and most common, method, dilute sulfuric acid is slowly added to some sodium sulfite in a flask, and gentle heat is applied to the mixture (Fig. 43):



If a bisulfite is used, the final reaction is:



If copper is used, the metal, in the form of turnings, is heated strongly with concentrated sulfuric acid:

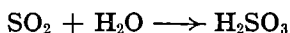


In all cases, the liberated sulfur dioxide is collected by the downward displacement of air.

### Properties

1. *Physical*.—Sulfur dioxide is a colorless gas having a strong, suffocating odor. It is more than twice as heavy as air. It is extremely soluble in water, and can be very easily liquefied.

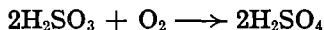
2. *Chemical*.—Sulfur dioxide is a stable compound. It neither burns nor supports combustion. It combines with water to form sulfurous acid, and is therefore called *sulfurous anhydride*.



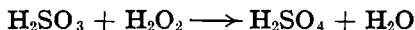
Sulfurous acid is an unstable compound, as shown by its strong odor of sulfur dioxide. Like all acids it turns litmus red, has a sour taste, and reacts with bases, forming sulfites:



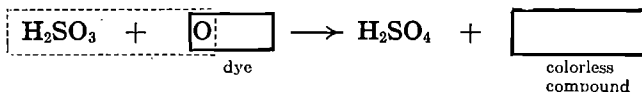
Sulfurous acid slowly takes up oxygen when in contact with the air, forming sulfuric acid:



In a similar manner, it is able to remove oxygen from certain compounds, thus acting as a reducing agent. The action with hydrogen peroxide is as follows:



Sulfurous acid also acts as a bleaching agent. In some cases, the reducing action of the acid results in the removal of color:



In other cases, the sulfur dioxide combines with the coloring matter, forming colorless compounds.

**Uses of Sulfur Dioxide**

1. *For Bleaching.*—Moist sulfur dioxide is used to bleach straw, wood pulp (for making paper), wool, and silk. Dried and canned fruits, such as apricots and cherries, are sometimes bleached with sulfur dioxide.

2. *For Preserving Foods.*—Sulfur dioxide kills bacteria, and thus tends to prevent fermentation. Since the gas may be harmful to human beings, this use is of questionable practicability.

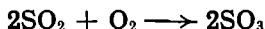
3. *For Disinfecting.*—Sulfur candles are sometimes used to disinfect rooms. Since the sulfur dioxide may corrode metal fixtures and bleach wall paper, its use for this purpose is limited.

4. *For Refrigeration.*—Sulfur dioxide has certain definite advantages over ammonia as a refrigerant and is coming into wider use as a substitute for this gas.

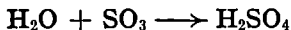
5. *For Making Sulfuric Acid.*—This is by far the most important use of sulfur dioxide.

**SULFUR TRIOXIDE (SO<sub>3</sub>)**

**Preparation and Properties.**—Sulfur trioxide is made by passing a mixture of sulfur dioxide and air over heated, finely divided platinum, which acts as a catalyst:



Sulfur trioxide is a white solid, melting at room temperature to a colorless liquid which fumes in moist air. Sulfur trioxide dissolves in water to form sulfuric acid, and is therefore an acid anhydride:



It dissolves readily in concentrated sulfuric acid to form *fuming sulfuric acid*:



Sulfur trioxide finds its chief use in the manufacture of sulfuric acid by the contact process.

**SULFURIC ACID (H<sub>2</sub>SO<sub>4</sub>)**

**Preparation.**—Sulfuric acid is made by the *lead chamber process* and, more recently, by the *contact process*. In both processes, sulfur dioxide is first made by burning sulfur or by roasting a metallic



sulfide. In the older process, the change to sulfur trioxide is brought about through the catalytic action of oxides of nitrogen in large lead chambers, hence the name "lead chamber" process. The contact process employs platinum as the catalyst to effect the oxidation. "Contact" acid is highly concentrated (about 98%), while "chamber" acid is relatively dilute (60 to 70%).

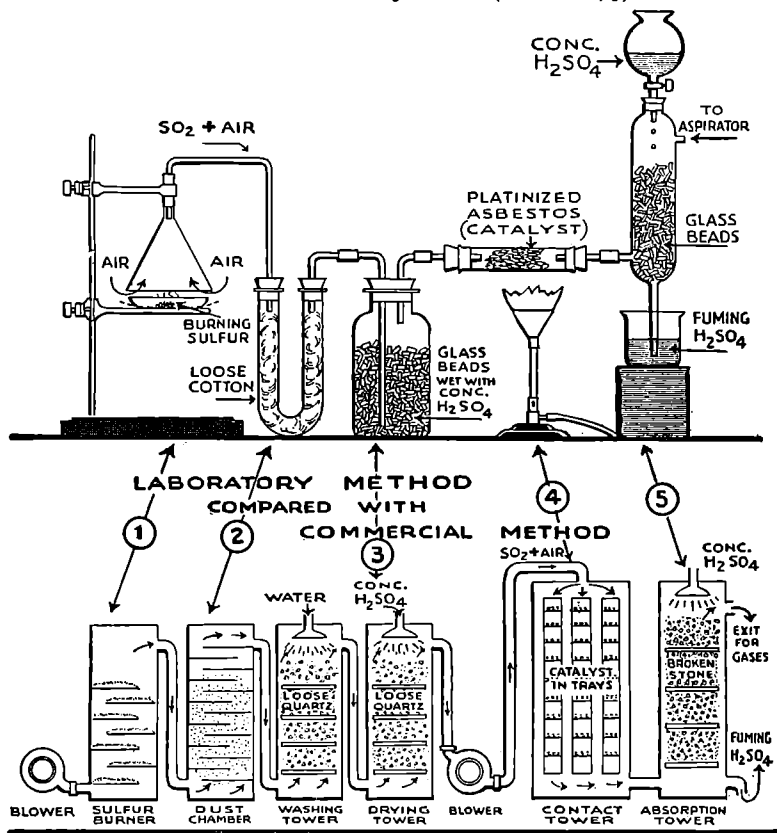


FIG. 44. The contact process for making sulfuric acid.

**Contact Process.**—Fig. 44 illustrates the contact process for manufacturing sulfuric acid as carried out in the laboratory and on a commercial scale. The numbers 1, 2, 3, etc. indicate related apparatus and steps in both methods. These steps are as follows:

1. Sulfur dioxide is produced by burning sulfur or some ore of sulfur, and is mixed with an excess of air.

2. The mixture of gases is passed through the dust chamber to remove sulfur dust and other impurities.

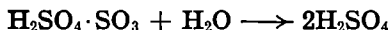
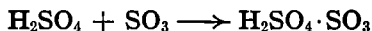
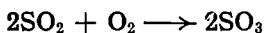
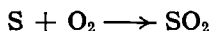
3. The gases are further purified by washing with water, and are then dried by contact with concentrated sulfuric acid.

4. The gases are then passed over the catalyst (platinized asbestos, iron oxide, or certain complex vanadium compounds) at a temperature of about 400°C., thus forming sulfur trioxide.

5. The sulfur trioxide is absorbed by concentrated sulfuric acid, producing fuming sulfuric acid.

6. The fuming sulfuric acid is then diluted with water, forming concentrated sulfuric acid.

The reactions which take place in the contact process may be represented as follows:

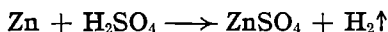


**Physical Properties.**—Concentrated sulfuric acid is a syrupy, corrosive liquid called *oil of vitriol*. It has a specific gravity of 1.8 and a boiling point of 338°C., and is thus the heaviest and the least volatile of all ordinary acids.

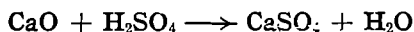
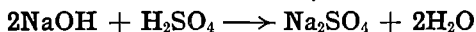
Much heat is evolved when sulfuric acid is mixed with water. Therefore, in making a dilute solution of the acid, care must be taken to avoid dangerous spattering that would result if steam were produced. This is done by pouring the acid slowly into the water (never the reverse), accompanied by constant stirring.

**Chemical Properties.**—The behavior of sulfuric acid varies, depending on whether it is dilute or concentrated.

1. *Dilute* sulfuric acid has the usual properties of a strong acid. It reacts with metals, liberating hydrogen:



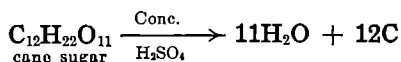
It neutralizes bases and metallic oxides (basic anhydrides), yielding salts and water:



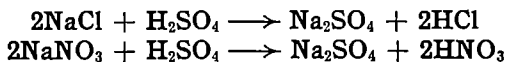
2. *Concentrated* sulfuric acid exhibits certain special oxidizing and dehydrating properties. When hot, it acts with metals as an oxidizing agent, usually liberating sulfur dioxide:



Concentrated sulfuric acid has so strong an affinity for water that it often abstracts hydrogen and oxygen (in the proportion of 2 to 1, as in water) from certain compounds called *carbohydrates*. Thus wood, paper, and sugar are charred by concentrated sulfuric acid because of the removal of hydrogen and oxygen, leaving a black residue which is principally carbon. This action is called *dehydration*:



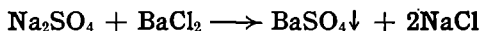
Concentrated sulfuric acid reacts with many salts, forming volatile acids. This is the basis for the preparation of most of the common acids:



**Uses.**—Sulfuric acid has been called the “king of chemicals” since, directly or indirectly, it is used in practically every important industry. The following table lists a few of its outstanding applications, and shows how these uses are related to the properties of sulfuric acid.

PROPERTIES	USES DEPENDING ON THESE PROPERTIES
Stability and high boiling point	Manufacture of other acids, such as hydrochloric acid and nitric acid. Converting insoluble calcium phosphates ( <i>phosphate rock</i> ) into soluble phosphates for fertilizers.
Great affinity for water	Dehydrating agent in making nitrocellulose, nitroglycerine, and various dyes and drugs. Drying gases, both in the laboratory and commercially.
Oxidizing power	Refining petroleum, by removing substances which would give the product a dark color.
Electrical conductivity in dilute solutions	Making storage batteries, in electroplating, and refining of copper, silver, and gold.
Acid properties	“Pickling,” or removing oxide impurities from the surface of metals before they are tinned or galvanized.

**Test for the Sulfate Ion.**—Most sulfates are soluble in water. Barium sulfate is a white solid which is insoluble in water and in dilute acids. Hence, its formation may serve as a test for the sulfate ion. Barium chloride solution is added to the solution being tested. The formation of a white precipitate, which does not dissolve in hydrochloric acid, indicates the presence of sulfuric acid or a soluble sulfate:



### QUESTIONS

1. In the Frasch process for obtaining sulfur, explain the reason for using (a) superheated water, (b) compressed air.
2. (a) Name three different forms of sulfur. (b) Why are they called allotropic forms?
3. Given a yellow solid. How would you prove that it is sulfur?
4. Explain, according to the Electron Theory, (a) why sulfur usually acts as a non-metal; (b) how sulfur may sometimes act as a metal.
5. Why does a silver spoon tarnish after contact with egg yolk?
6. How would you prove that hydrogen sulfide is produced when organic matter decays?
7. "Zinc oxide" paint retains its whiteness for a much longer time than does "white lead" paint. Explain.
8. Why is sulfur dioxide, rather than chlorine, used in bleaching wool?
9. (a) How is sulfur dioxide collected in the laboratory? (b) Give two reasons why this method is used.
10. (a) Why is sulfur trioxide called sulfuric anhydride? (b) Write the equation to explain this.
11. Sulfuric acid is used in the manufacture of sulfuric acid. Explain the reason for this.
12. Write the equation for the action of hot, concentrated sulfuric acid on carbon.
13. Explain why sulfuric acid is used in the preparation of other acids.
14. The formula for starch is  $\text{C}_6\text{H}_{10}\text{O}_5$ . Explain why starch turns black when concentrated sulfuric acid is poured on it.
15. How would you distinguish between sodium chloride and sodium sulfate?

**COMPLETION TEST**

*Supply the word or words required to complete each of the following statements.*

1. Rhombic sulfur is . . . . . (soluble, insoluble) in carbon disulfide.
2. The commercial variety of sulfur is known as . . . . .
3. The water solution of hydrogen sulfide is a weak . . . . .
4. The process of heating a sulfur ore with plenty of air is called . . . . .
5. Sulfur dioxide is able to bleach only when it is . . . . .
6. When sulfur dioxide is bubbled through a solution of sodium hydroxide, . . . . . is formed.
7. Concentrated sulfuric acid chars sugar because it is a . . . . . agent.
8. Sulfates are salts of . . . . . acid.
9. In the contact process of preparing sulfuric acid, the catalyst is usually . . . . .
10. The product obtained in the contact process is . . . . . sulfuric acid; the product obtained in the chamber process is . . . . . sulfuric acid.

**MULTIPLE CHOICE TEST**

*In each of the following, select the word or words that best completes the statement.*

1. A crystal of prismatic sulfur is (cubic, needle-like, double-pyramid) in form.
2. When sulfur vapor supports combustion, the element acts as a (metal, non-metal, ion).
3. (Hydrogen sulfide, Sulfur dioxide, Flowers of sulfur) is used in analytical chemistry.
4. "White lead" paint turns dark because of the action of (sulfur dioxide, carbon dioxide, hydrogen sulfide) in the air.
5. Sulfur dioxide has a (sweet, nauseating, choking) odor.
6. Sulfur trioxide dissolves in concentrated sulfuric acid, forming (dilute sulfuric acid, fuming sulfuric acid, hydrosulfuric acid).
7. After a bottle half full of concentrated sulfuric acid has been exposed to the air for several days it will be found to be (full, empty, unchanged).

8. When a gas is bubbled through concentrated sulfuric acid, the latter acts as a (reducing, catalytic, dehydrating) agent.

9. To test for a sulfate (barium chloride, lead acetate paper, hydrogen sulfide) is used.

10. Sulfuric acid is used to prepare other acids because of its (low cost, high boiling point, high specific gravity).

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

#### A

- ( ) Barium sulfate
- ( ) Fuming sulfuric acid
- ( ) Sodium sulfite
- ( ) Copper sulfide
- ( ) Hot, concentrated sulfuric acid
- ( ) Sulfur trioxide
- ( ) Flowers of sulfur
- ( ) Dilute sulfuric acid
- ( ) Sulfur dioxide
- ( ) Hydrosulfuric acid

#### B

- 1. Sulfurous anhydride
- 2. Salt of sulfurous acid
- 3. Roasting
- 4. Hydrogen sulfide in solution
- 5. Mostly rhombic
- 6. Sulfuric anhydride
- 7. White precipitate
- 8. Black precipitate
- 9.  $\text{H}_2\text{SO}_4 \cdot \text{SO}_3$
- 10. Electrolyte
- 11. Oxidizing agent
- 12. Louisiana

## CHAPTER 13

### CARBON AND SOME CARBIDES

**Importance.**—Among the elements that make up the earth's crust, carbon ranks far from the top in order of abundance. In order of importance, however, it has often been assigned first place. Coal, which is our most valuable source of power, is largely carbon, in both the free and the combined state. Carbon, in the form of diamond, is the most highly prized among all precious stones because of its unmatched brilliance. Moreover, the diamond is the hardest substance known to man, and many of its uses depend on this property. Graphite, another form of free carbon, is employed in making lead pencils, and is also an important lubricant. Carbon, as coke, is a widely used fuel and reducing agent. In the forms of charcoal, boneblack, and lampblack, carbon is employed industrially for a variety of purposes. The two oxides of carbon, the monoxide and the dioxide, play an important part in both commercial and natural processes.

Carbon combined with calcium is the source of acetylene, a valuable fuel gas. Carbon combined with silicon gives us carborundum, the universal grinding agent. Limestone, marble, and other carbonates are carbon compounds of great commercial importance. In combination with hydrogen, this element forms a great number of hydrocarbons, many of them found in petroleum. In addition, there are compounds of carbon, hydrogen and oxygen, often with one or more other elements; the number of these compounds runs into hundreds of thousands. Many are found in the living tissues of plants and animals. Many others have been synthesized in the laboratory. So vast is the field of carbon compounds that the subject is treated separately under the heading of *Organic Chemistry*.

**Occurrence.**—Carbon, in both the free and the combined state, is very extensively distributed in nature.

1. In the *free state* carbon is found in coal, diamond and graphite.

2. In the *combined state* carbon is found in the carbon dioxide of the air; in all carbonate rocks, such as limestone and marble

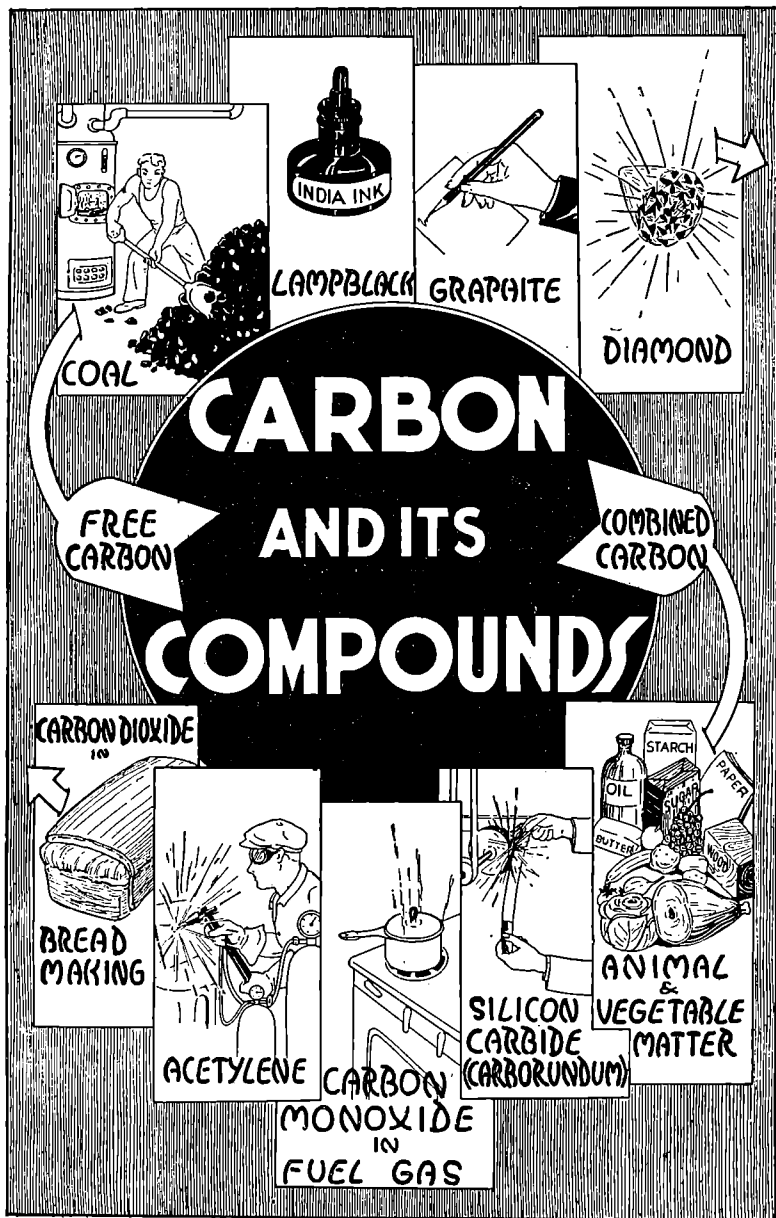


FIG. 45.



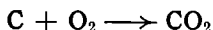
( $\text{CaCO}_3$ ); in marsh gas or methane ( $\text{CH}_4$ ), which is the principal constituent of natural gas; in petroleum, which is almost wholly a mixture of compounds of hydrogen and carbon (hydrocarbons); and in all animal and vegetable matter such as starch, fats, oils, sugar, protein, and cellulose.

**Allotropic Forms of Carbon.**—There are at least eight varieties of carbon, more or less impure, which either exist free in nature or can be prepared artificially. Among these, three definite allotropic forms are recognized. *Diamond* and *graphite* are crystalline varieties. *Amorphous* carbon is the shapeless variety, which includes such commercial products as coke, charcoal, boneblack, and lampblack. The latter is the purest of the amorphous varieties.

### Crystalline Forms of Carbon

1. *Diamond.*—Diamonds are found principally in South Africa, Brazil, and Borneo. Microscopic diamonds were first made artificially in 1893 by Moissan. He dissolved carbon in molten iron, and then, by cooling the mass, subjected the carbon to tremendous pressure, thus imitating the process as it probably takes place in nature. No successful method has ever been developed to manufacture diamonds on a commercially practical scale.

The diamond occurs in the form of octahedral (eight-sided) crystals. It is brilliant and transparent when pure, but is sometimes colored by small amounts of metallic oxides, and may even be black. Because of its ability to reflect and refract light to a high degree, the diamond is used as a jewel. It is brittle, has a density of 3.5, does not conduct electricity, and is insoluble in all common liquids. The diamond is the hardest substance known, and is therefore employed extensively for cutting glass and precious stones, and in rock-drilling, grinding and engraving. Like all other forms of carbon, the diamond combines with oxygen at a high temperature, forming carbon dioxide:



2. *Graphite.*—Graphite is found in New York State, Canada, Ceylon and Siberia. In its natural form, it always occurs mixed with impurities. A superior form of graphite, containing practically no impurities, is made synthetically in the electric furnace by a method invented by the American chemist, Acheson. Anthracite coal is packed loosely in the furnace and an electric current

is sent through it. The high resistance encountered by the current generates tremendous heat, which converts the coal into graphite.

Graphite is a soft, black solid, made up of minute crystals or scales that slide over each other very smoothly. It is virtually infusible, readily conducts the electric current, and is insoluble in all common liquids. Owing to its unusual smoothness and its exceedingly high melting point, graphite is used to lubricate heavy or high-speed machinery. Its great resistance to fusion also renders it useful for making crucibles. When mixed with clay, it becomes the "lead" of lead pencils. Because it is an excellent conductor of electricity, graphite is used to make electrodes for many types of electrical devices, such as arc lamps and electric furnaces. It is also used as an ingredient in stove polishes and certain paints, and as a means of preventing the formation of boiler scale.

### **Amorphous Forms of Carbon**

1. *Coke*.—Coke is made by the destructive distillation of bituminous coal. Destructive distillation is the process of decomposing an organic substance by heating it in the absence of air, and then condensing the resulting vapors. Coke is a steel-gray, hard, porous solid. It combines with oxygen vigorously, producing little flame and no smoke, but evolving much heat; thus, coke is an ideal fuel. It readily removes oxygen from most metallic oxides, and is therefore employed as a reducing agent in metallurgy, and in the manufacture of water gas. It is also used in making carborundum and calcium carbide.

2. *Charcoal*.—Charcoal is made by the destructive distillation of wood. When the wood is heated, certain volatile products are driven off, the most important of which are wood alcohol, acetic acid, and acetone, while charcoal is left behind as a residue. It is a black, porous solid, which possesses the cellular structure of wood. It has the ability to adsorb large volumes of gases, and is therefore employed in gas masks for both industrial and military purposes. A special variety of charcoal to be used for such purposes is made by the destructive distillation of fruit stones and cocoanut shells. Charcoal is also employed as a fuel, because it burns with no smoke and practically no flame. In addition, it serves as a filtering medium for purifying water; as a reducing agent; and as a raw material in the manufacture of gunpowder.

3. *Boneblack*.—Boneblack, or animal charcoal, is made by the

destructive distillation of bones. It is a granular or powdery black solid, consisting of about 10% amorphous carbon and 80% calcium phosphate. Like charcoal, it is very porous, and is therefore used in the refining of sugar because of its ability to remove objectionable gases and coloring matter by adsorption (page 270).

4. *Lampblack*.—Lampblack, a variety of soot, is prepared by burning oil or natural gas in a limited supply of air. The incandescent carbon of the yellow flame is condensed on special surfaces that are maintained at a low temperature by circulating water. Lampblack is a soft, finely divided, black, greasy powder, and is practically pure carbon. It is used as a pigment in making printer's ink and India ink, as well as black paint and shoe polish. It is also used as a "filler" in the manufacture of rubber.

**Coal**.—Coal is fossilized vegetable matter. It is found to some extent in widely scattered regions throughout the world, but the most important deposits are in North America and Asia. Coal was formed underground ages ago by the slow decomposition of wood and vegetation out of contact with the air, and under the influence of high temperature and tremendous pressure. As a result of the chemical action taking place, more and more of the oxygen and hydrogen present in the vegetable matter, together with some other elements, was forced out in the form of volatile products. At the point of greatest decomposition, the remaining mass consisted of nearly pure carbon, with only very little volatile matter left. This end product of the series of changes is hard coal as we know it today.

**Varieties of Coal**.—There are two main varieties of coal, as well as several forms of minor importance. They vary widely in composition, the chief differences consisting in the percentages of free carbon, combined carbon, and volatile matter which they contain.

1. *Anthracite* or *hard coal* contains as high as 90% carbon, nearly all of it in the uncombined state. It is a black, hard, shiny solid. It is ignited with difficulty, and burns with little flame and the evolution of much heat. By far the greater part of the anthracite coal mined is used as a fuel, mostly for domestic purposes. Some is consumed in the manufacture of water gas and of graphite.

2. *Bituminous* or *soft coal* is a black, crumbly solid containing about 75% carbon, a considerable part of which is combined with hydrogen, oxygen, nitrogen, and sulfur. It has a larger percentage

of volatile matter than anthracite. When ignited, it burns with a smoky flame, but because of its greater hydrogen content, yields more heat than does the hard coal. Bituminous coal is used extensively as fuel. When subjected to destructive distillation, it decomposes into coke, coal gas, ammonia, and coal tar.

3. *Lignite* contains even less carbon than does bituminous, and correspondingly more of the other elements. It retains the cellular structure of the wood from which it originated.

4. *Peat* is a brownish-black surface deposit made up of moss, leaves, and other plant debris loosely held together. When its moisture has been pressed out, peat is used as a low-grade fuel.

**Properties of Carbon.**—With but very few exceptions, the physical and chemical properties are the same for all the different forms of carbon.

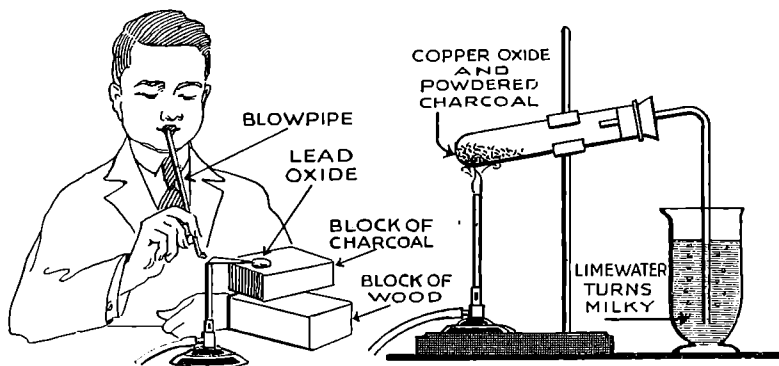
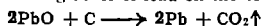
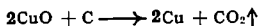


FIG. 46. Carbon as a reducing agent.

A mixture of lead oxide and charcoal is heated on a charcoal block by means of a blowpipe. Carbon dioxide is given off, leaving a small globule of lead on the block:



A mixture of copper oxide and powdered charcoal is heated. The gas given off is passed through limewater, and turns it milky. The residue in the test tube is metallic copper:



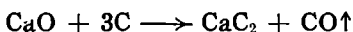
1. *Physical Properties.*—All the allotropic forms of carbon are solid, odorless, and tasteless. All are black, except the diamond. They are all insoluble in ordinary solvents. However, molten iron can absorb 1% or 2% of carbon which, on cooling, may separate as graphite, or react with the iron, forming iron carbide. This property is made use of in manufacturing various grades of iron and steel (page 222 ff.).

2. *Chemical Properties.*—Carbon is inert in the presence of acids or bases. It burns in oxygen, forming carbon dioxide. At low or moderate temperatures, carbon is relatively inactive. At high temperatures, it readily removes oxygen from metallic oxides; hence it is an excellent reducing agent (Fig. 46). At the extremely high temperature of the electric furnace ( $3000^{\circ}\text{C}.$ ), carbon combines with many metals and some non-metals, forming compounds such as calcium carbide ( $\text{CaC}_2$ ), carborundum ( $\text{SiC}$ ), and carbon disulfide ( $\text{CS}_2$ ).

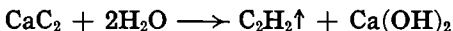
The table below lists some of the more important properties of the various forms of carbon, and indicates certain uses which are dependent on these properties.

VARIETY OF CARBON	PROPERTIES	USES BASED ON PROPERTIES
Diamond	Reflects and refracts light. Hardest known substance.	In jewelry. As an abrasive and cutting material.
Graphite	Very soft and smooth. Difficult to melt. Conducts electricity.	As a lubricant. In making crucibles. In making electrodes.
Bituminous coal	Burns with much heat. Decomposed by destructive distillation.	As a fuel. In making coal gas.
Anthracite coal	Burns with a nearly smokeless flame. Contains a high percentage of free carbon.	As a fuel for domestic purposes. As a reducing agent in making water gas.
Coke	Burns with a hot, smokeless flame. Has a strong attraction for oxygen.	As a domestic and commercial fuel. As a reducing agent in the blast furnace.
Charcoal	Almost pure carbon. Adsorbs dissolved impurities. Adsorbs gases.	As a fuel. As a filtration medium. In making gas masks.
Boneblack	Very porous, therefore has great power of adsorption.	To remove colors and odors from solutions.
Lampblack	Very finely divided black powder.	In making India ink and printer's ink.

**Calcium Carbide ( $\text{CaC}_2$ ).**—Calcium carbide is made by heating a mixture of coke and lime to a high temperature in an electric furnace:



The resulting product is a grayish solid which reacts with water, forming acetylene ( $\text{C}_2\text{H}_2$ ), a gas used extensively for the production of light and heat:



Calcium carbide is also used to produce calcium cyanamide in the fixation of atmospheric nitrogen (page 124).

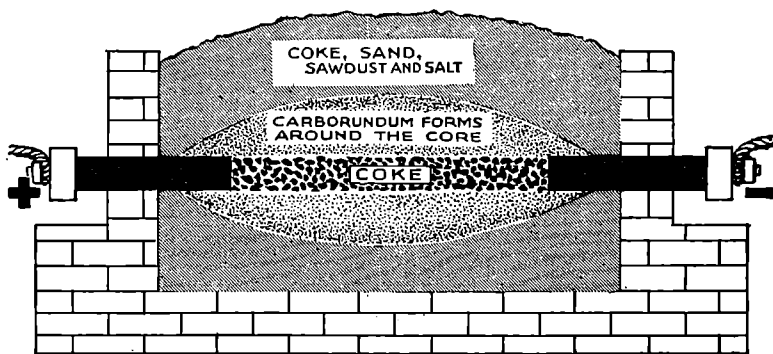
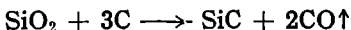


FIG. 47. The manufacture of carborundum. The electric furnace is used to make this and other compounds of carbon because it is a convenient source of high temperatures.

**Silicon Carbide ( $\text{SiC}$ ).**—Silicon carbide, more commonly known as *carborundum*, is made by passing a strong current of electricity through a mixture of sand, coke, sawdust, and salt in an electric furnace of the resistance type (Fig. 47). The sand and coke react according to the following equation:

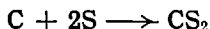


The salt helps the mixture to fuse more readily, while the sawdust renders the mass porous, thus permitting the escape of gases.

Carborundum consists of sharp, flat crystals having a deeply purple iridescence. It is almost as hard as diamond, and is unaffected by all common reagents. Because of its extreme hard-

ness, carborundum is used as an abrasive for cutting, grinding, and polishing metals, stone and glass.

**Carbon Disulfide (CS<sub>2</sub>).**—Carbon disulfide is made by bringing together charcoal and sulfur vapor in the high temperature of the electric furnace:



Carbon disulfide is a heavy, inflammable liquid. Owing to the presence of impurities, it usually has a disagreeable odor. It is used to dissolve rubber, sulfur, and phosphorus. It is also valuable as an insecticide, and in the manufacture of rayon, which is used extensively as a substitute for silk.

### QUESTIONS

1. Name three natural and four artificial forms of carbon.
2. Describe a chemical test to distinguish between manganese dioxide and powdered charcoal.
3. (a) What chemical property is common to all forms of carbon?  
(b) Write the equation illustrating this property.
4. Explain how anthracite coal is formed in nature.
5. Is there any "lead" in a lead pencil? Explain.
6. Why is there a difference in the amount of smoke produced when anthracite and bituminous coal burn?
7. (a) How would you remove the color from a solution of brown sugar? (b) Explain the action involved.
8. What is the simplest method of showing that a compound contains carbon?
9. Mention one property of graphite upon which each of its following uses depends: (a) lubricant, (b) crucibles for melting metals, (c) electrodes for electric furnaces, (d) lead pencils, (e) stove polish.
10. State the difference between free carbon and combined carbon, illustrating each with an example.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Two crystalline varieties of carbon are . . . . .; four amorphous varieties are . . . . .
2. Bituminous coal contains more . . . . . matter than anthracite coal.
3. Two reasons why the diamond is used as a gem are . . . . .

4. Two uses of charcoal are .....
5. Organic chemistry is the study of the compounds of .....
6. Two products formed when lead oxide and carbon are heated together are .....
7. Destructive distillation is employed to obtain coke from ....., charcoal from ....., and boneblack from .....
8. When a solution of brown sugar is passed through boneblack, the ..... of the solution is removed.
9. Carbon disulfide is prepared by heating together ..... and ..... vapor.
10. The common name of silicon carbide is ..... Its chief property is its extreme .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it:*

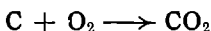
- | A                    | B   |
|----------------------|---|
| ( ) Acheson          | 1. Solvent for rubber                     |
| ( ) Lampblack        | 2. Artificial abrasive                    |
| ( ) Carbon disulfide | 3. Powerful reducing agent                |
| ( ) Bituminous coal  | 4. Low-grade coal                         |
| ( ) Lignite          | 5. To make acetylene                      |
| ( ) Electric furnace | 6. Synthetic graphite                     |
| ( ) Boneblack        | 7. For gas masks                          |
| ( ) Coke             | 8. Rock drill                             |
| ( ) Calcium carbide  | 9. Decomposed by destructive distillation |
| ( ) Powdered diamond | 10. To make India ink                     |
|                      | 11. High temperature                      |
|                      | 12. Contains 80% calcium phosphate        |



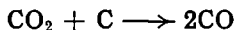
## CHAPTER 14

### OXIDES OF CARBON

Carbon may unite with oxygen in two different ways, depending on the conditions. At ordinary temperatures, carbon burns in air or oxygen, forming carbon dioxide:



If the carbon is red-hot, and carbon dioxide passes through it, as happens in a coal stove, carbon monoxide is formed:



Although so closely related in composition, these two oxides of carbon are directly opposite in most of their properties. The table at the end of this chapter summarizes the principal points of similarity and difference.

#### CARBON DIOXIDE (CO<sub>2</sub>)

**Occurrence.**—Carbon dioxide is present in the air to the extent of about .04% by volume. Certain natural spring water, such as Vichy and Seltzer, is charged with carbon dioxide (carbonated). Some gas wells evolve enormous volumes of gaseous mixtures containing as high as 95% carbon dioxide. Minerals like marble, limestone, and calcite, all of which are forms of calcium carbonate (CaCO<sub>3</sub>), contain chemically united carbon dioxide in vast amounts.

**Sources of Carbon Dioxide in the Air.**—In our study of the atmosphere (page 109), we learned that carbon dioxide is produced in nature by: (1) respiration of animals and plants, (2) combustion of fuels containing carbon, (3) decay of animal and vegetable matter, and (4) volcanic action. All these processes would soon raise the amount of carbon dioxide in the air to the danger point, if it were not for the fact that the gas is used by green plants to manufacture starch. The carbon dioxide cycle operates in such a way as to maintain a balance in nature.

**Preparation**

1. *Action of an Acid on a Carbonate (Laboratory Method).*—The apparatus is set up as shown in Fig. 48. Some marble chips (calcium carbonate) are placed in the bottle, and dilute hydrochloric acid is added through the thistle tube. Carbon dioxide is

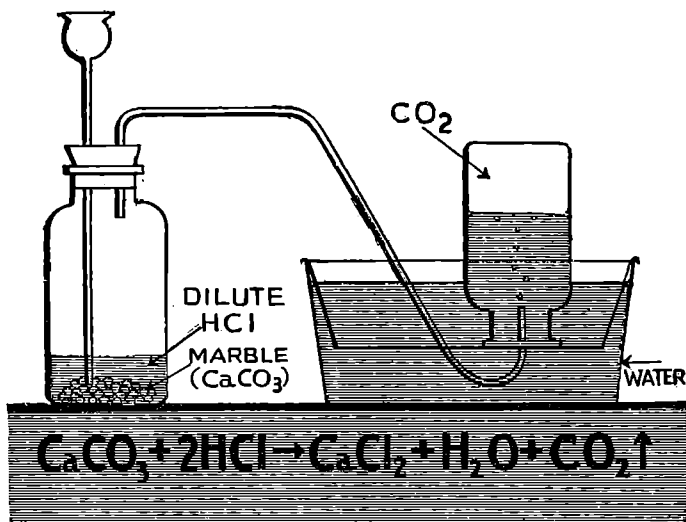
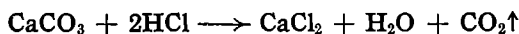
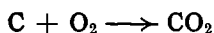


FIG. 48. Laboratory preparation of carbon dioxide.

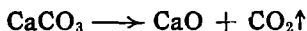
immediately evolved, and may be collected either by displacement of air or by displacement of water, as shown in the diagram. The equation is:



2. *Burning Carbon (Commercial Method).*—The simplest method of preparing carbon dioxide is by the direct union of the two elements. Coke is burned with a good supply of air, and the carbon dioxide is then separated from the other products:

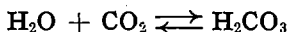


3. *Heating a Carbonate (Commercial Method).*—When limestone (CaCO<sub>3</sub>) is heated, carbon dioxide and quicklime (CaO) are obtained:

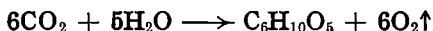


**Physical Properties.**—Carbon dioxide is a colorless gas, having a slight odor and taste. It is about  $1\frac{1}{2}$  times as heavy as air, and may be poured from one vessel into another like water. Carbon dioxide is moderately soluble in water. Soda water and similar carbonated beverages contain an excess of the gas held in solution under pressure. When the pressure is removed, the gas is liberated, thus producing an effervescence. Carbon dioxide is easily liquefied and solidified. When liquefied carbon dioxide is allowed to escape, it evaporates rapidly, causing such a sharp drop in temperature that a portion of the liquid freezes into a white solid resembling snow. This "snow," when pressed into compact form, is the familiar "dry ice" used as a refrigerant.

**Chemical Properties.**—Carbon dioxide is a stable and inactive gas. It does not burn or support combustion, and is therefore employed to extinguish fires. Carbon dioxide combines with water, forming the weak carbonic acid, and is therefore the anhydride of this acid:



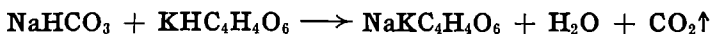
Carbonic acid is unstable; when warmed, it breaks up into carbon dioxide and water. In the leaves of green plants, carbon dioxide combines with water, under the influence of sunlight and chlorophyll, forming starch (process of *photosynthesis*):



### Uses

1. *Carbonated Beverages.*—About 90% of the carbon dioxide made in this country is used in the soft drink industry.

2. *Leavening Agent.*—In the baking of bread, carbon dioxide is formed in the dough by the action of yeast on sugar (page 186), or by the action of water on baking powder. The latter consists of a mixture of sodium bicarbonate (baking soda) and some acid substance, such as potassium acid tartrate, commonly known as cream of tartar ( $\text{KHC}_4\text{H}_4\text{O}_6$ ). The following reaction takes place in the dough:



The bubbles of gas expand and cause the dough to rise, thus making the bread light and digestible.

3. *Plant Food.*—Leaves absorb carbon dioxide from the air, forming starch and liberating free oxygen. As a result of this

process of photosynthesis, animals are able to secure their food supply, and at the same time to obtain oxygen for respiration.

4. *Solvay Process.*—Large quantities of sodium carbonate and sodium bicarbonate are made by this process (page 204).

5. *Refrigeration.*—For this purpose, liquid carbon dioxide is sometimes used instead of ammonia, especially on ships. “Dry ice” (solid carbon dioxide) is used to preserve ice cream and other perishable foods during shipment. This new refrigerant is superior to ice for these purposes, because (a) it has a much lower temperature ( $-79^{\circ}\text{C}.$ ), and (b) it produces no liquid on melting, but changes directly into a gas.

6. *Fire Extinguishing.*—Carbon dioxide does not support combustion, and, since it can be generated in large quantities both quickly and economically, it is used in several ways as a means of putting out fires.

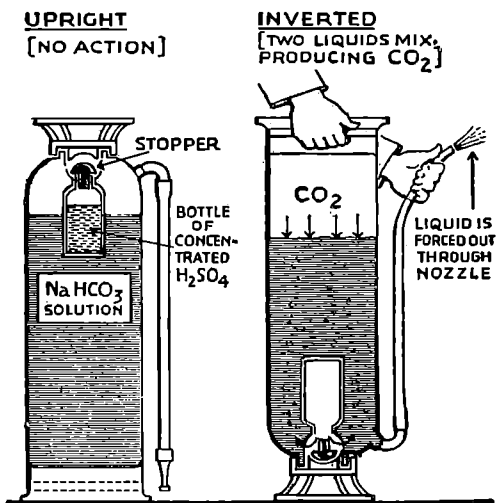


FIG. 49. The ordinary carbon dioxide fire extinguisher.

**Carbon Dioxide Fire Extinguisher.**—The ordinary carbon dioxide or soda-acid fire extinguisher (Fig. 49) contains a concentrated solution of sodium bicarbonate ( $\text{NaHCO}_3$ ) and a bottle of sulfuric acid fitted with a loose lead stopper. When the extinguisher is inverted, acid pours out of the bottle and reacts with the sodium bicarbonate solution, producing carbon dioxide gas under pressure:



The stream of liquid and carbon dioxide forced out of the nozzle is usually successful in putting out small fires.

**Foamite Method for Extinguishing Fires.**—For larger fires, especially oil conflagrations, the foamite method has been found effective. A foam containing carbon dioxide is formed by the action between solutions of aluminum sulfate and sodium bicarbonate containing licorice extract:



The jelly-like aluminum hydroxide and licorice extract form a mixture which is puffed up by the carbon dioxide bubbles, creating a tough, light foam. This foam covers the burning substance and smothers the fire by excluding oxygen (Fig. 50).

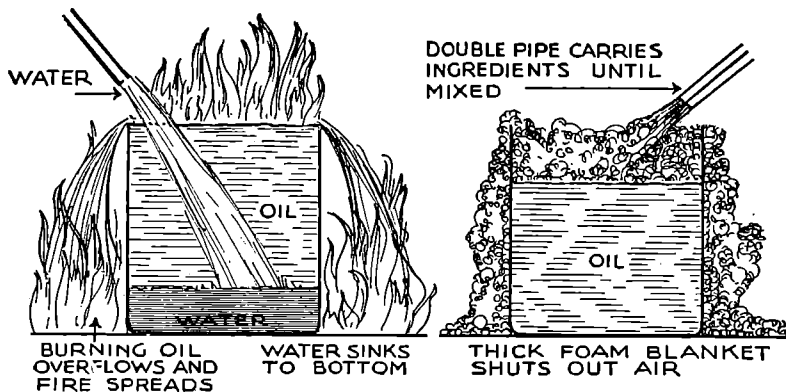
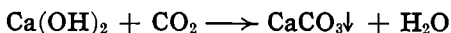


FIG. 50. The *wrong* way and the *right* way of fighting an oil fire.

**Test for Carbon Dioxide and Carbonates.**—To prove that a gas is carbon dioxide, allow it to bubble through clear limewater. Carbon dioxide is the only gas which will turn the limewater milky. This is the method used to demonstrate that exhaled air contains carbon dioxide:

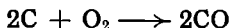


To test for a carbonate, add hydrochloric acid to the substance. If a gas is given off which turns limewater milky, the substance being tested is a carbonate.

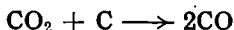
## CARBON MONOXIDE (CO)

**Preparation**

1. *Burning Carbon in a Limited Supply of Air.*—When carbon is made to burn with an insufficient air supply, it is not completely oxidized, thus forming the monoxide instead of the dioxide:



2. *Reducing Carbon Dioxide with Hot Carbon.*—Highly heated carbon is a powerful reducing agent. It converts carbon dioxide to carbon monoxide:



This reaction takes place in the ordinary coal stove (Fig. 51). Near the bottom, where there is a plentiful air supply, carbon dioxide

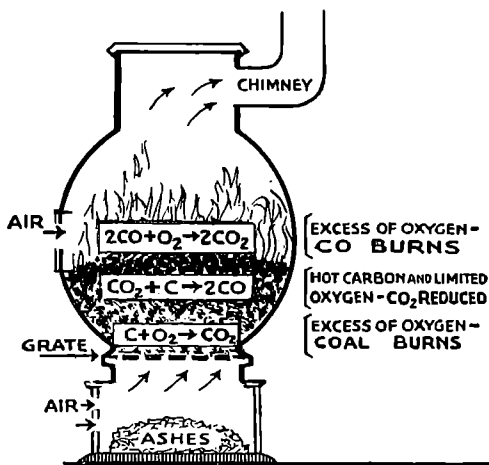
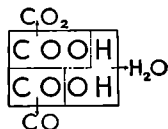


FIG. 51. Three chemical actions that take place in a coal stove.

is formed. This gas, as it passes through the heated coal above, is reduced to carbon monoxide, which continues its way to the top. There it burns with the familiar blue flame seen over a coal fire.

3. *Heating Oxalic Acid with Concentrated Sulfuric Acid (Laboratory Method).*—Oxalic acid ( $\text{H}_2\text{C}_2\text{O}_4$ ) is an organic acid whose formula is sometimes written  $(\text{COOH})_2$ . The sulfuric acid acts as a dehydrating agent, extracting water and leaving carbon

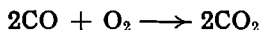
dioxide and carbon monoxide. The accompanying diagram shows how the molecule of oxalic acid is decomposed.



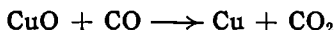
The carbon dioxide is absorbed by passing the mixture of gases through a concentrated solution of sodium hydroxide, in which carbon monoxide is insoluble. The carbon monoxide is collected by displacement of water.

**Physical Properties.**—Carbon monoxide is a colorless, odorless, and tasteless gas. It is somewhat lighter than air, and only slightly soluble in water. Carbon monoxide is extremely poisonous.

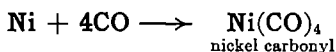
**Chemical Properties.**—The chemical properties of carbon monoxide depend chiefly on the fact that it is able to take on an additional atom of oxygen. In doing so, it burns with a hot, blue flame, forming carbon dioxide:



Carbon monoxide acts as a powerful reducing agent, principally in metallurgy, by removing oxygen from metallic oxides, and so liberating the free metal:

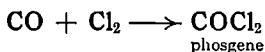
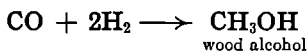


It combines with certain metals at high temperatures, forming *carbonyls*:



Nickel carbonyl is employed in the metallurgy of nickel.

In the presence of catalysts, carbon monoxide combines directly with hydrogen, yielding wood alcohol; it also combines with chlorine, yielding phosgene, a poison gas used in warfare.



**Physiological Properties.**—Carbon monoxide causes more accidental deaths each year than does any other poison known to civilized man. A concentration of as little as 0.1% in the atmosphere is dangerous, and may bring on death within a short time. The poisonous character of carbon monoxide is due to the strong

attraction which it has for the hemoglobin in the red corpuscles of the blood. When the gas enters into combination with the hemoglobin, it renders the corpuscles incapable of absorbing oxygen and of carrying this vital substance to the body cells. The chief causes of carbon monoxide poisoning are exposure to the following:

1. *Automobile Exhaust Gases.*—These gases contain a rather high percentage of carbon monoxide, resulting from the incomplete combustion of the gasoline.

2. *Producer and Blast Furnace Gases.*—These are industrial fuels which contain as high as 30% carbon monoxide.

3. *Escaping Gases from Heating Appliances.*—These result from leaking tubes and fixtures in the home, as well as from the evolution of carbon monoxide by a coal stove in which an improper adjustment of drafts has resulted in incomplete combustion of the fuel.

**Uses.**—Carbon monoxide is used as a fuel, being an important constituent of water gas and of producer gas. It serves as a reducing agent in the metallurgy of iron. As explained above, carbon monoxide is employed in making wood alcohol and phosgene, and also in extracting nickel from its ore.

**Test for Carbon Monoxide.**—A given gas is carbon monoxide if it burns with a pale-blue flame, and if the product of this burning proves to be carbon dioxide by forming a white precipitate when bubbled through clear limewater.

**Comparison of the Oxides of Carbon.**—Although the two oxides of carbon are so closely related in composition, we have

CARBON DIOXIDE	CARBON MONOXIDE
Colorless.	Colorless.
Has very slight odor and taste.	Odorless and tasteless.
About $1\frac{1}{2}$ times as heavy as air.	Somewhat lighter than air.
Moderately soluble in water.	Very slightly soluble in water.
Reacts with water, forming carbonic acid.	Does not react with water.
Liquefied with ease.	Liquefied with difficulty.
Has no tendency to combine with more oxygen.	Can combine with more oxygen, hence is a reducing agent.
Does not burn.	Burns with a hot flame.
Used in putting out fires.	Used as a fuel.
Non-poisonous.	Extremely poisonous, even in small amounts.



seen that, in many respects, their properties are strikingly dissimilar. Carbon dioxide, being stable, is rather inactive chemically. Its fellow-compound is of an unstable molecular structure; therefore, it is "eager" to acquire another atom of oxygen and so to achieve the stable structure of carbon dioxide. This explains why carbon monoxide is so extremely active chemically. The table on page 160 contrasts in summary form the chief properties of the two oxides.

### QUESTIONS

1. (a) Name two methods for obtaining carbon dioxide on a commercial scale. (b) Write the equation for each method.

2. In the laboratory preparation of carbon dioxide, describe two methods by which the gas may be collected and give reasons why either method may be used.

3. (a) Write the equation for the process of photosynthesis. (b) Explain why this process cannot be duplicated in the laboratory.

4. (a) Why is it not wise to throw water on an oil fire? (b) Why does an ordinary soda-acid fire extinguisher fail to put out such a fire? (c) Explain an effective method for using carbon dioxide to extinguish an oil fire.

5. (a) Carbon dioxide is the anhydride of what acid? (b) Write the equation for the formation of this acid.

6. Given five bottles, each containing one of the following gases: (a) carbon dioxide, (b) carbon monoxide, (c) hydrogen, (d) nitrogen, and (e) nitric oxide. How would you identify each gas?

7. Explain why it is dangerous to run an automobile engine in a closed garage.

8. Explain why carbon monoxide is formed when (a) coal burns in a stove, (b) a pot of cold water is placed over a gas flame.

9. Why does carbon monoxide act as a reducing agent?

10. Tobacco smoke is believed to contain a small quantity of carbon monoxide resulting from incomplete combustion. What precautions should a smoker take to guard himself against this poison?

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Carbon dioxide is evolved in nature by the following three processes: . . . . . Green plants remove carbon dioxide from the air by a process called . . . . .

2. Paper contains the element . . . . ., as shown by the fact that when it burns it produces . . . . . gas which, on being passed through limewater, forms a . . . . . precipitate.

3. The stream from a soda-acid fire extinguisher puts out a fire because it contains . . . . . and . . . . ., which shut out . . . . .

4. Dry ice is . . . . . in . . . . . form, and is used as a . . . . .

5.  $\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \longrightarrow$  . . . . .

6. When . . . . . is passed over red-hot . . . . ., carbon monoxide is formed. This process is called . . . . .

7. The poisonous properties of carbon monoxide are due to the strong affinity which it has for the . . . . . in the . . . . ., thus preventing . . . . . from being carried to the . . . . . of the body.

8. Carbon monoxide is an excellent . . . . . agent because of its strong attraction for . . . . .

9. In the preparation of carbon monoxide from oxalic acid, the sulfuric acid acts as a . . . . . agent.

10. The percentage compositions of the two oxides of carbon illustrate the Law of . . . . .

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

#### A

- ( ) Marble
- ( ) Carbonated water in nature
- ( ) Uses aluminum sulfate
- ( )  $\text{CO}_2 + \text{C} \longrightarrow 2\text{CO}$
- ( ) Laboratory preparation of CO
- ( ) Carbon dioxide cycle
- ( ) Reducing agent in metallurgy
- ( ) Carries oxygen in the blood
- ( ) Green coloring matter in plants
- ( ) Dry ice

#### B

- 1. An exchange between plants and animals
- 2. Chlorophyll
- 3. Hemoglobin
- 4. Vichy
- 5. Limewater
- 6. Changes directly to a gas
- 7. Calcium carbonate
- 8. Foamite fire extinguisher
- 9. Sodium bicarbonate
- 10. Oxalic acid
- 11. Carbon monoxide
- 12. In the center of a coal fire

## CHAPTER 15

### LIQUID AND GASEOUS FUELS

**Definition.**—A *fuel* is a substance which is burned in air to produce heat and power for domestic and industrial purposes. Carbon and hydrogen are the important heat-producing elements in every fuel.

**Desirable Properties of Fuels.**—The following considerations determine the practical value of a substance which is to be used as a fuel: (1) its price and the readiness with which it may be obtained; (2) its calorific, or heat-producing, value; (3) its ease of ignition; (4) the convenience with which it may be fed to the fire; (5) the cost of transportation and storage; and (6) the amount of smoke and other waste materials produced.

**Classes of Fuels.**—The common fuels may be grouped according to the following table:

	SOLIDS	LIQUIDS	GASES
Natural Fuels	Wood Peat Coal	Petroleum	Natural gas
Artificial Fuels	Charcoal Coke	Gasoline Kerosene Alcohol	Coal gas Water gas Producer gas Acetylene

NOTE.—The principal solid fuels mentioned in this table have already been treated in Chapter 13.

### LIQUID FUELS

**Petroleum.**—Next to water, petroleum or crude oil is the most important liquid occurring in nature. It is a dark-colored, oily, inflammable substance having an unpleasant odor, and is found deep below the earth's surface in many different regions. It is

believed to have been formed by the decomposition of marine organisms, the remains of which collected in natural underground reservoirs. Petroleum is a mixture of many hydrocarbons (compounds of hydrogen and carbon) ranging in formula from  $C_5H_{12}$  to  $C_{60}H_{122}$ .

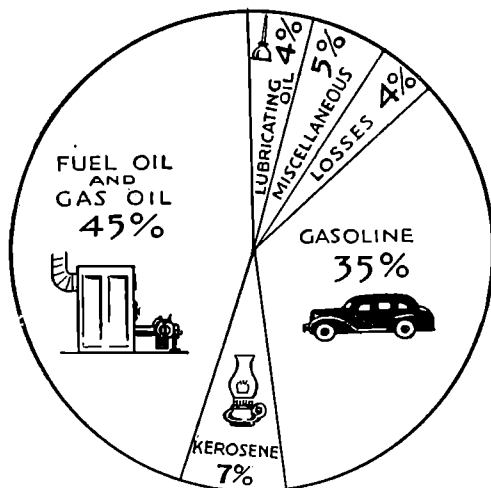


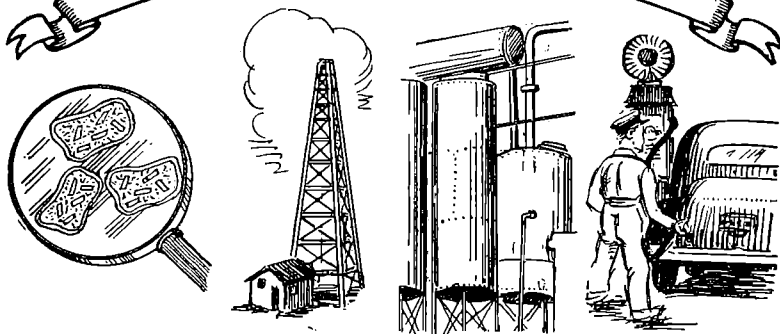
Fig. 52. What happens to petroleum.

**Refining of Petroleum.**—The crude petroleum is transported to the oil refineries, mainly by means of pipe lines, which are sometimes as long as 1000 miles. At the refineries, the oil is subjected to *fractional distillation*, which is a process of separating by distillation a mixture of liquids having different boiling points. Each fraction

is then redistilled to obtain a better separation of the products. The table below summarizes, in a general way, the chief products obtained from petroleum. The relative amounts of these products are represented graphically in Fig. 52.

PRODUCT	RANGE OF BOILING POINTS	USES
Petroleum ether . . . . .	40° — 70°C.	Solvent
Naphtha . . . . .	80° — 120°C.	Solvent, fuel
Gasoline . . . . .	60° — 190°C.	Solvent, fuel
Kerosene . . . . .	150° — 250°C.	Illuminant, fuel
Fuel oil . . . . .	250° — 350°C.	Power and heat
Gas oil . . . . .	290° — 400°C.	"Enriching" water gas
Lubricating oils . . . . .	400° — 450°C.	Lubrication
Vaseline	Above 450°C.	{ Medicinal purposes Candles Paving and roofing
Paraffin		
Asphalt		
Petroleum coke . . . . .	Solid residue	Fuel

# STORY OF GASOLINE



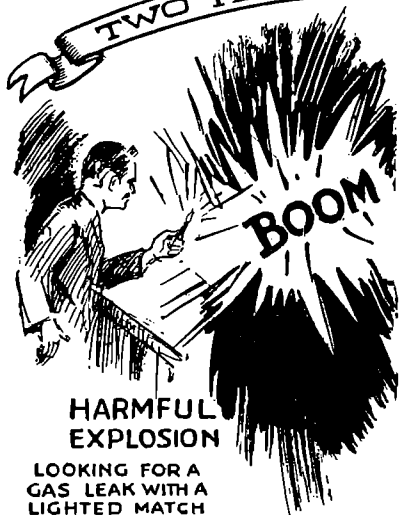
• **DIATOMS**  
MICROSCOPIC MARINE ANIMALS LIVED AND DIED COUNTLESS AGES AGO. THEIR REMAINS FORMED

**PETROLEUM**  
WHICH IS PUMPED FROM THE EARTH AND SHIPPED TO THE

**REFINERY**  
WHERE IT IS SUBJECT-ED TO A PROCESS OF SEPARATION, YIELDING AMONG MANY OTHER SUBSTANCES

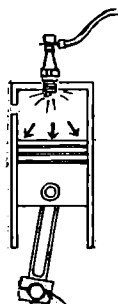
**GASOLINE**  
THE MOST VALU-ABLE SINGLE PRO-DUCT OBTAIN-ED FROM PETROLEUM

# TWO KINDS OF EXPLOSIONS



## HARMFUL EXPLOSION

LOOKING FOR A GAS LEAK WITH A LIGHTED MATCH RESULTS IN A HARMFUL EXPLOSION



## USEFUL EXPLOSION

IGNITING A MIXTURE OF GASOLINE VAPOR AND AIR IN AN AUTOMOBILE ENGINE RESULTS IN A USEFUL EXPLOSION

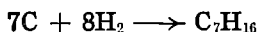
FIG. 53.

**Gasoline.**—Gasoline, which is today the most important single product obtained from petroleum, is a mixture of volatile hydrocarbons, chiefly *hexane* ( $C_6H_{14}$ ) and *heptane* ( $C_7H_{16}$ ). Its average boiling point is about  $80^\circ C$ . Gasoline is used almost exclusively in the internal combustion engine. Here, it is vaporized in the carburetor, the vapor is mixed with the proper amount of air, and the mixture is exploded in the cylinders. The energy of this explosion is the source of power in the motor.

**Increasing the Production of Gasoline.**—The phenomenal growth of the automobile industry has stimulated the demand for gasoline to such a degree that chemists have been forced to devise new methods for producing the fuel. This has been done both by increasing the yield from petroleum, and by deriving the substance from other sources. One method of each type is now in successful use.

1. *Cracking of Oils (Burton Process).*—By subjecting the heavier fractions of petroleum to high temperature and pressure, the more complex hydrocarbon molecules are “cracked” or broken down into the lighter and more volatile molecules of which gasoline is composed. By this means, the yield of gasoline has been increased nearly fourfold.

2. *Hydrogenation of Coal (Bergius Process).*—By heating low-grade coal with hydrogen under extremely high pressure, and in the presence of a catalyst, a fairly satisfactory synthetic gasoline has been produced. The reaction may be represented by the equation:



### GASEOUS FUELS

**Natural Gas.**—Natural gas is believed to have been formed as a result of the same processes of decomposition as were probably responsible for the formation of petroleum. This gas is chiefly methane ( $CH_4$ ), but contains also some hydrogen and carbon monoxide. Over half of the gas used in this country for heating and lighting is natural gas.

**Coal Gas.**—Coal gas is made by heating soft coal out of contact with the air, thereby decomposing it into simpler substances (*destructive distillation*). It contains about 40% methane, 40% hydrogen, and 10% carbon monoxide. In addition to coal gas, some other important products resulting from this process are ammonia, coal tar, and coke. From the coal tar are obtained

many hydrocarbons and other compounds which serve as the starting point for the manufacture of drugs, dyes, explosives, solvents, and a great variety of other valuable products.

A gas plant (Fig. 54) consists of a group of large retorts or ovens where the coal is heated, and a series of chambers through which the vaporized products of decomposition pass. The other volatile substances are separated from the coal gas by washing with water

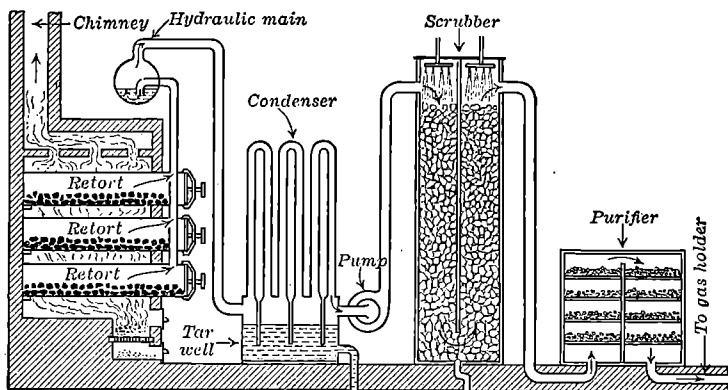
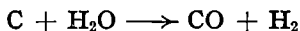


FIG. 54. Plant for making coal gas.

and by condensation. Some impurities of an acid nature, such as compounds of sulfur, are removed from the gas by neutralization with iron oxide or lime.

**Water Gas.**—Water gas is made by reducing steam with white-hot coke (Fig. 55):



After this action has gone on for a few minutes, the coke must be reheated. This is accomplished by a shifting of valves, which allows a forced draft of air to burn some of the coke and thus heat the rest of the coke white-hot again. Since both carbon monoxide and hydrogen burn with an almost colorless flame, water gas cannot be used for illuminating purposes unless it has first been *enriched*. In this process, the gas is mixed in a hot fire-brick chamber (carburetor) with the vapors of gas-oils from petroleum. In the superheater, a still hotter chamber, the hydrocarbons of the gas-oils are cracked, forming simpler compounds which remain in

the vapor state, and which serve to give a luminous flame when the gas is burned.

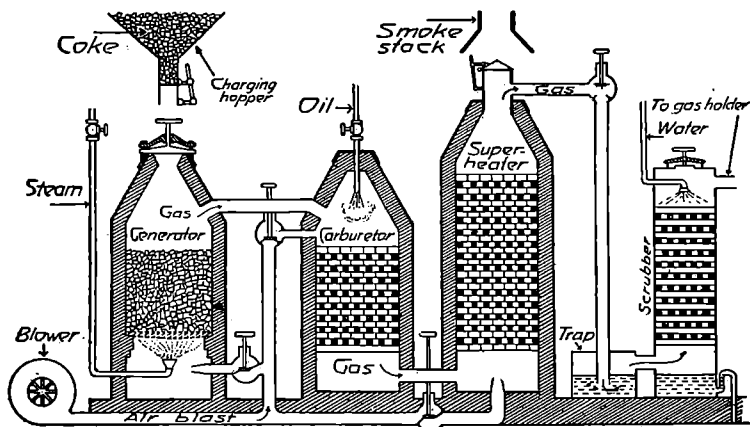
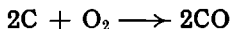


FIG. 55. Plant for making water gas.

**Producer Gas.**—Producer gas is made by burning low-grade coal in a limited supply of air (Fig. 56):



Sometimes steam may be admitted with the air; this is

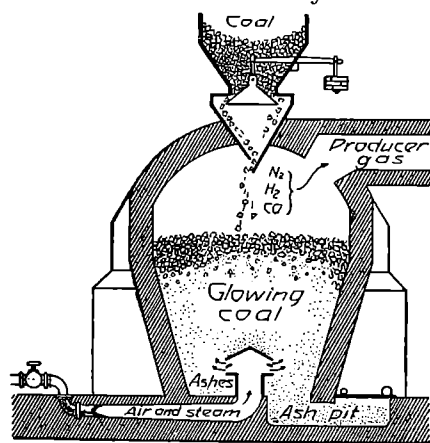
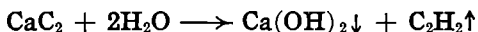


FIG. 56. Apparatus for making producer gas.

then reduced to hydrogen by being passed through the bed of red-hot coal, as in the manufacture of water gas. The main fuel constituent of producer gas is carbon monoxide, together with some hydrogen. It contains as high as 63% inert, unburnable nitrogen, which comes from the air used in making the gas. Producer gas, because it is so cheap, is one of the chief fuel gases employed in many industrial processes.



**Acetylene.**—Acetylene ( $C_2H_2$ ) is made by the action of water on calcium carbide ( $CaC_2$ ):



It is produced in automatic generators which permit water to remain in contact with calcium carbide only long enough to evolve the gas as rapidly as is required. In a limited supply of air,

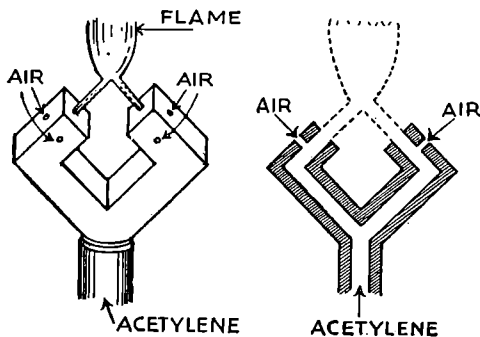
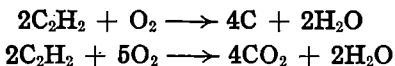


FIG. 57. Acetylene burner.

acetylene burns incompletely with a smoky, yellow flame. In an excess of air, it burns completely with a dazzling white light. These actions may be represented by the following equations:



A special burner must be used (Fig. 57) to insure an adequate air supply. The chief industrial use of acetylene is in the *oxy-acetylene torch*, which can attain a temperature of  $3500^\circ C$ . This torch is employed for cutting and welding metals.

## FLAMES

**The Nature of a Flame.**—A flame is a burning gas. Some fuels produce a flame directly, while others must be partially changed in order to produce a flame when burning.

1. A gaseous fuel like *acetylene* always produces a flame.
2. A liquid fuel like *gasoline* simply vaporizes and burns with a flame.

3. Some solid fuels like *paraffin* melt and vaporize, as in a burning candle.

4. *Wood* and *soft coal* are solid fuels that are decomposed by heat, yielding volatile products which burn with a flame.

5. *Coke* and *charcoal* contain little volatile matter, and therefore burn with practically no flame.

6. Hot particles of *iron* produce sparks but no flame, because the iron does not vaporize.

**The Color of a Flame.**—Most fuels contain hydrocarbons, which are decomposed into hydrogen and carbon at the high temperature of combustion. The hydrogen burns immediately, while the free carbon is heated until it becomes incandescent (red-hot or white-hot) before it begins to burn. This produces the familiar yellow color that is visible in most flames. Frequently, a compound, when vaporized in a flame, imparts to it a particular color which serves to identify the metallic portion of the compound. Thus, the following metals, in the form of their compounds, impart to a flame the colors indicated: sodium, *yellow*; potassium, *violet*; barium, *green*; strontium, *bright-red* (see *Flame Tests*, page 321).

**The Bunsen Burner.**—The Bunsen burner is a device for mixing gas with air and burning the mixture so as to produce a non-

luminous flame having a high temperature. The construction of the burner is shown in Fig. 58. Gas is admitted through the inlet in the base and, in passing through the small hole in the *spud*, it increases its speed of flow. This causes air to be drawn in through the hole in the *ring*. The latter may be adjusted so as to regulate the amount of air admitted. The gas and air are then mixed in the *barrel*, and burn at the top.

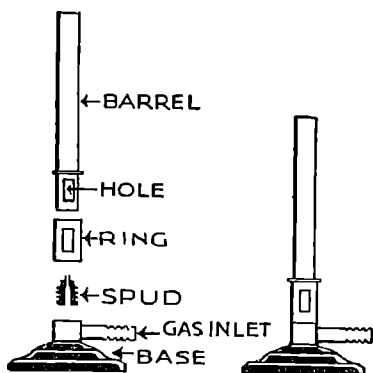


FIG. 58. Bunsen burner.

**The Parts of a Bunsen Flame.**—The non-luminous Bunsen flame has two distinct regions (Fig. 59): the *inner cone*, where the mixture of gas and air has not yet begun to burn; and the *outer*

*cone*, where the combustion takes place. A substance may be *oxidized* by placing it in the upper part of the outer cone, because here the temperature is very high, and there is an excess of air. A substance may be *reduced* by placing it just above the tip of the inner cone, because here the gas is beginning to burn, and oxygen is needed. In the laboratory, a *blowpipe* (page 148) is used to produce oxidizing and reducing flames, and to direct them where they are wanted.

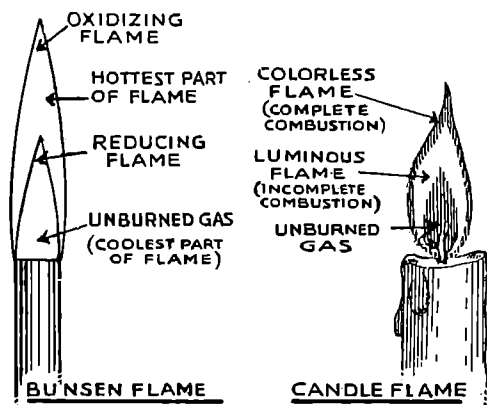


FIG. 59. Flames.

**The Parts of a Candle Flame.**—The candle flame has three distinct regions (Fig. 59). Surrounding the wick, is the *inner cone* of unburned gas resulting from the vaporization of the paraffin. The *middle cone*, the largest one of the three, contains incandescent carbon, and is yellow in color. The *outer cone* is the region of complete combustion, and is nearly invisible.

**Explosive Mixtures.**—An explosive mixture is a combination of gas and air in approximately the correct proportions for complete combustion. When such a mixture is ignited, the entire mass burns instantaneously, with a violent expansion in volume. Explosive mixtures occur in:

1. A *mine*, where the air is mixed with methane.
2. A *gas-filled room*, where the air is mixed with illuminating gas.
3. A *gasoline tank*, if the air becomes mixed with gasoline vapor.

4. An *internal combustion engine*. In this case air and gasoline in vapor form are mixed in the right proportion and exploded in the cylinder by means of an electric spark. The resulting explosion furnishes the energy to operate the engine.

### QUESTIONS

1. (a) Explain the difference between fractional distillation and destructive distillation. (b) Give an industrial example of each.

2. Name five products derived from petroleum in the order in which they are liberated during the refining process.

3. (a) Why is gasoline better than kerosene for use in an automobile engine? (b) Why is kerosene better than gasoline for use in an oil lamp?

4. (a) Name two products formed when kerosene burns. (b) How would you prove that they are formed?

5. What is meant by "lean" and "rich" mixtures in the operation of an automobile?

6. Which is a greater source of danger, escaping water gas or escaping coal gas? Explain.

7. (a) Under what conditions will acetylene burn with (1) a smoky, yellow flame, (2) a dazzling, white flame? (b) Write the equation for the reaction in each case.

8. How does the presence of about 60% nitrogen in producer gas affect its value as a fuel?

9. How would you adjust the burner of your gas stove if the flames had yellow tips?

10. Some explosive mixtures do useful work, while others cause damage. Give two examples of each kind.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Petroleum consists mainly of organic compounds called . . . . .

2. When acetylene burns, a great deal of heat is evolved because the gas contains a high percentage of the element . . . . .

3. The process of heating hydrocarbons having high boiling points to convert them into compounds having lower boiling points is called . . . . .

4. The names and formulas of the two principal compounds in gasoline are .....

5. Water gas is made by the reaction between ..... and .....  
The two chief substances present in water gas are .....

6.  $\text{CaC}_2 + \text{H}_2\text{O} \longrightarrow$  .....

7. A gaseous product, derived from soft coal, that turns moistened red litmus paper blue is .....

8. The poisonous substance present in producer gas is .....

9. The oxidizing portion of a Bunsen flame is the ..... part of the ..... cone.

10. The ..... color observed in the luminous portion of a candle flame is due to the presence of .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

#### A

- ( ) A natural solid fuel
- ( ) Acetylene
- ( ) Gas oil
- ( ) To make synthetic gasoline
- ( ) Producer gas
- ( ) Temperature of 3500°C.
- ( ) From petroleum oil
- ( ) Regulates the air supply
- ( ) Burns without a flame
- ( ) Natural gas

#### B

- 1. Hydrogenation of coal
- 2. Made from low-grade coal
- 3. Bunsen burner collar
- 4. Explosive mixture
- 5. Peat
- 6. Charcoal
- 7. Destructive distillation
- 8. To "enrich" water gas
- 9. Chiefly methane
- 10. Made from calcium carbide
- 11. Paraffin
- 12. Oxy-acetylene flame

## CHAPTER 16

### ORGANIC CHEMISTRY

**Importance.**—Organic chemistry is the study of the compounds of carbon. This field is so vast that it embraces every phase of our modern civilization—indeed the very existence of life itself. The cell, which is the unit of all living matter, is made up of complex substances, all of which are compounds of carbon. The food we eat consists almost entirely of organic compounds. The textile fibers for our clothing, the wood for our furniture and buildings, the paper in our books and newspapers, the gasoline that runs our automobiles are all carbon compounds, as are thousands of other everyday materials such as dyes, medicines, drugs, perfumes, soaps, rubber, explosives, etc.

It was formerly believed that organic compounds could be made only with the aid of some mysterious force that is present in the living plant or animal organism; hence the name "organic." This concept was shattered when, in 1828, the German chemist, Wöhler, succeeded in synthesizing urea from inorganic substances. Nevertheless, we still cling to the term "organic chemistry," which now has come to mean simply the *chemistry of the carbon compounds*.

Various estimates place the number of organic compounds now known at anywhere from 225,000 to 500,000. As a striking contrast to these figures, all known compounds of all the elements other than carbon total up to no more than 26,000. Since carbon is not a very active element, it is surprising to find that it is present in so many compounds. The chief reason is that carbon atoms have the peculiar property of combining with each other, forming *chains* or *rings*, which contain from two to as many as sixty atoms of carbon.

**Structural Formulas.**—The study of organic compounds has been immeasurably simplified by representing their formulas in a graphic manner. A structural formula consists of *symbols*, to represent atoms, and *lines* or *bonds* to represent valence. Thus

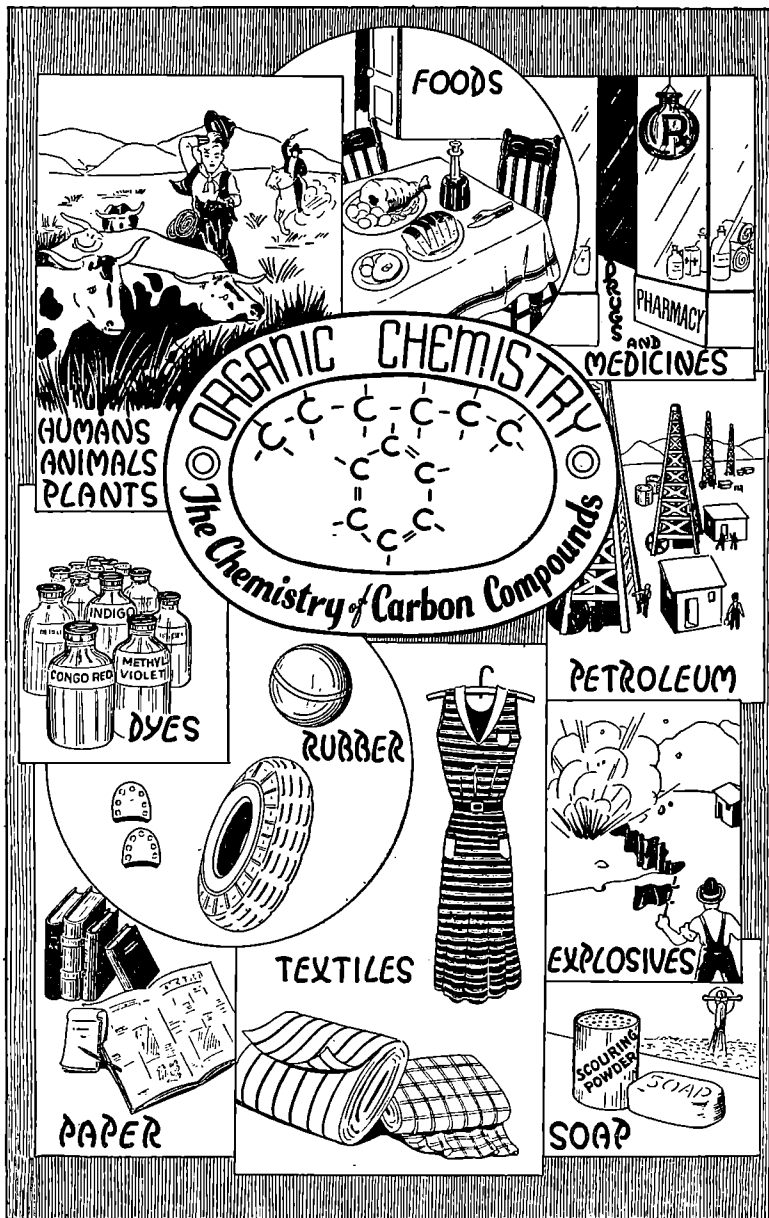
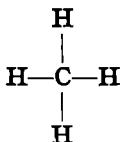
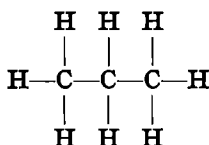


FIG. 60.

the structural formula of methane, the simplest organic compound, is



This indicates that one atom of tetravalent carbon is combined with four atoms of univalent hydrogen. The structural formula of propane ( $\text{C}_3\text{H}_8$ ) is written thus:



**Classification.**—Practically all organic compounds are grouped in the following classes:

- |                 |              |                  |
|-----------------|--------------|------------------|
| 1. Hydrocarbons | 3. Aldehydes | 5. Esters        |
| 2. Alcohols     | 4. Acids     | 6. Carbohydrates |

### HYDROCARBONS

A hydrocarbon is a compound containing only hydrogen and carbon. Hydrocarbons occur extensively in petroleum and fuel gases. Methane, acetylene, benzene, turpentine, and naphthalene are all important hydrocarbons. Methane is the simplest of all, its formula being  $\text{CH}_4$ . It is the initial compound of the *methane* or *paraffin* series, the first five members of which are:

Methane	$\text{CH}_4$	Butane	$\text{C}_4\text{H}_{10}$
Ethane	$\text{C}_2\text{H}_6$	Pentane	$\text{C}_5\text{H}_{12}$
Propane	$\text{C}_3\text{H}_8$		

**Homologous Series.**—A study of the composition and structure of organic compounds reveals that they may be conveniently grouped into families. In the methane series above, it will be noted that the difference between any two successive members is  $\text{CH}_2$ . Such a group is called an homologous series. It should be further noted that the number of hydrogen atoms in a molecule

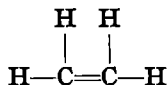


of each of these compounds is twice the number of carbon atoms plus 2. The general formula for the members of the methane series is, therefore,  $C_nH_{2n+2}$ . In general, there is a gradation of properties such as physical state, boiling point, etc., as we pass from one member of an homologous series to the next.

**Methane.**—Methane is a colorless, odorless gas which burns with a faintly luminous flame. It is formed by the decomposition of vegetable matter under water in marshy regions, hence its common name, *marsh gas*. Methane is frequently found in coal mines, where it is a great source of danger because it forms explosive mixtures with air. Miners call the gas *fire damp*. Methane is the principal constituent of natural gas, and is one of the important ingredients present in most fuel gases.

**Saturated and Unsaturated Compounds.**—In the members of the methane series, each carbon atom is joined to an adjacent carbon atom by a single valence bond. Such compounds are said to be *saturated*. There exist other carbon compounds where a carbon atom has one or more of its four valences unsatisfied. These unsatisfied carbon atoms form what is known as *unsaturated* compounds. Their structural formulas show neighboring carbon atoms joined by double or triple bonds. Ethylene and acetylene are examples of unsaturated compounds.

**Ethylene Series.**—In this series, two adjacent carbon atoms are linked by a double bond. The general formula of a member of this series is  $C_nH_{2n}$ . The first compound of this group is *ethylene* ( $C_2H_4$ ). Its structural formula is



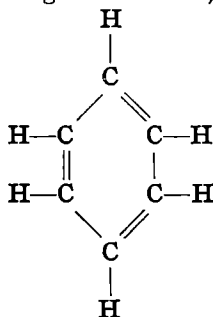
Ethylene is a colorless gas which burns with a luminous flame. It is used to ripen fruit artificially, and has recently been found to possess valuable anaesthetic properties.

**Acetylene Series.**—In this series, two adjacent carbon atoms are linked by a triple bond. The general formula is  $C_nH_{2n-2}$ . The first member of this group is *acetylene* ( $C_2H_2$ ). Its structural formula is



The preparation and properties of acetylene have already been treated in the chapter on liquid and gaseous fuels.

**Benzene Series.**—Many organic compounds are known in which the carbon atoms are united, not in a straight line *chain*, as in the members of the series so far treated, but in a hexagonal *ring* formation. One such group is known as the benzene or *aromatic* series. The first in this group is *benzene*, or *benzol* ( $C_6H_6$ ). Its structural formula is shown in the diagram. Benzene is obtained from coal tar by fractional distillation. It is a colorless, volatile liquid, having a characteristic odor. Benzene is a good solvent for fats, gums, rubber, and resins. It is used in paint and varnish removers, and in the manufacture of numerous aniline dyes. (Care should be taken not to confuse *benzene*, which is a definite chemical compound, derived from soft coal, with *benzine*, which is a mixture of hydrocarbons obtained from petroleum.)



**Other Hydrocarbons.**—Other hydrocarbons of considerable commercial importance follow:

1. *Toluol* or *toluene* ( $C_7H_8$ ) is a colorless liquid made from coal tar. It is used in the preparation of drugs, dyes, perfumes, saccharine, and trinitrotoluene (T.N.T.).

2. *Naphthalene* ( $C_{10}H_8$ ) is a white, insoluble solid, having a characteristic odor. It is used in making "moth balls," and in the preparation of indigo.

3. *Turpentine* ( $C_{10}H_{16}$ ) is a liquid obtained by distilling the sap of pine trees. It is used in medicine, and in making paints and varnishes.

**Substitution Products.**—Part or all of the hydrogen in methane ( $CH_4$ ) may be replaced by chlorine, forming the following compounds: (1) monochlormethane or methyl chloride ( $CH_3Cl$ ); (2) dichlormethane ( $CH_2Cl_2$ ); (3) trichlormethane or chloroform ( $CHCl_3$ ); (4) tetrachlormethane or carbon tetrachloride ( $CCl_4$ ). Other elements or radicals, such as Br, OH,  $NO_3$ ,  $NH_2$ , etc., can be used to replace the hydrogen in methane and other hydrocarbons, thus giving rise to an enormous number of new compounds. Such derivatives are called *substitution products*. Some important substitution products of methane follow:

1. *Chloroform* ( $CHCl_3$ ) is prepared by the reaction between bleaching powder and ethyl alcohol. It is a heavy, oily, volatile

liquid with a pleasant, sweetish odor. Chloroform is used as an anaesthetic, and as a solvent for fats.

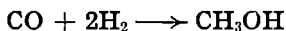
2. *Iodoform* ( $\text{CHI}_3$ ) is prepared by the reaction between iodine and ethyl alcohol in the presence of an alkali. It is a yellow crystalline solid with a penetrating odor. Iodoform is used as an anti-septic.

3. *Carbon tetrachloride* ( $\text{CCl}_4$ ) is prepared by the reaction between chlorine and carbon disulfide ( $\text{CS}_2$ ). It is a heavy, colorless liquid, having an odor resembling that of chloroform. It is a good solvent for grease, and, since it is non-inflammable, it makes a safe fluid for "dry-cleaning." It neither burns nor supports combustion, hence is useful for extinguishing fires. It has the further advantage of being a non-conductor of electricity, and may therefore be employed to fight fires in the region of electrical apparatus, where it would be dangerous to use water.

## ALCOHOLS

An alcohol is an organic compound containing one or more *hydroxyl* (OH) groups. Alcohols may be regarded as substitution products derived from hydrocarbons by replacing hydrogen atoms with hydroxyl groups.

**Methyl Alcohol.**—Methyl alcohol ( $\text{CH}_3\text{OH}$ ), commonly called *wood alcohol* or *methanol*, is related to methane because of the substitution of a hydrogen atom by an OH group. It is prepared commercially by the destructive distillation of wood. In a synthetic process that has recently been perfected, carbon monoxide and hydrogen are made to react with the help of zinc oxide as a catalyst to form wood alcohol:



Wood alcohol is a colorless liquid having a characteristic odor. It is violently poisonous, and even small quantities have been known to cause blindness and death. Wood alcohol is used as a solvent in making varnish and shellac, as a fuel in alcohol lamps, and as a *denaturing agent*, by which ethyl alcohol is made unfit for beverage purposes, while it remains suitable for use as "anti-freeze" in automobile radiators and for other industrial uses.

**Ethyl Alcohol.**—Many alcohols are known, but the term "alcohol" is most often applied to ethyl alcohol (*ethanol*,  $\text{C}_2\text{H}_5\text{OH}$ ).

It may be regarded as a substitution product of ethane, one hydrogen atom having been replaced by an OH group. Originally, most of our ethyl alcohol was derived from corn, barley, and rye; hence the name *grain alcohol*. Practically all the alcohol now produced in this country is made by fermenting molasses obtained as a by-product from the sugar refinery.

Fermentation is a chemical change brought about by the action of microscopic plants such as bacteria, yeasts, and molds. These tiny organisms form complex organic catalysts called *enzymes*. The molasses solution, containing principally glucose ( $C_6H_{12}O_6$ ), is treated with yeast, and undergoes alcoholic fermentation under the influence of *zymase*, the enzyme present in the yeast. The products of the action are carbon dioxide and alcohol:



Almost pure alcohol is obtained by fractional distillation of the resulting mixture.

Ethyl alcohol is a colorless liquid having an agreeable odor and a boiling point of  $78^\circ C$ . Like methyl alcohol, it burns with a hot, colorless flame, forming water and carbon dioxide. It is miscible with water in all proportions. In its usefulness as a solvent, it is second only to water. It is employed in the manufacture of dyes, drugs, extracts, perfumes, varnishes, and a host of important organic products. All tinctures, (*e.g.*, tincture of iodine) are alcoholic solutions.

Alcohol is intoxicating when taken internally, and, when used in excess, is definitely poisonous. It constitutes the important ingredient in a variety of fermented and distilled beverages, such as beer, wine, whiskey, gin, brandy, etc. To prevent the illegal use of alcohol for beverage purposes, the government requires that industrial alcohol be "denatured." This is accomplished by the addition of small amounts of other substances, such as methanol, benzene, and pyridine, to render the alcohol poisonous, or offensive in odor or taste. Industrial alcohol, completely denatured, is free from the restrictions and the tax imposed by the government upon ethyl alcohol intended for use in beverages.

**Glycerine.**—Glycerine (*glycerol*) is an alcohol with the formula  $C_3H_5(OH)_3$ . It may be regarded as a substitution product of propane in which three hydrogen atoms have been replaced by OH groups. Glycerine is obtained as a by-product in the manufacture

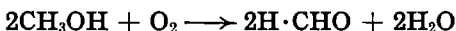
of soap. It is a thick, syrupy liquid having a sweetish taste. Glycerine finds its greatest use in the manufacture of the high explosives, *nitroglycerine* and *dynamite*. Large quantities are also used in a variety of toilet preparations, for softening leather, and as an "anti-freeze" for automobile radiators.

**Phenol.**—Phenol (carbolic acid) has the formula  $C_6H_5OH$ , and is a hydroxyl derivative of benzene. It is obtained principally from coal tar. Phenol is a colorless, crystalline solid with a characteristic odor. It is extremely poisonous. The chief uses of phenol are in the manufacture of disinfectants, explosives, and synthetic plastics, such as *bakelite*.

### ALDEHYDES

An aldehyde is an organic compound having the characteristic group  $-CHO$ . Aldehydes are formed from alcohols by partial oxidation.

*Formaldehyde* ( $H \cdot CHO$ ) is the simplest and most important aldehyde. It is prepared by passing a mixture of methyl alcohol and air over hot copper gauze:

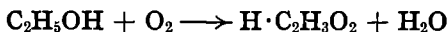


Formaldehyde is a colorless gas having a strong pungent odor. A 40% solution of the gas in water is known as *formalin*. This preparation has powerful disinfecting properties, and is used to preserve anatomical specimens. A large amount of formaldehyde is now used with phenol in the manufacture of bakelite.

### ACIDS

An organic acid is a compound containing one or more *carboxyl* ( $-COOH$ ) groups. Although the molecule of such an acid often contains many hydrogen atoms, only the hydrogen present in the carboxyl group becomes the hydrogen ion in dilute water solution, and may be replaced by a metal, forming a salt. An organic acid may be prepared by careful oxidation of the corresponding alcohol. Following are the more common organic acids:

1. *Acetic acid* ( $H \cdot C_2H_3O_2$  or  $CH_3COOH$ ) is the most important organic acid. It is obtained commercially by the destructive distillation of wood. It may also be prepared by the oxidation of ethyl alcohol in *acetic acid fermentation*:



This action is caused by the enzymes in "mother of vinegar," present in the slimy substance often seen in a vinegar bottle. Acetic acid is a colorless liquid having a strong pungent odor. Since it does not readily ionize in water, it is a weak acid. It is employed in the production of vinegar, white lead, and various organic compounds such as cellulose acetate, used in making artificial silk and photographic film.

2. *Oxalic acid* ( $\text{HOOC}-\text{COOH}$  or  $\text{H}_2\text{C}_2\text{O}_4$ ) is obtained by heating sawdust with caustic soda. It is a poisonous, crystalline solid. Being a reducing agent, oxalic acid is useful for removing iron rust and ink spots from white cloth, for bleaching flax and straw, and for cleaning brass and copper.

3. *Tartaric acid* ( $\text{H}_2\cdot\text{C}_4\text{H}_4\text{O}_6$ ) is present in grapes and other fruits in the form of the salt, potassium acid tartrate (*cream of tartar*,  $\text{KHC}_4\text{H}_4\text{O}_6$ ). The acid is a white crystalline solid, soluble in water. It is used in making soft drinks and Seidlitz powder, and in dyeing. The salt is widely used as the acidic constituent in many baking powders.

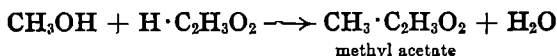
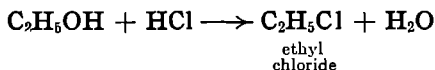
4. *Citric acid* ( $\text{H}_3\cdot\text{C}_6\text{H}_5\text{O}_7$ ) is the acid to which the citrus fruits—lemons, oranges, grapefruit, and limes—owe their characteristic sour taste. It is used in making soft drinks and citrate of magnesia.

5. *Lactic acid* ( $\text{H}\cdot\text{C}_3\text{H}_5\text{O}_3$ ) is the acid present in sour milk. When sour milk is used as a leavening agent in baking, this acid reacts with baking soda ( $\text{NaHCO}_3$ ), generating carbon dioxide.

6. *Formic acid* ( $\text{H}\cdot\text{COOH}$ ), the simplest organic acid, is made by the oxidation of formaldehyde. It is a colorless liquid with a pungent odor. Formic acid is the irritating substance present in the sting of certain insects.

## ESTERS

The reaction between a base and an acid, called *neutralization*, produces water and a salt. In a like manner, the reaction between an alcohol and an acid, called *esterification*, produces water and an *ester* or *etheral salt*.



in the process of esterification, concentrated sulfuric acid is usually employed to absorb the water formed in the reaction.

**Important Esters.**—Most esters are volatile liquids or non-crystalline solids having a pleasant, fruity odor. Following are some important members of this class of compounds:

1. *Nitroglycerine* (glyceryl trinitrate) is made by the reaction between glycerine and nitric acid:

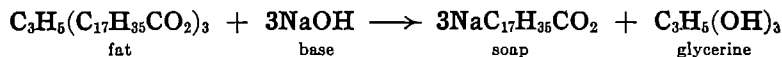


This is an oily liquid which explodes very easily and therefore must be handled with great caution. The liquid is usually soaked in some absorbent material such as sawdust, the resulting product being called "dynamite." In this form it is less sensitive to shock.

2. *Trinitrotoluene* (T.N.T.), having the formula  $\text{C}_7\text{H}_5(\text{NO}_2)_3$ , is a violent explosive made by treating toluene ( $\text{C}_7\text{H}_8$ ) with nitric acid.

3. *Fats and oils*, from both vegetable and animal sources, are esters of glycerine and fatty acids. Beef fat is largely *stearin*, or glyceryl stearate, with the formula  $\text{C}_3\text{H}_5(\text{C}_{17}\text{H}_{35}\text{CO}_2)_3$ . Vegetable oils, such as olive oil and cottonseed oil, are mainly *olein*. These oils are frequently converted into solid fats by the addition of hydrogen (*hydrogenation*). "Crisco," a lard substitute, is made from cottonseed oil in this way.

**Soap.**—A soap is a metallic salt of a fatty acid. Ordinary soap is prepared by boiling a fat with sodium hydroxide. The process is called *saponification*. The following is one of the main reactions involved:



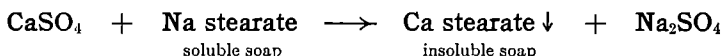
Saponification is carried out in huge kettles about fifty feet high, each kettle holding enough material to make a million pounds of soap. When the action is completed, salt is added to make a strong brine. The soap is thus rendered insoluble, and floats to the top, where it is removed.

**Kinds of Soap.**—The soap manufacturer turns out a large assortment of cleansing products which are intended for a variety of uses. *Hard soap* is the ordinary type of soap, made with sodium

hydroxide. *Soft soap* is made with potassium hydroxide. *Castile soap* is made from olive oil, and consists largely of sodium oleate. *Floating soap* contains tiny bubbles of air, which render it buoyant. *Laundry soap* usually contains a little free alkali to help it remove grease. *Soap powder* is finely ground soap mixed with washing soda ( $\text{Na}_2\text{CO}_3$ ). *Scouring powder* is soap powder mixed with some gritty material, such as fine sand.

**Cleansing Action of Soap.**—Soap in solution helps to remove dirt in two ways. (1) The soap emulsifies the grease that is present with the dirt, thus making it possible for the water to wash away both the grease and the dirt. (2) The particles of dirt are *adsorbed* on the surface of the soap bubbles, and are thus removed. Both actions result from the fact that soap forms a colloidal suspension in water (page 267 ff.).

Soap does not lather in hard water, hence such water cannot be used for laundry purposes. Hard water usually contains dissolved salts of calcium and magnesium. These salts combine with ordinary soap, forming an insoluble scum which floats on the surface. Thus, when calcium sulfate reacts with sodium stearate, insoluble calcium stearate is formed:



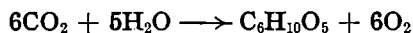
An excessively large amount of soap must therefore be used in order first to precipitate all of the calcium and magnesium ions before any cleansing action can start. There are several more practical ways of rendering hard water soft (page 214).

## CARBOHYDRATES

A carbohydrate is a compound containing carbon, hydrogen, and oxygen, the last two elements being present in the ratio of 2:1, just as in water. *Starch*, having the formula  $\text{C}_6\text{H}_{10}\text{O}_5$ , illustrates the relationship. If a carbohydrate is treated with concentrated sulfuric acid, which is a powerful dehydrating agent, the hydrogen and oxygen are removed, leaving a black residue that is chiefly carbon. Other important carbohydrates are *cellulose* and *sugar*.



**Starch.**—Starch( $C_6H_{10}O_5$ )<sub>x</sub>\* is found in the roots, seeds and leaves of plants. It is formed from carbon dioxide and water in the process of photosynthesis (page 155):



Most of our starch is obtained from corn and potatoes.

Starch is a white powder which does not dissolve in cold water, because each individual particle is surrounded by a thin coating of insoluble material. Boiling with water causes the starch grains to swell up and burst open. The resulting mixture is a colloidal suspension known as "starch paste." A drop or two of iodine solution added to starch paste produces a deep-blue color. This constitutes an extremely delicate test for either starch or iodine.

Starch is one of our most important food elements. It is used in the laundry for stiffening clothes; and is an essential raw material in the manufacture of *alcohol* (page 186), *glucose*, which is an important sugar, and *dextrin*, which is used in making adhesives.

**Cellulose.**—Cellulose is a carbohydrate having the formula ( $C_6H_{10}O_5$ )<sub>y</sub>. Plant stems, wood, and all vegetable fibers are mainly cellulose. Cotton and linen, which are nearly pure cellulose, are employed extensively in the textile industry. Paper, another important cellulose product, is manufactured from wood pulp or rags. A great variety of very valuable materials is produced as a result of the reaction of cellulose with certain chemicals.

1. *Nitrocellulose* is made by the action of concentrated nitric acid on cellulose under the dehydrating influence of concentrated sulfuric acid. By altering the conditions of the reaction, it is possible to control the degree of nitration, thus yielding a number of different cellulose nitrates. Gun-cotton, a powerful explosive, is a highly nitrated cellulose.

2. *Pyroxylin* is a less highly nitrated cellulose than gun-cotton, and finds extensive use in quick-drying lacquers for automobiles, furniture, etc.

3. *Collodion* is made by dissolving pyroxylin in a mixture of

\* The subscript  $x$  in the formula for starch represents a number whose value has not as yet been determined. The same is true of the subscript  $y$  in the formula for cellulose [ $(C_6H_{10}O_5)_y$ ]. Both of these compounds have the same basic formula, which indicates the ratio of the carbon, hydrogen, and oxygen atoms that are present. Actually, the molecule of each compound has a complex structure representing some unknown multiple of this basic formula.

alcohol and ether. This solution is the familiar "New-skin," used to cover small cuts.

4. *Celluloid* is made by mixing pyroxylin with camphor and alcohol. It is used in making toilet articles, toys, ornaments, etc.

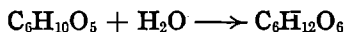
5. *Cellulose acetate* is made by the action of acetic acid on cellulose. Unlike the cellulose nitrates, it is relatively non-inflammable, and is therefore used extensively in making "safety" moving-picture film.

6. *Artificial silk*, more properly called *silk substitute*, is almost pure cellulose that has been precipitated from solutions of its compounds. The various methods employed in making substitutes for silk are taken up in detail under the study of textiles in Chapter 25.

7. *Cellophane* is a thin, highly transparent film that is made by regenerating the cellulose from a viscose solution similar to that employed in manufacturing the silk substitute known as rayon. It is used extensively for wrapping purposes, artificial flowers, and sausage casing.

**Sugars.**—The term "sugars" refers to a number of carbohydrates, the best known of which are *glucose*, *fructose*, and *sucrose*. All sugars have a sweet taste, and are soluble in water.

1. *Glucose*.—Glucose ( $C_6H_{12}O_6$ ), also known as *dextrose* or *grape sugar*, occurs in grapes and in many other fruits. It is made commercially by the hydrolysis of starch. In this process, corn starch is boiled with water in the presence of a small amount of hydrochloric acid, which acts as a catalyst:



Starch may also be changed to glucose by the action of an enzyme called *diastase*.

Glucose is used extensively in making candies, jams, table syrups, and other food products. In the process of digestion, starch and cane sugar are both changed to glucose. As explained previously, the fermentation of glucose is an important source of both alcohol and carbon dioxide:



2. *Sucrose*.—Sucrose ( $C_{12}H_{22}O_{11}$ ) is ordinary cane sugar. It occurs in the sugar cane, in certain fruits and vegetables, and in maple sap. Most of our common sugar is obtained from sugar cane. The cane is first crushed to extract the juice, which is

treated with lime and then with carbon dioxide to remove albuminous impurities. The liquid is filtered, and then evaporated, the latter process being carried on in vacuum pans to prevent scorching. The product obtained is raw sugar. In the refining process the sugar is redissolved, filtered through boneblack to remove coloring matter, evaporated under reduced pressure, and then recrystallized. The resulting product is granulated sugar.

Sucrose is a white, sweet, crystalline solid which melts at  $160^{\circ}\text{C}$ . When it is heated to  $210^{\circ}\text{C}$ ., some water is driven off, leaving *caramel*, which serves as a coloring and flavoring material in making confectionery. An enormous amount of sugar is used as a food, the annual per capita consumption in the United States being nearly 100 pounds.

When a solution of sucrose is boiled with a small amount of acid, the process of *inversion* takes place, in which the sugar combines with water, forming dextrose (*glucose*) and levulose (*fructose*).

The mixture of the two simple sugars thus formed is called *invert sugar*. Inversion may also be caused by an enzyme known as *invertase*, which is present in yeast. Since the latter substance also contains the enzyme *zymase*, which, we have already learned, causes alcoholic fermentation (page 180), it is possible to obtain alcohol and carbon dioxide from a solution of cane sugar solely by the action of yeast.

**3. Other Sugars.**—A large number of other sugars are known, many of which are of great practical importance. Among the most useful are: (a) *fructose* ( $\text{C}_6\text{H}_{12}\text{O}_6$ ), more correctly known as *levulose*, which is found in honey and in ripe fruits; (b) *lactose*, ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ), commonly known as *milk sugar*, which is found in small quantities in cow's milk, and is used in the preparation of infants' foods; and (c) *maltose* ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ), known as *malt sugar*, which is formed during the sprouting of grains, such as barley. Yeast causes maltose to ferment, forming alcohol and carbon dioxide.

## QUESTIONS

1. Why does carbon form such an enormous number of compounds?
2. What is meant by an *homologous series*? Illustrate by reference to the methane series.
3. Distinguish between *benzene* and *benzine*.
4. Give the formulas, properties, and uses of three substitution products of methane.

5. (a) What is the difference between a *hydrocarbon* and a *carbohydrate*? (b) Give two examples of each.
6. Compare methyl alcohol with ethyl alcohol as to their effects when taken internally.
7. (a) Define *fermentation*. (b) Write the equation for alcoholic fermentation. (c) What causes alcoholic fermentation?
8. (a) Why is ethyl alcohol denatured? (b) What are the properties of an ideal denaturing substance?
9. (a) Give the formula of glycerine. (b) Show how propyl alcohol ( $C_3H_7OH$ ) and glycerine are derived from the same hydrocarbon.
10. (a) Give the name and formula of the organic acid present in vinegar. (b) Name one other source of this acid.
11. Show by equations how (a) starch is converted into glucose; (b) glucose is converted into alcohol; (c) alcohol is converted into acetic acid.
12. Give the name and formula of the group which is present in all organic acids.
13. Why is stearic acid called a fatty acid?
14. (a) Name three compounds used in the laboratory in the preparation of ethyl acetate. (b) Give the purpose of each. (c) Write the equation in each case.
15. (a) State three important differences between an ester and an inorganic salt. (b) Name the corresponding process for the preparation of each.
16. (a) What is a soap? (b) Describe the manufacture of hard soap. (c) Write an equation for the reaction involved.
17. Give two reasons why soap is a good cleansing agent.
18. Give the general composition of a fat or oil.
19. (a) What happens when hard water is used for laundry purposes? (b) Write the equation for one reaction involved.
20. Show the importance of rain, fire, and sunshine in the process of starch-making in nature.
21. (a) Give two important uses of starch. (b) Describe a test for starch.
22. (a) Mention two sources of cellulose. (b) Name two manufactured articles that are mainly cellulose.

23. Why is cellulose acetate preferable to nitrocellulose in the manufacture of motion picture film?

24. (a) Mention two methods of preparing glucose. (b) Write the equations involved. (c) State two uses of glucose.

25. Describe briefly the steps involved in the extraction and refining of cane sugar.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Organic chemistry is the chemistry of the compounds of .....
2. In the structural formula of a molecule, ..... represent atoms, and ..... represent the valence bonds between atoms.
3. Acetylene is called a ..... (saturated, unsaturated) compound because its carbon atoms are joined by a ..... valence bond.
4. Phenol is derived from ..... by the replacement of a atom by a ..... group.
5. Most of our glycerine is obtained as a by-product in the manufacture of .....
6. Ethyl alcohol is useful as a fuel because .....
7. The irritating substance present in many insect stings is ....., the formula of which is .....
8. Vinegar has a sour taste owing to the presence of .....
9. Grape juice contains an acid salt of ..... The formula of this salt is .....
10. Lemons, oranges, and grapefruit contain ..... acid.
11. In the process of esterification, the alcohol acts like an inorganic .....
12. Another name for esters is .....
13. Beef fat is composed chiefly of the ester .....
14. Vegetable oils are converted into solid fats by a process called .....
15. Soft soap is a ..... salt of a ..... acid.
16. Starch is made in nature by a process called .....
17. In the hydrolysis of starch to form ....., the catalytic agent used is .....
18. The process by which glucose is converted into alcohol is called .....
19. When cane sugar is heated, it is converted into .....

20. Complete the following table by filling in the blank spaces:

NAME	FORMULA	SOURCE	USE
	$(C_6H_{10}O_5)_y$		
Glyceryl trinitrate			
	$C_6H_6$		
Hard soap			
	$CH_3OH$		
Iodoform			
	$CCl_4$		
Acetylene			
	$H \cdot C_2H_3O_2$		
Naphthalene			

### MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

- | A                      | B                             |
|------------------------|-------------------------------|
| ( ) Toluene            | 1. In soft soap               |
| ( ) Nitrocellulose     | 2. Sugar                      |
| ( ) Zymase             | 3. Denatured alcohol          |
| ( ) Acetic acid        | 4. In vinegar                 |
| ( ) Bakelite           | 5. An unsaturated hydrocarbon |
| ( ) An anti-freeze     | 6. Formaldehyde               |
| ( ) Methane            | 7. An enzyme in yeast         |
| ( ) Ethylene           | 8. From coal tar              |
| ( ) Olein              | 9. A synthetic plastic        |
| ( ) Potassium stearate | 10. In smokeless powder       |
|                        | 11. An ester                  |
|                        | 12. Marsh gas                 |

## CHAPTER 17

### METALS AND NON-METALS

**Classification of Elements.**—Thus far, we have dealt almost entirely with a class of elements known as the *non-metals*. This group constitutes only about one-fifth of the ninety-two elements which are known to be the building blocks of all matter. We now begin the study of another class of substances, the *metals*, which (outside of a few inert elements that are considered neither metals nor non-metals) comprise all the remaining elements, or nearly four-fifths of the entire family. The great bulk of the metals are extremely rare, and of no practical importance. Only about fifteen members of this group are the common metals encountered in everyday life, and these we shall study in some detail.

**Comparison of Metals and Non-Metals.**—In the chapter on the Electron Theory, we took up the differences in the structure of the metal and the non-metal atom. Fig. 61 illustrates these

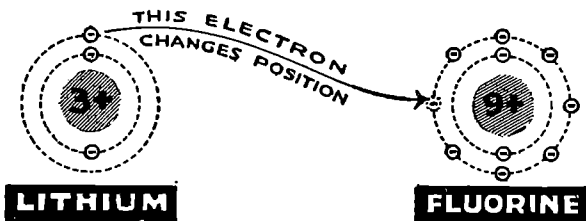


FIG. 61. *Lithium* is a typical *metal* because it has *one* electron in its outside ring which it is ready to lend to a non-metal atom. Its valence is  $+1$ . *Fluorine* is a typical *non-metal*, because it has *seven* electrons in its outside ring and is therefore "eager" to borrow another electron from a metal atom. Its valence is  $-1$ .

dissimilarities. Such variation in the number and arrangement of electrons naturally results in marked differences in properties. The table on page 192 summarizes the principal facts relating to these two classes of elements.

METALS	NON-METALS
<p>Are base formers. Have a metallic luster. Are crystalline in structure.</p> <p>Conduct heat and electricity very well. Are malleable and ductile. Unite with oxygen to form oxides which are <i>basic anhydrides</i>. Combine with non-metals to form salts. <i>Lose or lend</i> electrons during chemical reactions. Form <i>positive ions</i> in solution. Are liberated at the <i>cathode</i> during electrolysis. About 70 of the elements are metals, all of which are solids, except mercury, which is a liquid. <i>Examples:</i> Sodium, calcium, iron, copper, gold.</p>	<p>Are acid formers. Have no metallic luster. Solids may be crystalline or amorphous. Do not easily conduct heat and electricity. Solids are brittle and are not ductile. Unite with oxygen to form oxides which are <i>acid anhydrides</i>. Combine with metals to form salts.</p> <p><i>Gain or borrow</i> electrons during chemical reactions. Form <i>negative ions</i> in solution. Are liberated at the <i>anode</i> during electrolysis. 20 of the elements are non-metals, of which 10 are gases, 9 are solids, and 1, bromine, is a liquid. <i>Examples:</i> Oxygen, chlorine, nitrogen, sulfur, carbon.</p>

We have already seen that some elements, notably carbon, nitrogen and sulfur, may behave either as metals or as non-metals. (Such elements are said to be *amphoteric*—page 58.) The “noble” gases of the atmosphere, namely helium, argon, neon, krypton, and xenon, do not combine with any other element, and therefore cannot be regarded as either metals or non-metals.

**Occurrence of Metals.**—Very inactive metals like gold and platinum exist in the free or native state. Most metals, however, occur in minerals, chemically combined with other elements. A *mineral* is a solid inorganic substance found in the earth. An *ore* is a mineral from which an element, usually a metal, is extracted. For example, the mineral *hematite* is principally iron oxide ( $\text{Fe}_2\text{O}_3$ ), and is the chief ore of iron. The most common ores are oxides, carbonates and sulfides.

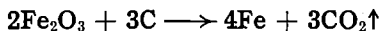
**Extraction of Metals.**—The process of extracting a metal from its ore is called *metallurgy*. There are two general methods commonly employed:

1. *Electrolysis.*—The more active metals, such as sodium,

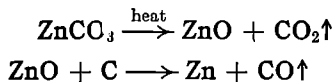


magnesium, and aluminum, are obtained by the electrolysis of their melted chlorides, hydroxides, or oxides.

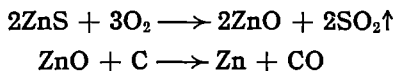
2. *Reduction by Carbon.*—This method is employed for the less active metals, such as iron, zinc, lead, tin, and copper. (a) If the ore is an *oxide*, it is heated with carbon, which removes the oxygen, leaving the metal in the free state:



(b) If the ore is a *carbonate*, it is first heated to change it to an oxide, which is in turn reduced with carbon:



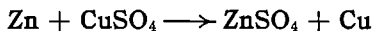
(c) If the ore is a *sulfide*, it is first changed to an oxide by roasting, that is, by heating it in contact with air. The oxide is then reduced in the usual way:



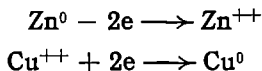
NOTE.—Oxides are sometimes reduced with aluminum (see *Thermit*, page 236).

**Relative Activity of Metals.**—The common metals are found to vary greatly in their chemical activity. The table on page 194 shows the order of activity, and indicates how this activity is reflected in certain outstanding characteristics of these elements.

**Replacement of Metals.**—When a strip of zinc is dipped into a copper sulfate solution, it becomes coated with copper, while some zinc enters the solution in the form of zinc sulfate:



The electronic equations for the reaction may be written:



Since zinc is more active than copper, being higher up on the activity list, it has a greater tendency to lose electrons. Hence, two electrons leave the neutral zinc atom, which becomes a zinc ion, and go over to the positively charged copper ion, which becomes a neutral copper atom, depositing as metallic copper. The

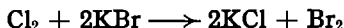
METALS	ORDER OF ACTIVITY	OCCUR-RENCE	MOST COMMON ORE	EASE OF REDUCTION	DISSOLVED BY
1. Potassium	<i>Highest</i> activity, or <i>greatest</i> tendency to lose electrons, decreasing to ↓ <i>lowest</i> activity, or <i>least</i> tendency to lose electrons	Only in combined form	KCl	Difficult (The oxides of these metals cannot be reduced by carbon.)  Moderate (The oxides of these metals are reduced commercially by carbon.)	Cold water
2. Sodium			NaCl		
3. Calcium			CaCO <sub>3</sub>		
4. Magnesium			MgCl <sub>2</sub>		
5. Aluminum			Al <sub>2</sub> O <sub>3</sub>		
6. Zinc			ZnO		
7. Iron			Fe <sub>2</sub> O <sub>3</sub>		
8. Nickel			NiS		
9. Tin			SnO <sub>2</sub>		
10. Lead			PbS		
11. HYDROGEN					
12. Copper	Free and combined	Cu <sub>2</sub> S	Easy  Very easy	HNO <sub>3</sub> or hot H <sub>2</sub> SO <sub>4</sub>	
13. Mercury		HgS			
14. Silver		Ag			
15. Platinum	Free only	Pt	Need no reduction	Aqua regia	
16. Gold		Au			

sulfate ion does not undergo any change. In a similar fashion, copper will replace silver in a silver nitrate solution, neutral silver atoms depositing out of solution, while positively charged copper ions go into solution. Therefore, the activity series of metals is also known as the *replacement series* or *electrochemical series*. Each metal in the list will replace from its salt solution any metal occurring below it, but will, in turn, be replaced by any metal above it.

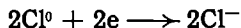
Hydrogen is included in the replacement series of the metals because, although a non-metal, it forms positive ions in solution, and acts as a metal chemically. The metals *above* hydrogen replace it from dilute acids, while the metals *below* hydrogen do not replace it from acids.

**Replacement of Non-Metals.**—Certain of the non-metals also can be arranged in the order of their ability to replace each other in solution. The halogens are a good example of such a series. When

chlorine is added to a potassium bromide solution, free bromine is liberated and potassium chloride is formed:



The electronic equations may be written as follows:



Since chlorine is a more active non-metal than bromine, it has a greater tendency to gain electrons. Hence, each atom takes one electron from a bromine ion, thus converting itself into a chlorine ion. The bromine ion, meanwhile, becomes a neutral atom, and unites with another bromine atom, forming a free bromine molecule, which is liberated.

**Alloys.**—An alloy is a uniform mixture of two or more metals. The ingredients of an alloy are usually mixed while in the molten condition, and the mixture is allowed to harden. The nature of the alloy depends on its physical structure, as seen under the microscope. Three classes of alloys are known:

1. The metals may exist in the form of a mixture of crystals, as in the case of a gold-silver alloy. This may be called a *solid solution*.

2. One or more of the metals may *crystallize out* from a saturated solution, as in the case of some copper and silver alloys.

3. The metals may form a definite *chemical compound* with each other, as in the case of the alloy "sodamalgam," which has the formula  $\text{NaHg}_2$ . Other compounds of this type have the formulas  $\text{Cu}_3\text{Al}$ ,  $\text{Fe}_3\text{C}$ , and  $\text{Cu}_2\text{Zn}_5$ .

**Properties of Alloys.**—As a rule, alloys differ radically in properties from their constituent metals. Some of these properties are:

1. *Hardness.*—An alloy is usually harder than any of its ingredients. Thus, pure gold is relatively soft, but the addition of copper makes it sufficiently hard and durable for use in jewelry and coins.

2. *Melting Point.*—The melting point of an alloy is lower than that of one of its components, and sometimes lower than the melting points of all the components. Thus, brass, made of copper and zinc, has a melting point between that of copper and zinc; while solder, made of lead and tin, melts at a lower temperature than either of these two metals.

3. *Chemical Activity*.—Usually, the alloy is far less active than its constituents. Thus, sodium amalgam (sodium and mercury) decomposes water very slowly, whereas pure sodium reacts vigorously with water. This property of alloys is made use of in protecting a metal from corrosion. For example, stainless steel is made by adding a small amount of chromium to the steel to form an alloy.

4. *Electrical Conductivity*.—An alloy is not as good a conductor of electricity as any of its constituents, which means that its *resistance* is higher. This property is made use of in the electric toaster, in which the heating element is nichrome, an alloy of nickel, chromium and iron.

5. *Color*.—Various combinations of metals are employed to produce alloys with distinctive colors. Thus the addition of silver, cadmium, and even iron to gold yields the white, green, and blue golds used in jewelry.

NOTE.—Individual alloys are taken up in greater detail in the discussions of the respective metals. For a complete list of alloys, their composition and their uses, see the table in the Appendix.

**Weathering of Metals**.—With the exception of the so-called "royal metals," gold, platinum, etc., all metals tarnish or rust, owing to the slow action of oxygen, water vapor, carbon dioxide, and other substances present in the air. When the forces of decay penetrate into the interior of the metal, the process is called *corrosion*. This is best illustrated by the rusting of iron, in which the rust first formed appears to act catalytically in hastening the ultimate disintegration of the metal, unless checked by artificial means. Since iron is our most important metal, this rusting action results in a loss estimated at many millions of dollars each year.

**Prevention of Rust**.—Certain metals are attacked by the air, as a result of which a layer of compound is formed on the surface that acts like a coat of paint in preventing further action by the air. Aluminum, lead, copper, and zinc are examples of such self-protecting metals. Other metals must be treated in various ways to obtain the same result. The following processes for protecting iron from rusting illustrate the typical methods employed:

1. *Covering the Metal with a Non-Metallic Coating*.—This is done by painting, oiling, enameling or covering with a thin layer of the oxide of the metal (e.g.,  $\text{Fe}_3\text{O}_4$  on *Russia iron*).

2. *Covering the Metal with a Layer of a More Resistant Metal.*—This is accomplished by coating with zinc (*galvanizing*), coating with tin (*tinning*), and *electroplating* with copper, nickel or chromium.

3. *Alloying with Other Metals.*—Steel is rendered “stainless” by being alloyed with chromium. Silicon is added to iron to make non-rusting *duriron*.

### QUESTIONS

1. (a) Which of the following elements are considered metals: zinc, aluminum, sulfur, bromine, potassium, carbon, oxygen, hydrogen, neon, mercury? (b) What general properties are the basis of this classification?

2. (a) From the list of elements in question 1, select two which may act either as metals or as non-metals. (b) What name is applied to this property? (c) How is it explained in terms of the Electron Theory?

3. (a) Name three metals that are found free in nature, and three that are found combined. (b) Compare the chemical activity of the first group with that of the second.

4. Name three physical properties and three chemical properties characteristic of non-metals.

5. (a) Define *metallurgy*. (b) Name two general methods used in metallurgy, and illustrate each with an equation.

6. Arrange the following metals in the order of their replacing power: platinum, calcium, mercury, zinc, silver, copper, magnesium.

7. (a) Why is hydrogen included in the electrochemical series? (b) Name three metals that react with acids to yield hydrogen, and three metals that do not.

8. (a) According to the Electron Theory, why does chlorine replace iodine in a solution of potassium iodide? (b) Write the electronic equation for this replacement.

9. (a) Define *alloy*. (b) Compare four properties of alloys with the corresponding properties of the component metals.

10. (a) State five practical methods by which the rusting of iron may be prevented. (b) How is the result obtained in each case?

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. A non-metal is a ..... of electrons, forms ..... ions in solution, and unites with oxygen to form ..... anhydride.
2. Iron is obtained from its oxide ore by removing the .....; in this process ..... is used as a ..... agent.
3. To obtain a metal from its sulfide ore, the latter is first converted into the ....., which is then reduced with .....
4. Metals near the ..... (top, bottom) of the replacement series do not corrode because of their (higher, lower) ..... activity.
5. Bromine can replace ....., but cannot replace ..... from compounds in solution.
6. A metal which lends electrons more easily than does zinc is .....
7. The melting point of an alloy is always ..... (higher than, lower than, the same as) that of one of its components.
8. Gold is alloyed with copper to give it the property of .....
9. Iron rusts more easily than other common metals because .....
10. Three metals which protect themselves from corrosion are .....

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A  | B                            |
|--|------------------------------|
| ( ) Nitrogen                             | 1. Sodamalgam                |
| ( ) Non-metal having metallic properties | 2. Chief ore of iron         |
| ( ) Made with mercury                    | 3. Noble gas                 |
| ( ) Replaces iodine in solution          | 4. Replaces gold in solution |
| ( ) Solder                               | 5. Acid former               |
| ( ) Hematite                             | 6. Galvanizing               |
| ( ) Copper                               | 7. Electrolysis              |
| ( ) Is roasted                           | 8. Hydrogen                  |
| ( ) To protect iron from rusting         | 9. An alloy                  |
| ( ) A process in metallurgy              | 10. Zinc sulfide ore         |
|  | 11. Chlorine                 |
|  | 12. Acid anhydride           |

## CHAPTER 18

### SODIUM, POTASSIUM, AND THEIR COMPOUNDS

#### ALKALI METALS

**Importance.**—The chemical family known as the *alkali metals* heads the list of metallic elements. Lithium, sodium, and potassium are the most important members of the group. (Two rarer alkali metals are rubidium and cesium.) From the electronic structure of their atoms (Fig. 62), it can be seen that each has a single electron in its outer ring, which it can easily lend to some other atom. Therefore, these elements have a valence of 1, exhibit distinctly metallic properties, and are exceedingly active.

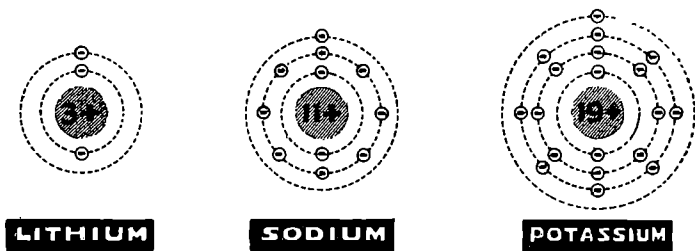


FIG. 62. Atomic structure of three alkali metals.

Although the alkali metals have a limited use (chiefly in the preparation of certain compounds), they are particularly interesting in that they contradict the popular conception of metals. Ordinarily, a metal is regarded as a substance that is hard, heavy, and rather difficult to melt. Yet sodium and potassium are soft enough to be molded with but slight pressure, are light enough to float on water, and melt below  $100^{\circ}\text{C}$ .

These two alkali metals are important chiefly for the large number of stable compounds which they form. Some of them, like sodium chloride and potassium chloride, are found abundantly in nature. Others, like sodium hydroxide and sodium carbonate, are manufactured on a large scale. These compounds of the alkali metals

are of great importance in fundamental industries and industrial processes.

**Occurrence.**—Sodium and potassium, because of their extreme activity, are never found free in nature. Combined with other elements, they occur in many compounds that are widely distributed. Some of the more important sodium compounds are *sodium chloride* (common table salt), *sodium nitrate* (Chile saltpeter), and *sodium tetraborate* (borax). The principal potassium compounds found in nature are *potassium chloride* (in the mineral carnallite,  $\text{MgCl}_2 \cdot \text{KCl} \cdot 6\text{H}_2\text{O}$ ), *potassium carbonate* (potash, found in the ashes of plants), and *potassium aluminum silicate* (feldspar).

**Preparation of Sodium and Potassium.**—Sodium and potassium are prepared today by essentially the same process as was used in 1807 by the English chemist, Davy, who first succeeded in isolating the two elements by the electrolysis of their fused (melted) hydroxides. An iron tank containing an iron anode and a graphite cathode is used. The equation for the preparation of sodium is:

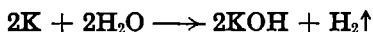
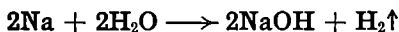


The sodium and hydrogen are liberated at the cathode, while the oxygen is released at the anode. The apparatus is so designed that the two gases are kept from mixing.

**Properties of Sodium and Potassium.**—These two alkali elements are strikingly similar to each other, both in appearance and in general behavior.

1. *Physical Properties.*—Sodium and potassium are soft enough to be cut with a knife. The freshly cut surface has a characteristic metallic luster, which soon darkens on exposure to the air, owing to the formation of the oxide. Both metals are excellent conductors of heat and electricity.

2. *Chemical Properties.*—Sodium and potassium combine so readily with the water vapor and the oxygen of the air that they must be kept immersed in kerosene. Both metals react violently with water, forming hydrogen and the hydroxide of the metal:



So much heat is evolved by the action that the metal melts into a globule which moves around rapidly on the surface of the water.



In the case of potassium, the heat generated is sufficient to set the hydrogen on fire.

**Uses.**—The free alkali metals have only a limited use. Sodium is employed in the manufacture of the oxidizing agent, sodium peroxide ( $\text{Na}_2\text{O}_2$ ), and of sodium cyanide ( $\text{NaCN}$ ), used to extract gold from its ores. It is employed as a reducing agent in the preparation of certain drugs and dyes, and as a catalytic agent in making synthetic rubber. Potassium, in the form of a thin film, is used in one type of photo-electric cell, or "electric eye." Light striking the metal causes a stream of electrons to be emitted, the strength of this flow varying with the intensity of the illumination. The variations in light are thus translated into a fluctuating electric current.

**Flame Tests.**—When certain metals or their compounds are heated and vaporized in the Bunsen burner flame, they impart a characteristic color to the flame. The colors produced are:

Sodium	<i>yellow</i>	Calcium	<i>orange-red</i>
Potassium	<i>violet</i>	Barium	<i>green</i>
Lithium	<i>crimson</i>	Strontium	<i>bright-red</i>

The test is made by dipping a clean wire made of platinum or nichrome into the substance being examined (either solid or water solution), and then holding it in the outer edge of the colorless flame. A blue cobalt glass is used in the test for a potassium compound in the presence of a sodium compound. The blue glass absorbs the dominant yellow sodium light, but allows the violet light of the vaporized potassium compound to pass through.

## COMPOUNDS OF SODIUM AND POTASSIUM

### SODIUM CHLORIDE

**Occurrence.**—Sodium chloride ( $\text{NaCl}$ ), ordinary table salt, is one of the most abundant and valuable compounds known to man. It is found in ocean water to the extent of about 3%, and in an even greater concentration in salt lakes, such as the Great Salt Lake and the Dead Sea. Beds of solid rock salt (*halite*) are found scattered in many localities throughout the world, chiefly in the United States, Germany, and Austria.

**Methods of Extraction.**—Salt is obtained from salt seas and the ocean by running the water into shallow basins and allowing it to

evaporate by the sun's heat. In cold regions, as in northern Russia, the water is frozen out, leaving a strong brine, which is then heated until the remaining water evaporates. Salt is extracted from salt-beds by a process very similar to coal-mining. Another method (used in New York State) consists of driving pipes into the salt beds, dissolving the salt with water, pumping out the salt solution, and then evaporating the water.

**Methods of Purification.**—Nearly pure salt is obtained commercially by a process called *fractional crystallization*. The various salts that occur mixed with sodium chloride are soluble in water to different degrees. When natural brine is evaporated, those compounds that are less soluble than sodium chloride are the first to crystallize out. The concentrated brine is then run into other vessels, in which the process of evaporation continues. The next substance to crystallize out of solution is the sodium chloride. The remaining liquid, containing all the more soluble substances, is drained away, leaving nearly pure salt.

Frequently, pure salt is obtained from crude salt solutions by bubbling hydrogen chloride gas through a saturated solution of the impure salt. Pure sodium chloride crystallizes out under these conditions.

**Properties.**—Sodium chloride is a white solid made up of cubical crystals. These contain no water of crystallization, but usually have water that is mechanically enclosed. The vaporization of the enclosed water causes the crystals to crackle and explode on being heated. This action is called *decrepitation*. Sodium chloride has the taste which is characteristic of most salts. It is unusual in that it is about equally soluble in cold and hot water. Pure salt is not deliquescent; common table salt gathers moisture from the air because it contains impurities, such as magnesium chloride.

**Uses.**—In small quantities sodium chloride is our most important mineral food. It is the source of the hydrochloric acid which is present in the human stomach and which is an essential factor in the process of digestion. Sodium chloride is used in great quantities in seasoning foods, and is employed as a preservative in packing meats and fish. Salt and ice form freezing mixtures, used in the transportation of perishable foods. Sodium chloride is the starting point for the manufacture of practically all of the very important sodium and chlorine compounds used today, from sodium hydroxide, sodium carbonate and sodium bicarbonate down to hydrochloric acid and even free chlorine.

## SODIUM HYDROXIDE

**Method of Preparation.**—Sodium hydroxide ( $\text{NaOH}$ ), a typical base, is manufactured in large quantities, and has a variety of uses. It is made commercially in two ways.

1. *Nelson Process (Vorce Process).*—When an electric current is passed through a strong brine solution, the three products formed are hydrogen, chlorine, and sodium hydroxide. This is the chief

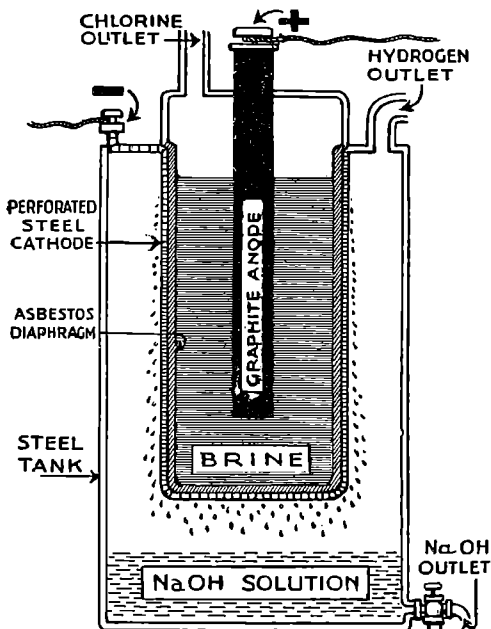
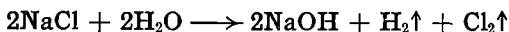


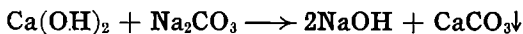
FIG. 63. Vorce cell for the electrolysis of sodium chloride.

commercial method for preparing chlorine (page 88) as well as sodium hydroxide. The Nelson (or Vorce) cell is the apparatus now in common use for the electrolysis of sodium chloride (Fig. 63). The cathode is a perforated steel trough lined with asbestos. The anode is a rod of graphite suspended in the center. Chlorine is liberated at the anode while sodium hydroxide is formed at the cathode. The asbestos diaphragm keeps these two products from mixing and thus reacting with each other. Hydrogen is also

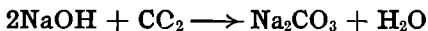
liberated at the cathode and escapes through a separate outlet. The equation for the reaction is:



2. *Soda-Lime Process.*—This is an older method, and consists of boiling solutions of sodium carbonate and calcium hydroxide. The calcium carbonate formed is insoluble, and is filtered out, leaving the sodium hydroxide:



**Properties.**—Sodium hydroxide is a white solid which is very soluble in water. It is highly corrosive in its action on the skin, hence its common name, *caustic soda*. Sodium hydroxide is extremely deliquescent, absorbing large quantities of moisture from the air and dissolving in the water thus absorbed. Its water solution has a bitter taste and a soapy feel, turns red litmus blue, and neutralizes acids—thus exhibiting all the characteristic properties of a typical strong base. Sodium hydroxide dissolves animal fibers, like wool and silk, but has little effect on vegetable fibers, like cotton and linen. Therefore, it can be used to identify various textiles. It absorbs carbon dioxide from the air, forming sodium carbonate:



**Uses.**—Sodium hydroxide finds its greatest industrial use in the manufacture of soap. It is also employed in refining petroleum, in making dyes, in mercerizing cotton, in making substitute silk (rayon) by the “viscose” process, and in preparing other sodium compounds. *Lye* is sodium hydroxide put up for household use. A strong solution of lye readily dissolves grease and is, therefore, a good cleansing agent for sinks and enamelware.

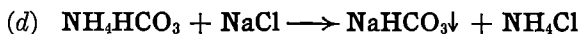
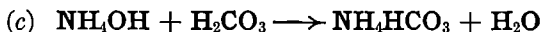
### SODIUM CARBONATE

**Method of Preparation.**—Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) is now made chiefly by the *Solvay Process*. This process is an excellent example of the remarkable efficiency of modern chemical methods. Since the raw materials are abundant, and the by-products are used over again in the preparation of the main product, the operation can be carried out with great economy. The steps in the Solvay Process are:

1. Limestone ( $\text{CaCO}_3$ ) is heated, forming calcium oxide and liberating carbon dioxide:



2. The carbon dioxide under pressure is passed through a saturated water solution of sodium chloride and ammonia. The reactions which take place may be expressed as follows:



The four reactions may be summarized as follows:

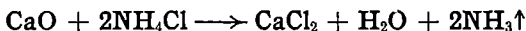


3. By an accurate control of the temperature, the sodium bicarbonate, which is nearly insoluble in the solution of ammonium chloride, is precipitated, and is separated from the liquid by filtering.

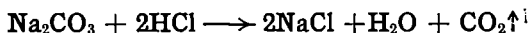
4. By heating the bicarbonate, half of its carbon dioxide is driven off, leaving sodium carbonate:



5. The ammonia is recovered from the ammonium chloride by use of the quicklime obtained when limestone was heated in the first step of the process:



**Properties and Uses.**—Sodium carbonate, when obtained from the slow evaporation of its water solution, occurs in the form of large transparent crystals, with the formula  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . This variety of the substance is common washing soda. On exposure to the air, the crystals gradually lose their water of crystallization (efflorescence) and crumble into a white powder. This dry, or anhydrous, sodium carbonate is called *soda ash*. Sodium carbonate in solution gives a basic reaction, owing to hydrolysis. When the substance is treated with an acid, carbon dioxide is liberated:



Sodium carbonate is used in making glass, soap and soap powders. It is also employed for softening water, refining petroleum, and for cleaning and dyeing.

### SODIUM BICARBONATE

**Method of Preparation.**—Sodium bicarbonate (baking soda,  $\text{NaHCO}_3$ ) is made in the Solvay Process, as explained above. It can be prepared in pure form by bubbling carbon dioxide into a saturated sodium carbonate solution:



**Properties and Uses.**—Sodium bicarbonate is a white solid which dissolves in water, forming a solution having mild basic properties (hydrolysis). When sodium bicarbonate is treated with an acid, carbon dioxide is liberated:

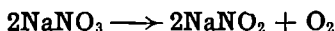


Sodium bicarbonate is the source of the carbon dioxide used in raising cake dough and some bread dough (see *Baking Powders*, page 155). It is employed to generate carbon dioxide in the soda-acid fire extinguisher (page 156). As bicarbonate of soda, it is used in medicine, particularly to counteract excessive acidity in the stomach.

### SODIUM NITRATE

**Occurrence and Method of Preparation.**—Sodium nitrate ( $\text{NaNO}_3$ ), known as *Chile saltpeter*, occurs in large deposits in Chile and Peru. These deposits were formed by the decay of organic matter in the presence of dry air and sodium salts. The crude mineral (known as *caliche*) contains about 60% sodium nitrate. It is dissolved in water, the insoluble impurities are allowed to settle, and the clear liquid is evaporated until crystallization is complete.

**Properties and Uses.**—Sodium nitrate is a white, crystalline, hygroscopic solid, very soluble in water. When heated, it liberates oxygen, forming sodium nitrite:



Great quantities of sodium nitrate are used in the manufacture of fertilizer, nitric acid, and potassium nitrate. In recent years, the

advances made in fixing atmospheric nitrogen (page 123) have caused a partial displacement of Chile saltpeter as the chief source of the world's nitrates.

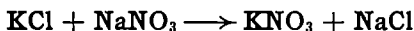
### POTASSIUM COMPOUNDS

The compounds of potassium are very similar to the corresponding sodium compounds, although they are somewhat more active. They do not occur as abundantly as the sodium compounds, and therefore are more expensive. Since potassium is essential to plant life, the compounds of this element find their chief use in the making of fertilizers. Following are the more important potassium compounds:

1. *Potassium chloride* (KCl) occurs in the minerals carnallite (KCl·MgCl<sub>2</sub>·6H<sub>2</sub>O), and sylvite (KCl), two of the potash salts obtained from the famous mines at Stassfurt, Germany. Potassium chloride is used as a fertilizer and in the preparation of other potassium compounds.

2. *Potassium hydroxide* (KOH) is commonly known as caustic potash. It is prepared by the electrolysis of potassium chloride in water solution. It is employed as the electrolyte in the Edison storage battery and also enters into the manufacture of liquid soaps.

3. *Potassium nitrate* (KNO<sub>3</sub>), is known as saltpeter. It is found in India and Persia, but most of the supply is prepared by mixing hot, concentrated solutions of potassium chloride and sodium nitrate:



Of the four salts present in this mixture, the least soluble in hot water is sodium chloride, and it therefore crystallizes out first. After the sodium chloride has been filtered off, the potassium nitrate is obtained by crystallizing the solution that remains. Potassium nitrate enters into the manufacture of fireworks and gunpowder. Other important uses are as a fertilizer, and as a preservative for meats, such as corned beef.

4. *Potassium chlorate* (KClO<sub>3</sub>) is a strong oxidizing agent. It is used in the laboratory preparation of oxygen, and industrially in the manufacture of matches, explosives and fireworks.

## QUESTIONS

1. State two ways in which sodium and potassium are alike, and two ways in which they are different.
2. Explain why sodium and potassium do not occur free in nature.
3. (a) Why must sodium and potassium be kept under kerosene?  
(b) Write two equations for the reactions which take place when sodium is allowed to remain in contact with the air.
4. Name three important industries which use salt as a raw material.
5. (a) Is sodium chloride deliquescent? (b) Explain why table salt becomes moist in damp weather.
6. Cold water is used to dissolve the salt in salt wells. Would it be better to use hot water? Explain.
7. With the aid of a labeled diagram, explain the operation of the Nelson cell used in the manufacture of sodium hydroxide.
8. Describe the method of preparing sodium hydroxide, starting with sodium carbonate.
9. When a water solution of sodium hydroxide is electrolyzed, sodium is not one of the products. Explain.
10. Write the equations for the reactions involved in the Solvay process for making sodium carbonate.
11. Woolen garments should not be washed in solutions of sodium carbonate. Explain the reason for this.
12. Describe how you would distinguish between sodium carbonate and sodium bicarbonate by chemical test.
13. Which gives a stronger basic reaction in solution, sodium carbonate or sodium bicarbonate? Explain with the aid of ionic equations.
14. Why is sodium bicarbonate an important ingredient in all baking powders?
15. Give the chemical name and formula of each of the following: caustic soda, soda ash, saltpeter, Chile saltpeter.

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Potassium is classed as a metal because it has the following three properties: . . . . .
2. Cobalt glass is used in the flame test to identify a . . . . . compound in the presence of a . . . . . compound.
3. Salt is the source of the compound . . . . . which is needed for digestion in the human stomach.



## SODIUM, POTASSIUM, AND THEIR COMPOUNDS 209

4. The common name for sodium hydroxide is .....; the kind used in the household is called .....

5. In buying sodium carbonate it is ..... (more, less) economical to get the powdered rather than the crystalline variety because .....

6. Two important uses of sodium bicarbonate are .....

7. A leading source of potassium is the compound ....., found in the mineral .....

8. Potassium carbonate gives a ..... reaction in solution. This is because it takes part in a process called .....

9. The oxidizing agent used in ordinary black gunpowder is .....

10. *Fill in the blanks in the following table:*

CHEMICAL NAME	COMMON NAME	FORMULA	USE
	Washing soda		
		$\text{NaHCO}_3$	
Sodium chloride			
	Caustic potash		
			To fix photographic film

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A                          | B                                    |
|----------------------------|--------------------------------------|
| ( ) Sodium nitrate         | 1. Nelson cell                       |
| ( ) Halite                 | 2. Crimson                           |
| ( ) Li, Na, K, Rb, Cs      | 3. Chile saltpeter                   |
| ( ) Impurity in table salt | 4. Used in soap making               |
| ( ) Electrolysis of salt   | 5. Soda ash                          |
| ( ) Calcium chloride       | 6. Alkali metals                     |
| ( ) Sodium hydroxide       | 7. A use for metallic potassium      |
| ( ) Photo-electric cell    | 8. Magnesium chloride                |
| ( ) Flame test for lithium | 9. Property of salt crystals         |
| ( ) Decrepitation          | 10. By-product of the Solvay process |
|                            | 11. Heating limestone                |
|                            | 12. Crude sodium chloride            |

## CHAPTER 19

### CALCIUM AND ITS COMPOUNDS

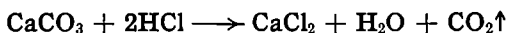
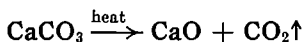
**Importance.**—Calcium, the principal member of the *alkali earth* group of metals, has only a limited commercial use as a free element. Its compounds, however, are of vast importance. They include such well-known minerals as marble and limestone, the latter being the basic raw material in the preparation of many widely used building materials, such as plaster, lime, mortar, and cement. In addition, calcium is an important ingredient in plant and animal foods. Bones and teeth contain a high percentage of this element in the form of calcium phosphate. Hence, it is essential that the diet, particularly that of a growing child, provide for an adequate amount of calcium.

**Occurrence, Preparation, and Properties of Calcium.**—Next to aluminum and iron, calcium is the most abundant metal, constituting 3.4% of the earth's crust. Since it is too active to exist as a free element, it occurs in a variety of compounds. The most important and abundant calcium salts are *calcium carbonate*, *calcium sulfate*, *calcium phosphate*, and *calcium fluoride*.

Calcium is prepared by the electrolysis of the fused chloride. It is a light, silver-white metal which tarnishes rapidly in moist air. It exhibits typically metallic properties; for example, it combines readily with oxygen, forming calcium oxide; and reacts with water, forming calcium hydroxide and liberating hydrogen.

**Calcium Carbonate.**—Calcium carbonate ( $\text{CaCO}_3$ ) is the most abundant compound of calcium found in nature. The pure crystalline form is called *calcite*. Other less pure forms are *marble*, *limestone*, and *chalk*. *Dolomite* is a common mineral, having the formula  $\text{CaCO}_3 \cdot \text{MgCO}_3$ .

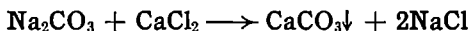
1. **Properties.**—All forms of calcium carbonate are white, amorphous solids, except calcite, which is crystalline and transparent. Calcium carbonate is easily decomposed by heat, and is readily acted on by acids:



It is insoluble in pure water, but dissolves slowly in water containing carbon dioxide in solution, forming soluble calcium bicarbonate:



2. *Uses.*—Natural calcium carbonate (marble and limestone) is used extensively for building and sculpture. Limestone is employed in the manufacture of glass and cement and in the production of lime. It is used as a flux in the reduction of iron ore in the blast furnace (page 221). In a finely ground form, it performs an important function by neutralizing the acid condition of certain soils. It provides a ready means of preparing carbon dioxide, both in the laboratory and commercially (page 154). Pure calcium carbonate, sometimes called *precipitated chalk*, is a finely divided powder used in tooth powders, tooth paste, polishes and putty. It is made by mixing solutions of sodium carbonate with a calcium salt:

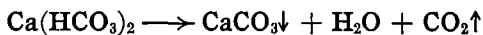


**Limestone Caves.**—In certain regions where large deposits of limestone occur, the slow solvent action of soil water containing carbon dioxide in solution has resulted in the formation of extensive caves. Some fine examples of such underground phenomena are the Mammoth Cave of Kentucky, the Luray Caverns of Virginia, and the Carlsbad Caverns of New Mexico.

The action which produces a cave involves the formation of soluble calcium bicarbonate:

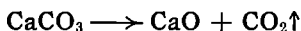


This action is readily reversible. As water containing calcium bicarbonate in solution drips from the roof of the cave, some water evaporates, and the compound is decomposed:



This results in the formation of projections called *stalactites*, which hang like icicles from the roof of the cave. In addition, some water may drop to the floor, and, because of further evaporation and decomposition, form corresponding upward projections called *stalagmites*. Eventually stalactites and stalagmites meet to form columns, and in time the cave is completely filled up again.

**Calcium Oxide.**—Calcium oxide ( $\text{CaO}$ , *quicklime*) is prepared by heating crushed limestone in a special furnace or *kiln* (Fig. 64):



Since this is a reversible reaction, complete decomposition of the limestone is attained only if the carbon dioxide formed is immediately removed. To effect the removal of this gas, a strong blast of hot air or steam is forced in at the bottom of the furnace.

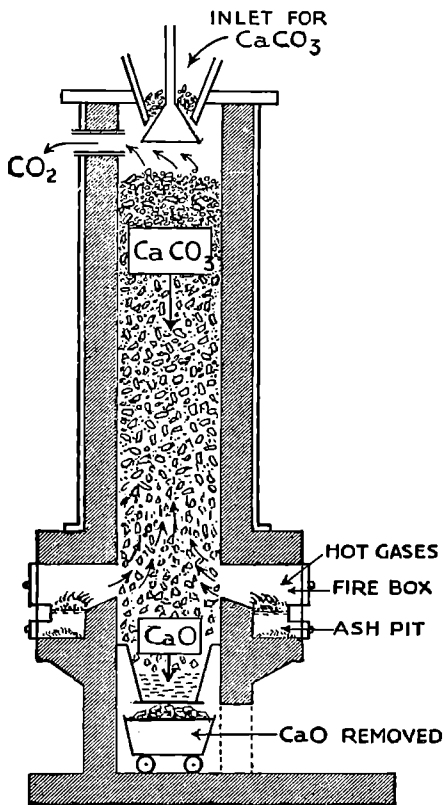
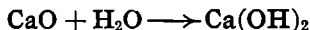
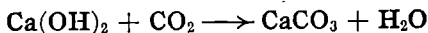
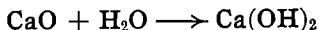


FIG. 64. A vertical lime kiln.

Calcium oxide is a white solid which melts only at extremely high temperatures. It reacts rapidly with water, forming calcium hydroxide (*slaked lime*):



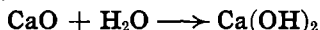
This "slaking of lime" produces enough heat to convert some of the water to steam, thus causing a vigorous bubbling in the mixture. When quicklime is exposed to the air, it gradually absorbs moisture and carbon dioxide, forming calcium carbonate:



This product is commonly known as *air-slaked lime*.

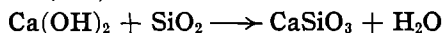
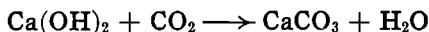
Calcium oxide is employed in the preparation of other calcium compounds such as calcium carbide (page 150) and bleaching powder. Its chief use is in the production of calcium hydroxide.

**Calcium Hydroxide.**—Calcium hydroxide,  $\text{Ca}(\text{OH})_2$ , (*slaked lime*) is prepared by the reaction between calcium oxide and water, as explained above:



Calcium hydroxide is a white powder, slightly soluble in water. Because it is a rather strong base and is comparatively cheap, it is used industrially for a variety of purposes, such as softening water, removing hair from hides, and making glass, sodium hydroxide, bleaching powder, mortar, and plaster. *Limewater* is a clear, saturated solution of calcium hydroxide in water. It is used as a moderately strong base in the laboratory and in medicines. *Milk of lime*, which is a suspension of calcium hydroxide in water, is used for whitewash.

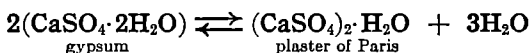
**Mortar.**—Lime mortar is prepared by mixing freshly slaked lime with sand and water to form a plastic mass. This material has the ability to adhere to bricks and stones, and therefore is widely used in building construction. The hardening or "setting" of mortar involves reactions which are, as yet, little understood. There is a slow evaporation of water, while, at the same time, an absorption of carbon dioxide from the air results in the formation of calcium carbonate on the surface. Within the mass the slaked lime is believed to combine, in part, with the sand ( $\text{SiO}_2$ ), forming calcium silicate. The equations are:



**Calcium Sulfate.**—Calcium sulfate occurs chiefly as *gypsum*, a soft, white crystalline solid, whose formula is  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ . This compound is employed as a fertilizer and as a raw material in the manufacture of cement. Its chief use, however, is in the preparation of *plaster of Paris*.

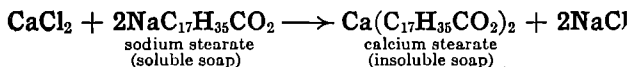
When gypsum is heated under carefully regulated conditions, three-quarters of its water of crystallization may be driven off, yielding plaster of Paris. On being mixed with water, plaster of Paris "sets" or hardens by recombining with the same quantity of water previously lost, thus forming a hard, crystalline mass of

artificial gypsum. Since these two reactions, are the reverse of each other, the equations may be combined into one as follows:



During the setting process, the mixture expands slightly; for this reason, it can be used as a molding medium for making accurate copies of articles such as medals, teeth, etc.

**Hard Water.**—Hard water contains certain mineral salts in solution, usually those of calcium and magnesium. Such water is undesirable for washing purposes, because it does not readily form a lather with soap. Instead, the dissolved salts react with the soluble soap, forming insoluble products:

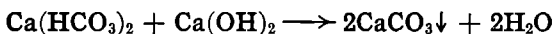


Hard water is also objectionable because it forms a mineral deposit known as "boiler scale" in steam boilers and heater pipes.

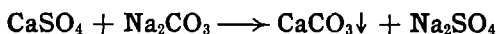
1. *Temporary hard water* contains in solution the bicarbonate of either calcium or magnesium. These impurities are often present in spring water of limestone regions because of the solvent action of water containing dissolved carbon dioxide. Small quantities of temporary hard water can be easily softened by boiling, which converts the soluble bicarbonates into insoluble carbonates:



On a large scale, temporary hard water can be softened by adding a calculated amount of calcium hydroxide, which precipitates insoluble calcium carbonate:

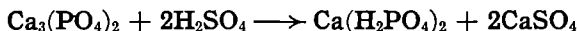


2. *Permanent hard water* usually contains calcium sulfate or chloride in solution, although the corresponding salts of magnesium and iron may also be present. It is called "permanent" hard water because boiling has no effect on it. The addition of washing soda, however, will precipitate the impurity in the form of an insoluble carbonate:



A highly efficient process for softening water on a large scale depends on the use of *permutit*, a complex salt containing sodium, silicon, and aluminum. The calcium sulfate in the water reacts with the permutit to form an insoluble calcium permutit, thus rendering the water soft. The calcium permutit is then converted back to sodium permutit by treating it with a solution of sodium chloride.

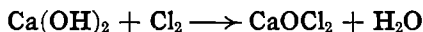
**Calcium Phosphate.**—Calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ , is present in large deposits of phosphate rock found in Tennessee, Florida, and South Carolina. These deposits are believed to have been formed from the skeletons of animals that lived in these regions during prehistoric ages. Calcium phosphate is the chief source of the phosphorus employed in making matches. The most important use of this compound, however, is as a fertilizer. Since calcium phosphate is insoluble in water, phosphate rock is first treated with sulfuric acid. This results in the formation of the soluble monocalcium acid phosphate:



The resulting mixture of the acid phosphate with the calcium sulfate is sold as fertilizer under the name of *superphosphate of lime*.

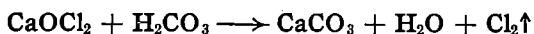
**Calcium Chloride.**—Calcium chloride ( $\text{CaCl}_2$ ) is the principal by-product in the Solvay Process for making sodium carbonate and sodium bicarbonate. Since it is an extremely deliquescent substance, it finds many uses as a drying agent, both in the laboratory and commercially. It is employed widely to settle the dust above dirt roads, which it does by absorbing moisture from the air. Calcium chloride is extremely soluble in water, yielding solutions that freeze at very low temperatures. Such solutions find extensive use in refrigeration and in the manufacture of ice.

**Bleaching Powder.**—Bleaching powder, or *chloride of lime*, has the formula  $\text{CaOCl}_2$ . It is prepared by passing chlorine over freshly slaked lime:



Bleaching powder is an unstable white powder, sparingly soluble in water. It reacts readily with acids, setting free chlorine. Even the exceedingly weak carbonic acid, formed by the union of water

and carbon dioxide absorbed from the air, is strong enough to decompose bleaching powder:



Because of its unstable character, bleaching powder provides a good source of free chlorine for bleaching and disinfecting (page 91).

### QUESTIONS

1. (a) Explain why calcium is not found free in nature. (b) Write the equation for the action of calcium on water.

2. Name four natural substances that consist chiefly of calcium carbonate.

3. Write equations for the following reactions: (a) calcium carbonate is heated strongly; (b) calcium carbonate is treated with dilute hydrochloric acid.

4. Explain (a) the formation of a limestone cave; (b) the growth of stalactites and stalagmites; (c) the gradual filling up of a cave. Write equations for the chemical actions involved.

5. A strong blast of air is blown into the bottom of the kiln during the production of calcium oxide from limestone. Explain why this is done.

6. (a) Why is calcium oxide called a basic anhydride? (b) Write an equation to explain your answer.

7. (a) Write the equations involved in obtaining water-slaked lime and air-slaked lime. (b) How would you distinguish between these two products?

8. (a) Name the ingredients in lime mortar, and state the use of each. (b) Write the equation for the chief action involved when mortar sets.

9. Explain the changes which take place in the following procedure, giving an equation in each case: (a) carbon dioxide is bubbled through clear limewater for a few seconds; (b) the bubbling is continued for a few minutes; (c) the resulting liquid is heated to the boiling point.

10. (a) Give the chemical name and formula of plaster of Paris. (b) Write the equation for the action involved when plaster of Paris sets.

11. Account for the formation of a white deposit on the bottom of a tea kettle in which hard water has been boiled.

12. (a) Name a compound often used for softening permanent hard water. (b) Write an equation to illustrate the action. (c) Explain how this action softens water.



13. (a) Why is phosphate rock treated with sulfuric acid in making phosphate fertilizer? (b) What is superphosphate of lime?

14. Calcium chloride is used to dry gases in the laboratory and to settle the dust on a country road. Name one property of this compound which is the basis for both these uses, and explain how it applies in each case.

15. Explain why chloride of lime is useful for (a) bleaching, (b) disinfecting.

### COMPLETION TEST

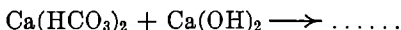
*Supply the word or words required to complete each of the following statements.*

1. The most abundant calcium compound found in nature is .....

2. A clear solution of calcium hydroxide is called .....; a suspension of calcium hydroxide in water is called .....

3. Temporary hard water may be rendered soft by ..... the water, or by adding .....

4. Complete and balance the following equation:



5. The application of heat converts ....., a soluble calcium compound, into ....., an insoluble calcium compound.

6. Hard water is objectionable for laundry purposes because .....

7. Plaster of Paris is made by driving out three-quarters of the ..... from ..... by the use of .....

8. Calcium in the diet supplies essential building material for ..... and .....

9. The salt most commonly found in permanent hard water is .....

10. Complete the following table of calcium compounds:

CHEMICAL NAME	COMMON NAME	FORMULA	USE
	Marble		
			Making mortar
Calcium hydroxide			
		$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A                               | B                                     |
|---------------------------------|---------------------------------------|
| ( ) Expands on setting          | 1. Precipitated chalk                 |
| ( ) Forms an insoluble soap     | 2. Phosphate rock                     |
| ( ) A reversible reaction       | 3. To remove temporary hardness       |
| ( ) A source of fertilizer      | 4. Heating limestone                  |
| ( ) Calcium chloride            | 5. Permutit                           |
| ( ) Pure calcium carbonate      | 6. Stalactite                         |
| ( ) Bleaching powder            | 7. Plaster of Paris                   |
| ( ) A commercial water-softener | 8. A by-product of the Solvay process |
| ( ) Limewater                   | 9. $\text{CaOCl}_2$                   |
| ( ) Boiling                     | 10. Air-slaked lime                   |
|                                 | 11. Hard water                        |
|                                 | 12. A mild base                       |

## CHAPTER 20

### IRON AND ITS COMPOUNDS

**Importance.**—Iron is the most valuable metal known to man. Almost every type of industrial activity in our modern machine age is dependent, in one way or another, on iron and steel. Vast quantities of the metal go into the manufacture of automobiles and railroad equipment, the erection of skyscrapers and bridges, the construction of ships, and the creation of a thousand different kinds of machines on which the productive capacity of our farms and factories depends. This accounts for the tremendous annual output of iron and steel—over one hundred million tons—of which about half is manufactured in the United States.

**Occurrence.**—Iron ranks fourth in abundance among the elements, being the most common metal in the earth's crust, with the sole exception of aluminum (Fig. 4, page 9). In its free state, iron is very rare, occurring chiefly in meteorites. In the combined form, however, iron occurs abundantly in many parts of the world. The principal ores of iron are *hematite* ( $\text{Fe}_2\text{O}_3$ ), *magnetite* ( $\text{Fe}_3\text{O}_4$ ), *limonite* ( $2\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$ ), and *siderite* ( $\text{FeCO}_3$ ).

In the United States, hematite is the most important iron ore. Vast deposits of this mineral are located in the regions about Lake Superior. In the open-pit mines of the Mesabi range in northern Minnesota, the ore is merely stripped from the surface by giant steam shovels. It is then transported, via the Great Lakes, to the iron and steel mills at Chicago, Gary, Cleveland, Buffalo, and Pittsburgh.

**Metallurgy of Iron.**—The chief operations involved in the process of converting iron ore into iron are the reduction of iron oxide with heated coke, and the removal of impurities by the use of a suitable flux. These processes are carried out in a *blast furnace*. The materials which go into the furnace are: (1) *iron ore*, usually hematite ( $\text{Fe}_2\text{O}_3$ ), to furnish the iron; (2) *coke*, a nearly pure, porous form of carbon, to supply the heat and act as a reducing agent; (3) *hot air*, to burn part of the coke in order to produce the necessary heat, and to form carbon monoxide, which acts as a reducing agent; and (4) *flux*, usually limestone ( $\text{CaCO}_3$ ),

# The Age of Steel

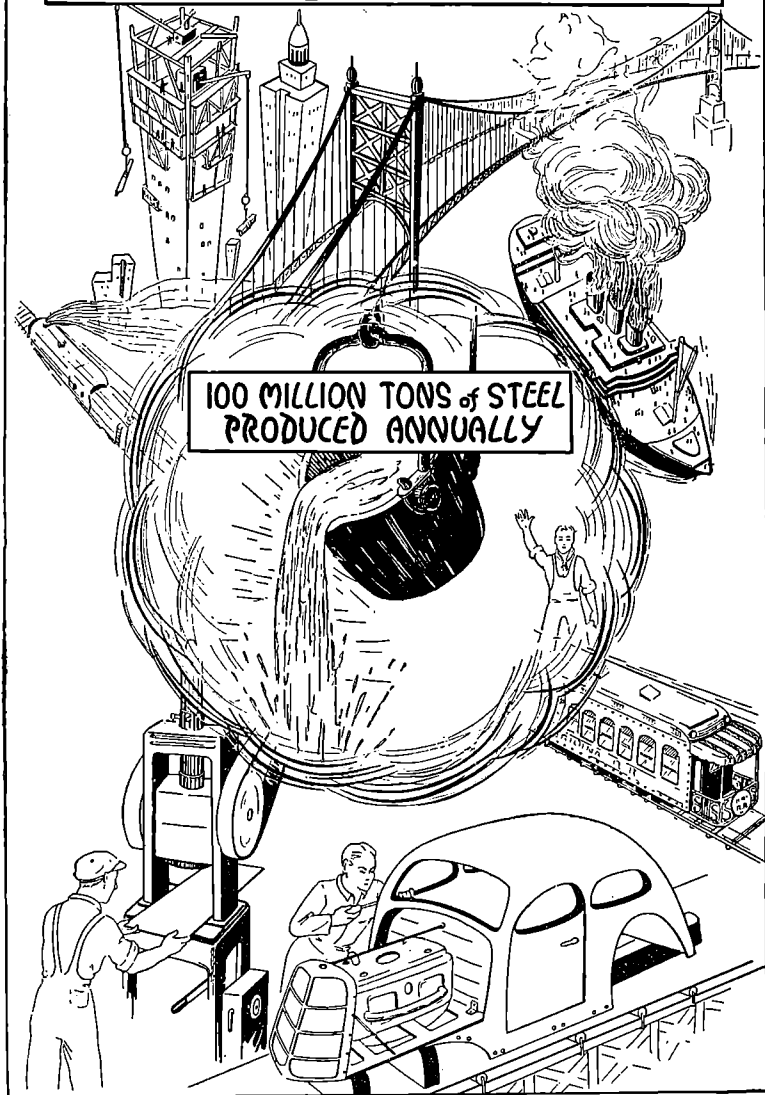
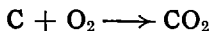


FIG. 65.

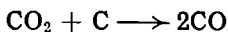
to combine with the impurities in the ore and form a liquid slag, which may be drawn off.

**The Blast Furnace.**—The metallurgy of iron is carried out in a huge cylindrical furnace about 100 feet high and 25 feet in diameter at its widest part (Fig. 66). It is built of wrought iron and is lined with fire brick. At the top, are two bell-shaped valves which can be opened alternately for the admission of the solid charge, thus preventing too great a loss of heat. Near the bottom, a series of pipes (*tuyères*) enters the furnace; through these pipes enormous volumes of heated air are blown. The waste gases resulting from the reactions in the furnace pass out through a pipe near the top. These gases contain about 25% carbon monoxide and are employed to heat the air that is forced into the furnace through the *tuyères*, and to operate the machinery of the plant. The molten iron and slag accumulate in the bottom of the furnace.

**Reactions in the Blast Furnace.**—The first operation which takes place is the combustion of some of the coke:



The heat produced by this reaction makes the upper layers of coke red-hot. As the carbon dioxide passes up through the heated mass, it is reduced to carbon monoxide:



This carbon monoxide, which is the chief reducing agent in the blast furnace, reacts with the iron ore, forming iron and carbon dioxide:

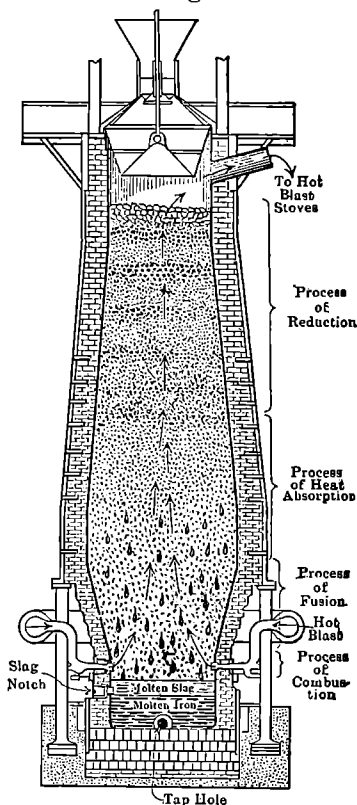
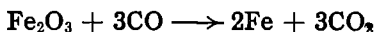
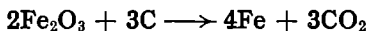


Fig. 66. Blast furnace.

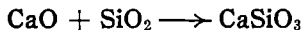
In addition, some ore is reduced directly by the highly heated coke in the lower part of the furnace:



The limestone, used as a flux, is decomposed by the intense heat:



The calcium oxide unites with the impurities in the ore, mainly sand and quartz ( $\text{SiO}_2$ ), forming a glass-like slag, calcium silicate:

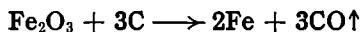


The molten slag floats above the molten iron that accumulates at the bottom of the furnace. The slag is drawn off through the "slag notch" every four hours. The liquid iron is run out through the "tap hole" every six hours, and is poured into molds, where it hardens into large bars of *pig iron* or *cast iron*.

The blast furnace operates continuously for months or years, being stopped only when repairs become necessary. In one working day a single blast furnace uses as much as 1000 tons of ore, 500 tons of coke, 250 tons of limestone, and 1200 tons of air, and turns out 600 tons of pig iron.

**Cast Iron.**—Cast iron, or pig iron, is the least pure of all forms of iron. It has a variable composition, containing usually about 93% iron, 5% carbon, and small amounts of silicon, sulfur, phosphorus, and manganese. Owing to the presence of these impurities, cast iron is a hard, grayish, brittle solid, possessing a crystalline structure. It melts readily at about 1150°C. It can be cast into various articles such as radiators, stove parts, and water pipes. Ordinary cast iron is not malleable, and cannot be tempered, welded, or forged. More than three-fourths of all the cast iron that comes from the blast furnace is used immediately for the production of wrought iron and steel.

**Wrought Iron.**—Wrought iron is the purest form of iron used commercially. It usually contains less than 0.5% of impurities, of which only about 0.1% is carbon. Wrought iron is made from cast iron by oxidizing most of the foreign matter contained in the latter. The process is carried out in a *reverberatory furnace*, where a powerful gas flame is used to melt the charge of cast iron. The furnace lining contains iron oxide, which provides oxygen for the removal of the carbon:



As the iron becomes purer, it turns thick and pasty. The metal is stirred (or *puddled*), and then drawn out of the furnace in the form of large lumps. These are hammered or squeezed to remove most of the liquid slag. The iron is then further rolled, a process which effects an even distribution of the remaining slag. This gives the wrought iron its characteristic *fibrous* structure.

Wrought iron is tough, flexible, malleable, and ductile. It can be forged and welded, but not tempered. It is easily magnetized or demagnetized. Wrought iron is well suited for making wire, chains, anchors, horseshoes, bolts, and cores of electromagnets. Because it is not easily corroded, wrought iron is useful for pipes and boiler tubes.

**Steel.**—Steel is a form of iron intermediate in purity between cast iron and wrought iron. It usually contains from 0.2% to 2% of carbon. Steel can be worked in a variety of ways, making it highly useful for a thousand different purposes. It can be tempered, forged into any desired shape, cast into molds, rolled into bars and sheets, and drawn into wire. Unlike wrought iron, steel, once magnetized, does not easily become demagnetized. The properties of steel depend not only upon the percentage of impurities (chiefly carbon) present, but also upon the method of manufacture and the heat treatment to which it is subjected.

#### Bessemer Process.

—The Bessemer converter provides a rapid and efficient method for making low-grade steel on a large scale. This device is a huge, egg-shaped cup made of wrought iron and lined with heat-resisting silica, which is an acid anhydride. It is supported on two axles, one of which is hollow to enable it to conduct a strong current of heated air up through the grate-like bottom (Fig. 67). The converter is first tipped forward to receive a charge of about fifteen tons of molten cast iron direct from the

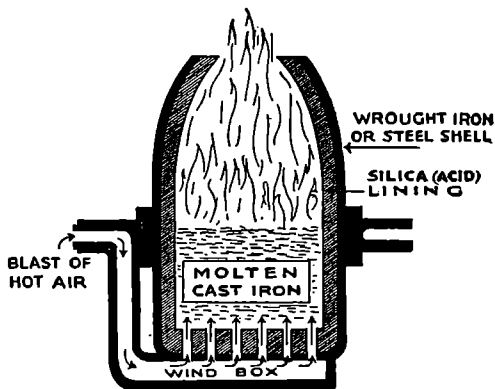
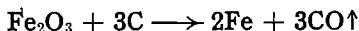


FIG. 67. Bessemer converter.

blast furnace. The huge pot is then righted, and a strong blast of heated air is blown through the liquid, which causes the carbon and other impurities to be burnt out. After about ten minutes, the oxidation is complete, as shown by a change in the size and color of the flame that emerges from the mouth of the converter. A weighed quantity of *spiegeleisen*, an alloy of iron and manganese that is rich in carbon, is now added. This provides the steel with just the correct amount of carbon. The manganese combines with any oxygen present, thus making the steel tougher. After a thorough mixing with the aid of the air blast, the finished steel is poured into a mold where it hardens, forming an *ingot*.

**Uses of Bessemer Steel.**—Bessemer steel is used to make wire, pipes, and sheet metal. It is also employed for certain types of structural work which do not involve excessive vibration or strain.

**Open-Hearth Process.**—A much higher grade of steel than that produced in the Bessemer converter is made in the open-hearth furnace (Fig. 68). This consists of a shallow bed, lined with some basic material like magnesium oxide, and capable of holding a charge of 60 to 200 tons of steel. Two sets of brick checker-work are located on each side of the furnace. Producer gas and air are blown in from one side, the mixture burning with a hot flame that beats down on the charge. The exhaust gases heat up the checker-work on the other side as they escape. The direction of the flow of gas and air is reversed every twenty minutes, so that the mixture always enters the furnace in a pre-heated condition. This *regenerative* principle results in the production of an extremely high temperature within the furnace. The charge used in this process consists of cast iron, scrap iron, limestone, and hematite ore. The action in the furnace involves mainly the oxidation of the impurities present in the cast iron. This is accomplished almost entirely by the hematite ( $\text{Fe}_2\text{O}_3$ ):



The resulting oxides of carbon, silicon, sulfur, and phosphorus escape as gases, or combine with the flux and basic lining, forming slag. At intervals, samples of steel are removed for analysis, until the composition is found to be exactly as desired. This usually takes from eight to twelve hours. Since a much more



precise control is possible than in the Bessemer process, a steel of finer quality is obtained. About three-quarters of all our steel is now made by the open-hearth process.

**Uses of Open-Hearth Steel.**—Open-hearth steel is used for armor plate, railroad tracks, girders for bridges and buildings, and for general structural work where exceptional strength is desired.

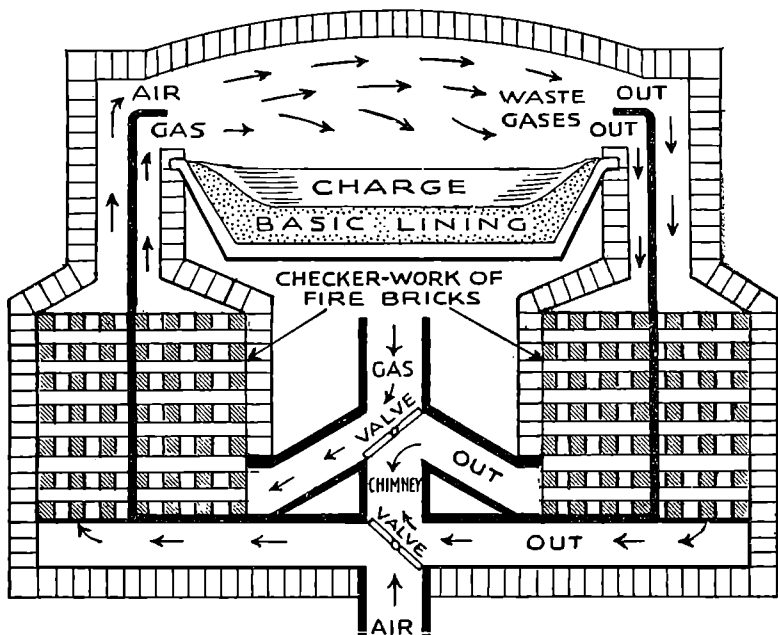


FIG. 68. Regenerative open-hearth furnace.

**Crucible Process.**—The crucible process is one of the oldest methods for making steel. Crucible steel is now made by heating several hundred pounds of wrought iron, scrap steel, and charcoal in a graphite vessel. The molten iron slowly absorbs some of the carbon, combining with it to form iron carbide ( $\text{Fe}_3\text{C}$ ). Since the charge in the crucible is a small one, the process can be controlled with great precision and the product is a high-carbon steel of superior quality. It is used for making penpoints, watch and clock springs, tools, and automobile parts.

**Electric Furnace Process.**—In recent years, the electrical method for producing high-grade steel has come into greater prominence. One type of furnace used for this purpose is shown in Fig. 69. The hearth is lined with some basic substance, usually dolomite. From the roof are suspended heavy graphite electrodes. The furnace is set on rollers, so that it can be tilted to discharge its contents. The charge consists of molten steel from the Bessemer converter, hematite, and a basic flux (usually limestone). A powerful current of electricity is passed through the slag that

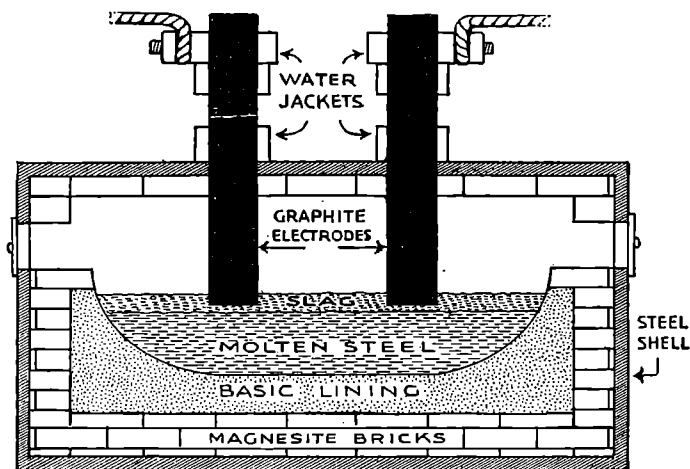


FIG. 69. Electric furnace for the manufacture of steel.

floats on the molten steel, thus generating an intense heat. This causes the oxidation of the acidic impurities by the flux and the basic lining. As in the case of the crucible process, the reaction can be very carefully controlled, thus yielding an unusually high-grade product. Electric furnace steel resembles crucible steel in composition and in some of its uses.

**Tempering of Steel.**—One of the chief reasons why steel is of such great importance is that when it is subjected to heat treatment it exhibits a remarkable range of properties. This heat treatment is called *tempering*, and is used to vary at will the toughness, strength, flexibility, and hardness of the steel.

1. *Hard steel* is made by heating a high-carbon steel above  $670^{\circ}\text{C}$ . and then cooling it suddenly by "quenching" in cold water or oil. This causes the iron carbide ( $\text{Fe}_3\text{C}$ ) which is dissolved in the hot steel to remain in the dissolved state. The product is therefore very hard and brittle.

2. *Softer steel* is made by heating steel as indicated above and allowing it to cool slowly. This permits a gradual crystallization of most of the iron carbide from solution. The resulting product is therefore softer but tougher.

If a hard steel is heated once more to a definite temperature and then cooled, it is left with any degree of hardness desired, depending on the temperature employed. This process is known as *drawing the temper*.

**Steel Alloys.**—In addition to tempering, the properties of steel are altered to a considerable degree by adding small amounts of other metals. In this way, special steels are produced which possess certain very desirable qualities, such as unusual hardness, toughness, and ability to resist corrosion.

1. *Manganese steel* contains about 12% of manganese. It is very hard and tough, and is used for making safes and rock crushers.

2. *Chrome steel* contains about 3% of chromium. It is very hard and elastic, and resists corrosion. It is used for making ball-bearings and roller-bearings, files, and armor plate. *Stainless steel* contains up to 18% of chromium.

3. *Nickel steel* is hard, tough and elastic, besides being resistant to corrosion. It is used for making armor plate, cannon, and bridges.

4. *Tungsten steel* is extremely hard, and is used for making cutting tools which must retain a sharp edge even at high temperatures.

5. *Molybdenum steel* possesses great hardness and tensile strength, and is used for making axles and high-speed lathe tools.

6. *Vanadium steel* is noted for its high tensile strength and elasticity. It is used for making axles, crank shafts, and gears for automobiles.

7. *Silicon steel* is very flexible. It is used for making automobile springs, and cores of transformers and electromagnets.

**Compounds of Iron.**—When combined with other elements, the atom of iron exhibits an irregularity in the arrangement of its

electrons. As a result it can lose either two or three electrons, thus forming the ions  $\text{Fe}^{++}$  or  $\text{Fe}^{+++}$ . Therefore, we have two series of iron compounds. Those in which the iron has a valence of 2 are called *ferrous* compounds, as ferrous chloride ( $\text{FeCl}_2$ ). Those in which the iron has a valence of 3 are called *ferric* compounds, as ferric chloride ( $\text{FeCl}_3$ ). Ferrous compounds are usually green, while ferric compounds are yellow, brown or red. Ferrous ions are not as stable as ferric ions; therefore, ferrous compounds have a tendency to change into the corresponding ferric compounds.

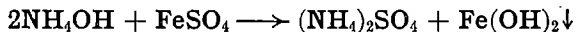
**Iron Oxides.**—Iron combines with oxygen in three different ways, forming the following compounds:

1. *Ferrous oxide* ( $\text{FeO}$ ) is a black powder which is unstable, readily taking on more oxygen from the air to form ferric oxide.

2. *Ferric oxide* ( $\text{Fe}_2\text{O}_3$ ) occurs abundantly in nature as *hematite*, the chief ore of iron. It is reddish in color, and is one of the products formed during the rusting of iron. Ferric oxide is used as a paint pigment (*Venetian red*), and as a polishing powder (*jeweler's rouge*).

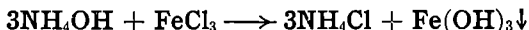
3. *Magnetic oxide of iron* ( $\text{Fe}_3\text{O}_4$ ) occurs in nature as the iron ore *magnetite* and as *lodestone*, which is a natural magnet. This oxide may be regarded chemically as ferro-ferric oxide ( $\text{FeO} \cdot \text{Fe}_2\text{O}_3$ ). It is a black powder, and may be prepared in the laboratory by passing steam over red-hot iron. It differs from  $\text{Fe}_2\text{O}_3$  in that it acts as a protective layer to prevent rusting. *Russia iron* is sheet iron that has been covered with a thin coating of  $\text{Fe}_3\text{O}_4$ .

**Iron Hydroxides.**—*Ferrous hydroxide*,  $\text{Fe}(\text{OH})_2$ , may be prepared by treating a ferrous salt solution with ammonium hydroxide:

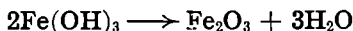


When first precipitated, the compound appears white, but the oxygen of the air soon changes it to a grayish-green color, and finally to reddish-brown, owing to the formation of ferric hydroxide.

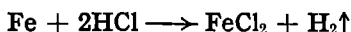
*Ferric hydroxide*,  $\text{Fe}(\text{OH})_3$ , may be prepared by treating a ferric salt solution with any soluble base.



Ferric hydroxide is a reddish-brown, flocculent precipitate. When heated red-hot, it loses water and changes to ferric oxide:

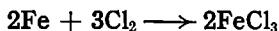


**Iron Chlorides.**—*Ferrous chloride* ( $\text{FeCl}_2$ ) is prepared by treating iron with hydrochloric acid:

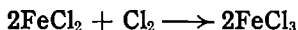


It is a pale-green, crystalline compound, unstable in air.

*Ferric chloride* ( $\text{FeCl}_3$ ) is prepared by the action of chlorine on hot iron:



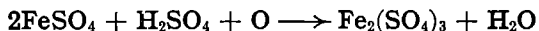
It may also be prepared by oxidizing ferrous chloride with chlorine:



Ferric chloride is a yellow, deliquescent salt, which dissolves easily in water. This solution gives an acid reaction, owing to hydrolysis. It is used as a mordant in dyeing, and as an astringent in medicine.

**Iron Sulfates.**—*Ferrous sulfate* ( $\text{FeSO}_4$ ) is the most familiar ferrous compound. It is formed as a by-product when steel plates are cleaned with sulfuric acid before being galvanized or tinned. The crystallized form ( $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ ) is known as *copperas* or *green vitriol*. Ferrous sulfate is used in making ink, in tanning, and as a mordant in dyeing.

*Ferric sulfate*,  $\text{Fe}_2(\text{SO}_4)_3$ , is prepared by oxidizing ferrous sulfate in the presence of sulfuric acid:

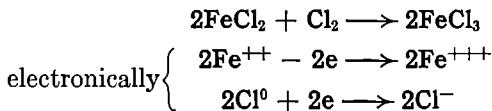


It is used to form a series of alums (page 237).

**Oxidation and Reduction.**—In a previous chapter we learned that oxidation is the union of a substance with oxygen, while reduction is the removal of oxygen from a compound by the use of some substance called a reducing agent. In our study of the Electron Theory, this conception was somewhat amplified. Oxidation was found to involve a *loss* of electrons, which means an *increase* in positive valence. Similarly, reduction involves a *gain* of electrons, which means a *decrease* in positive valence. This broader meaning is well illustrated by the conversion of ferrous to ferric salts and vice versa.

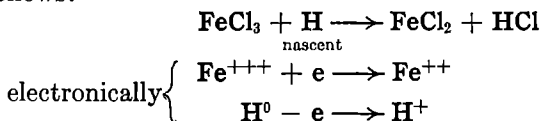
1. *Oxidation of Ferrous Chloride.*—When chlorine is bubbled

through a solution of ferrous chloride, a reaction takes place which may be represented as follows:



This shows that each atom of the neutral chlorine molecule takes one electron from a *ferrous* ion ( $\text{Fe}^{++}$ ), thus changing it to a *ferric* ion ( $\text{Fe}^{+++}$ ), while the chlorine atom ( $\text{Cl}^0$ ) becomes a chlorine ion ( $\text{Cl}^-$ ). The valence of the iron has thus been increased from 2 to 3, that is, the iron has been oxidized. A broad concept of oxidation is that it is: (a) combination with oxygen; or (b) loss of electrons; or (c) raising of positive valence.

2. *Reduction of Ferric Chloride.*—When nascent hydrogen is formed in a solution of ferric chloride by the action of zinc and hydrochloric acid, a reaction takes place which may be represented as follows:



This shows that each *ferric* ion ( $\text{Fe}^{+++}$ ) gains one electron from a neutral hydrogen atom ( $\text{H}^0$ ) and is thus changed to a *ferrous* ion ( $\text{Fe}^{++}$ ), while the hydrogen atom ( $\text{H}^0$ ) becomes a hydrogen ion ( $\text{H}^+$ ). The valence of the iron has thus been lowered from 3 to 2, that is, the iron has been reduced. A broad concept of reduction is that it is: (a) removal of oxygen; or (b) gain of electrons; or (c) lowering of positive valence.

**Iron Ink.**—Ferrous sulfate combines with tannic acid, which is extracted from nut galls (a certain "tumor" or growth found on trees) to form ferrous tannate. Ordinary fountain pen ink contains this compound, together with gum arabic and a small quantity of carboic acid to act as a mold-destroyer. Ferrous tannate is nearly colorless, but is readily oxidized by contact with the air, forming black, insoluble ferric tannate. In order to give an immediate color to the writing, a blue dye is added to the ink.

**Removal of Ink Stains.**—A freshly made iron ink stain can usually be removed by merely washing with water, since all the

ingredients of ink are soluble. After oxidation has taken place, forming insoluble ferric tannate, some reducing agent must be employed to convert the ferric salt to the soluble ferrous salt. Dilute oxalic acid or citric acid (present in lemon juice) will accomplish this result. After time is allowed for the reduction to take place, the fabric must be thoroughly washed. The same procedure is effective in removing rust stains.

**Blue Prints.**—Blue-print paper is paper which has been covered with a thin coating of potassium ferricyanide and ferric ammonium citrate, and then dried in the dark. Under such conditions the two substances do not react with each other. When the paper is to be used, it is covered by a “negative,” which consists of a black-line drawing, on translucent tracing paper. A strong light is then directed on the drawing, which causes the ferric ammonium citrate to be reduced to the corresponding ferrous salt. This reacts with the potassium ferricyanide, forming insoluble Turnbull’s blue. Wherever the black lines of the drawing have shut out the light, no reduction takes place. When the paper is washed in water, the unaffected salts are dissolved away, while the insoluble blue compound adheres to the surface. Thus the final print consists of white tracings against a deep blue background.

### QUESTIONS

1. (a) Give the chemical name, common name and formula of the chief ore of iron. (b) Where is this ore found in large quantities?
2. With the aid of a labeled diagram, describe the construction of the blast furnace.
3. Name the four substances that enter the blast furnace, and give a reason for the use of each.
4. (a) Describe the reactions that take place in the blast furnace. (b) Write six equations for the chemical actions involved.
5. What advantages are there in keeping the blast furnace in continuous operation?
6. What flux would you use to remove dolomite ( $MgCO_3$  and  $CaCO_3$ ) impurity from an iron ore?
7. What use is made of slag?
8. (a) What is the composition of the exhaust gas which leaves the top of the blast furnace? (b) Name two uses for this gas.
9. (a) What is the chief impurity present in cast iron? (b) State three properties and three uses of cast iron.

10. Wrought iron was probably the earliest form of iron used by man. Explain the reasons for this.
11. How does wrought iron differ from cast iron in (a) composition, (b) properties, (c) uses?
12. (a) Name the furnace used in making wrought iron, and explain the name. (b) Describe the process of making wrought iron.
13. Name five articles made of wrought iron, and explain why it is used in their manufacture.
14. With the aid of a labeled diagram describe the process for making Bessemer steel.
15. Cold air blown through the molten cast iron in the Bessemer converter causes the metal to become hotter. Explain.
16. In the Bessemer process all the carbon is first burned out, after which additional carbon is added. Explain.
17. With the aid of a labeled diagram, explain the construction and operation of the open-hearth furnace.
18. Explain why most of our steel is produced by the open-hearth process, rather than by the Bessemer process.
19. Why is the term "regenerative" applied to the open-hearth furnace?
20. (a) Explain how wrought iron is changed to crucible steel. (b) State four uses for crucible steel.
21. (a) Define *tempering*. (b) Explain how steel is tempered.
22. (a) Name four metals commonly alloyed with steel. (b) Name the special property which each imparts to steel. (c) State one use for each steel alloy.
23. (a) Explain how you would oxidize ferrous chloride to ferric chloride. (b) Define oxidation in terms of change of valence. (c) Explain what happens according to the Electron Theory.
24. (a) Explain how you would reduce ferric chloride to ferrous chloride. (b) Define reduction in terms of valence. (c) Explain what happens according to the Electron Theory.
25. Describe a chemical test to distinguish between a ferrous salt and a ferric salt.
26. An iron nail is dissolved in dilute sulfuric acid. Which is formed, ferrous sulfate or ferric sulfate? Explain.
27. Write an electronic equation for the reduction of ferric chloride by iron.
28. Explain the difference in stability of ferrous compounds and ferric compounds.
29. Fountain pen ink writes blue but turns black after exposure to the air. Explain.
30. Account for the blue background and the white lines on a blue print.



## COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. Meteorites contain about the only . . . . . iron found in nature.
2. The two uses of coke in the blast furnace are . . . . .
3. The purest commercial form of iron is . . . . .; the least pure form is . . . . .
4. In a metallurgical furnace, an acid flux is always used to remove a . . . . . impurity.
5. The substance formed when a flux combines with an impurity is called . . . . .
6. Cast iron cannot be welded or forged because . . . . .
7. Wrought iron is used for making electromagnet cores because . . . . .; steel is used for making permanent magnets because . . . . .
8. The Bessemer process takes . . . . . (more, less) time than does the open-hearth process.
9. Stainless steel is usually made by alloying the steel with . . . . .
10. Russia iron is sheet iron protected by a covering of . . . . .
11. When steam is passed over hot iron, the two products formed are . . . . .
12. When ferrous hydroxide is freshly prepared, it has a . . . . . color. On exposure to . . . . ., it soon turns . . . . ., owing to the formation of . . . . .

## MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

- | A                           | B                            |
|-----------------------------|------------------------------|
| ( ) Cast iron               | 1. Reverberatory furnace     |
| ( ) Calcium silicate        | 2. Calcium carbonate         |
| ( ) Cuprous                 | 3. Prussian blue             |
| ( ) Tuyère                  | 4. Ferro-ferricyanide        |
| ( ) For making wrought iron | 5. Slag                      |
| ( ) For making steel        | 6. An iron ore               |
| ( ) A flux                  | 7. Air pipe in blast furnace |
| ( ) Limonite                | 8. Magnetite                 |
| ( ) Turnbull's blue         | 9. Ferrous sulfate           |
| ( ) Oxidation               | 10. Open-hearth furnace      |
|                             | 11. Loss of electrons        |
|                             | 12. Very hard but brittle    |

## CHAPTER 21

### ALUMINUM AND ITS COMPOUNDS

**Importance.**—The great aluminum industry of the present day stands as one of the finest monuments to the triumph of the chemist over obstacles that seemed all but insurmountable. Although its compounds are extremely abundant, this element remained a mere laboratory curiosity until Charles M. Hall, a student at Oberlin College, Ohio, devised an ingenious method of extracting it from its ores by electrolysis. The present annual output of aluminum is 350 million pounds. Its price is now about 20 cents a pound, as against 90 dollars a pound when the metal was first placed on the market.

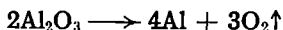
**Occurrence.**—Aluminum is the most common metallic element present in the earth's crust, being exceeded in abundance only by the non-metals, oxygen and silicon. In spite of its great abundance, it is never found free in nature. In the combined state, it is present in the form of complex silicates such as ordinary *clay*, *feldspar* and *mica*. The last two minerals occur in many rocks. Aluminum is present also in *bauxite* ( $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$ ), *cryolite* ( $\text{Na}_3\text{AlF}_6$ ), and *alundum* ( $\text{Al}_2\text{O}_3$ ).

**Preparation.**—Clay, the most abundant compound of aluminum, is, strangely enough, not the ore from which the metal is extracted commercially. The Hall electrolytic process, essentially the same as when it was first devised in 1886, employs *bauxite* as the electrolyte. The cell is an iron box having a lining of carbon, which acts as the cathode (Fig. 70). Suspended carbon rods act as anodes. The charge consists of pure *aluminum oxide* obtained from bauxite, and melted *cryolite*, which acts as a solvent, much as water is the solvent for sodium chloride in the electrolytic preparation of chlorine. The aluminum oxide dissociates in solution as follows:



The passage of the direct current of electricity results in the attraction of the  $\text{Al}^{+++}$  ions to the cathode, and the production of metal-

lic aluminum at the bottom of the box. The oxygen is similarly liberated at the carbon anodes:



The process is a continuous one, aluminum oxide being added at intervals, while aluminum is drawn off through the outlet at the bottom of the cell.

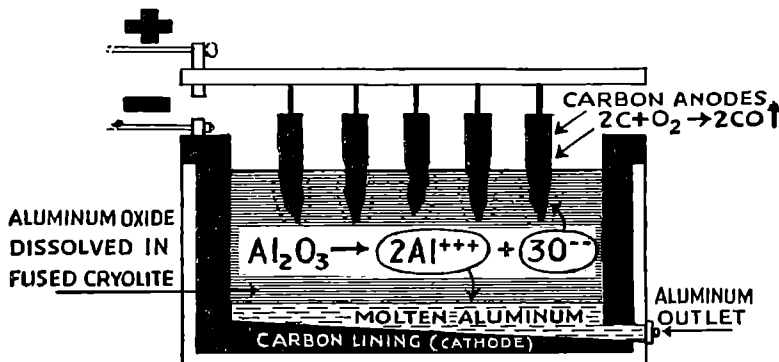
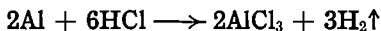


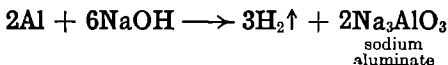
FIG. 70. Extraction of aluminum by the Hall Process.

**Properties.**—Aluminum is silvery-white in color, and is the lightest of the ordinary metals, being only one-third as heavy as iron. Among the common metals, only copper and steel surpass it in tensile strength. It is malleable and ductile, and is an excellent conductor of heat and electricity.

Chemically, aluminum is a moderately active metal, being oxidized quickly by contact with the air. However, the oxide thus formed acts as a protective coating which halts any further corrosion; in this respect, of course, it differs from ordinary iron rust. Owing to this protective coating, aluminum is not affected by hot or cold water. Sulfuric and nitric acids have only a very slight effect on the metal, but hydrochloric acid attacks it readily, causing a rapid evolution of hydrogen:



Aluminum is an amphoteric element, being dissolved by strong bases. This process results in the formation of a class of salts called *aluminates*:



Because of this reaction, aluminum utensils should not be cleaned with strongly alkaline substances, such as washing soda.

**Uses.**—The phenomenal reduction in the price of aluminum has opened vast new fields for the commercial utilization of this metal. Some of these uses are:

1. *Automobile and Aircraft Construction.*—These industries provide the largest single market for aluminum, consuming about one-third of the total annual output. When used for such construction purposes, aluminum is more serviceable if mixed with other metals. Among these alloys are *duralumin* (Al, Cu, Mg), which is used in the framework of airships; *aluminum bronze*, (Al, Cu); and *magnalium* (Mg, Al).

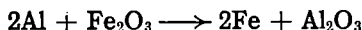
2. *Kitchen Utensils.*—Its lightness and freedom from tarnish make aluminum an ideal material for kitchen utensils.

3. *Electric Transmission Lines.*—Because of its lightness, aluminum is now replacing copper for use in long-distance electric transmission lines.

4. *Aluminum Foil.*—This material is replacing tin foil as a wrapper for food, candy, and other products. It is also used in flashlight bulbs.

5. *Reduction.*—In powder form, aluminum is used as a powerful reducing agent in the thermit process of welding iron.

**Thermit.**—Aluminum has a strong attraction for oxygen, and therefore is useful in the reduction of metallic oxides. Thermit is a mixture of powdered aluminum and the oxide of the metal desired (usually iron oxide). When this mixture is ignited by means of a magnesium ribbon fuse, reduction takes place, accompanied by the evolution of an enormous amount of heat, and the production of the molten metal:

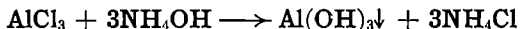


Thermit is used for welding railroad tracks, and for making emergency repairs on propeller shafts, locomotive parts, and other heavy machinery. The thermit action is employed to extract chromium, manganese, and other metals from their oxides.

**Aluminum Oxide.**—Aluminum oxide ( $\text{Al}_2\text{O}_3$ ) exists in several forms in nature, the most common ones being *bauxite*, *corundum*, and *emery*. *Alundum* is an artificial form of corundum. Both corundum and emery are extremely hard, and are used as abrasives. *Rubies* and *sapphires*, whether found in nature or prepared syn-

thetically, are almost entirely aluminum oxide. The colors of these gems are due to minute quantities of impurities.

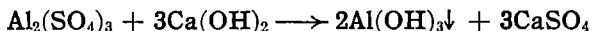
**Aluminum Hydroxide.**—Aluminum hydroxide,  $\text{Al}(\text{OH})_3$ , is a white, gelatinous solid, precipitated by the action between ammonium hydroxide and any soluble aluminum compound:



Aluminum hydroxide has several important industrial uses:

1. *Dyeing.*—Certain dyeing processes depend on the use of a *mordant*, which is the name given to any substance that helps the dye to stick to the cloth. A mordant does this by adsorbing the dye and forming an insoluble product called a "lake." Aluminum hydroxide is a good mordant.

2. *Purification of Water.*—For removing turbidity in water, aluminum hydroxide (*coagulum*) is formed by adding aluminum sulphate and lime to the water:



The gelatinous precipitate slowly settles to the bottom, carrying down suspended clay and other particles, and, incidentally, bacteria and organisms that cling to these particles.

**Alums.**—An alum is a double sulfate of a univalent metal and a trivalent metal. *Potassium alum*, the most common compound of this type, is potassium aluminum sulfate, with the formula  $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ . It is prepared by crystallizing a mixed solution of the two sulfates. It is used in forming the mordant, aluminum hydroxide; as a *styptic* to check the flow of blood from a small cut; and in certain baking powders. Other alums are soda alum,  $\text{NaAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ ; and chrome alum,  $\text{KCr}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ .

**Aluminum Silicates.**—Complex compounds of aluminum and silicon with other elements occur abundantly in nature. These minerals are employed in making ceramics (porcelain, china, earthenware, and other pottery) and cement.

1. *Ceramics.*—All forms of pottery are made from clay, which is an impure aluminum silicate, resulting from the weathering of rocks containing feldspar. The better grades of pottery, such as chinaware and porcelain, are made from *kaolin*, a pure form of clay. The material is mixed with a little water to form a plastic mass, which is molded and then baked in a kiln. In order to make the

pottery non-porous it must be *glazed*. Cheap pottery is glazed with a coating of sodium aluminum silicate, formed by the action of vaporized sodium chloride on the clay. Chinaware and porcelain are glazed by dipping the object into a suspension of kaolin mixed with ground feldspar and quartz, and then firing in the kiln. This forms a thicker, harder, and more resistant coating than the glaze produced from salt.

2. *Cement*.—Portland cement is made by strongly heating a finely powdered mixture of limestone and clay in a rotary kiln. A series of involved chemical actions take place, resulting in the formation of *clinker*, which is believed to be a complex calcium aluminum silicate. After cooling, this product is ground up with about 2% of gypsum, which prevents the cement from hardening too rapidly. Cement sets or hardens as a result of complex chemical changes involving a combination with water. *Concrete* is a mixture of cement, sand, crushed rock, and water. It hardens equally well under water and in air. *Reënforced concrete* has twisted steel rods or wire mesh embedded in it to give it great strength and rigidity. Its greatest use is in the construction of foundations, buildings, dams, bridges, and highways.

### QUESTIONS

1. Many metals are extracted from their ores by reduction with carbon. Explain why this cannot be done in the case of aluminum.

2. (a) With the aid of a labeled drawing, describe the electrolytic process for the metallurgy of aluminum. (b) State the purpose of each raw material used. (c) Give two reasons why aluminum drops to the bottom of the container.

3. Why is the metallurgy of aluminum in the United States carried on largely in the vicinity of Niagara Falls?

4. Explain why aluminum is far less subject to corrosion or rust than iron, although it is chemically more active than iron.

5. Explain why washing soda should not be used in cleaning aluminum utensils.

6. (a) Name an alloy of aluminum. (b) Give a use for this alloy, and explain why the alloy is superior to pure aluminum for this particular purpose.

7. (a) What is *thermit*? (b) For what purpose is it used? (c) Write an equation illustrating its use.

8. Since aluminum hydroxide is a very weak base, what effect would an aluminum sulfate solution have on litmus paper?

9. Compare the chemical actions involved in the hardening of (a) lime mortar, (b) plaster of Paris, (c) cement.

10. (a) What is *coagulum*? (b) Describe the method of purifying water by the use of alum.

### COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. The man chiefly responsible for the development of the aluminum industry is .....

2. In the electrolytic metallurgy of aluminum, ..... is the electrolyte and ..... is the solvent.

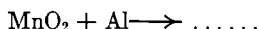
3. Aluminum is used to transmit electricity over long distances because .....

4. Aluminum reacts with concentrated ....., forming potassium aluminate.

5. Two reasons why aluminum is desirable for making kitchen utensils are .....

6. The formula of sodium alum is .....

7. *Complete and balance the following equation:*



8. A ruby consists almost entirely of the compound .....

9. Ordinary clay is the compound ..... in an impure state. A very pure variety of clay is commonly known as .....

10. Pottery is glazed in order to .....

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

A	B
( ) Aluminum bronze	1. $\text{Na}_3\text{AlF}_6$
( ) Porcelain	2. Common alum
( ) Aluminum hydroxide	3. A light but strong alloy
( ) Cryolite	4. Sodium aluminate
( ) For welding	5. Aluminum and copper
( ) Aluminum oxide	6. Made from kaolin
( ) Duralumin	7. A gelatinous substance
( ) $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$	8. A protective coating
( ) Clinker	9. Chrome alum
( ) Aluminum foil	10. A compound of Ca, Al and Si
	11. Thermit
	12. For flashlight bulbs

## CHAPTER 22

### COPPER AND ITS COMPOUNDS

**Importance.**—The history of copper is intimately bound up with the long history of human progress. Copper was the first metal which primitive man learned to use as a material for tools and weapons—a tremendous advance over the previous dependence on stone. The passage of the centuries has enhanced, rather than diminished, the importance of this reddish metal. Modern civilization is justly known as the “Age of Electricity,” but the electrical industries (and with them our entire economic life) would be paralyzed if they were deprived of the use of copper.

**Occurrence.**—Copper occurs free in rather abundant quantities, which explains its extensive use in primitive times. In the United States, there are rich deposits of free, or native, copper in northern Michigan. There are many ores of copper, some of them containing compounds that possess complex formulas. *Cuprite* ( $\text{Cu}_2\text{O}$ ) and *chalcocite* ( $\text{Cu}_2\text{S}$ ) are found in Arizona. The principal copper ore is *chalcopyrite*, ( $\text{CuFeS}_2$ ), which is mined in Montana and Utah. More than half of the world's copper supply is produced in the United States.

**Metallurgy.**—When copper occurs in the native state, the metal-bearing rock is crushed and heated. The molten metal is then drained away to be refined. The sulfide copper ores (which are the most common) are first roasted to effect a partial conversion of the sulfides to the oxide; the oxide, in turn, is then reduced by carbon in a blast furnace. The impurities, including iron, sulfur, silica, and others, must be removed. The final product is *blister copper*, which is nearly pure copper, containing small amounts of gold and silver.

The last step in the refining of copper involves the process of electrolysis (Fig. 71). The impure copper is in the form of large slabs, which are suspended as *anodes* in a bath of copper sulfate containing a little sulfuric acid. The *cathodes* are thin sheets of pure copper, hung between, and close to, the anodes. When the



current passes through the electrolyte, the copper ions (of the dissolved copper sulfate) are attracted to the cathodes. Here they gain electrons, become neutral copper ions, and deposit as a layer of pure metal. At the same time, copper from the anode goes into solution as copper ions. The gold and silver present in the slab of

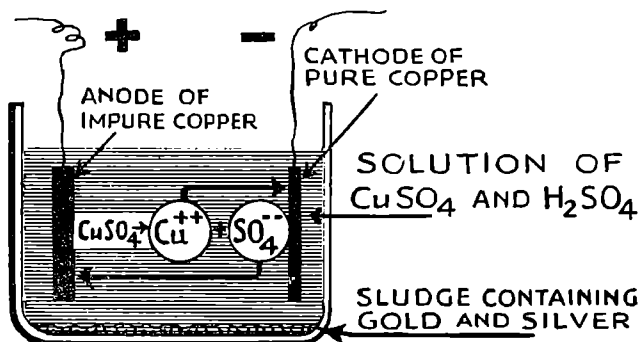
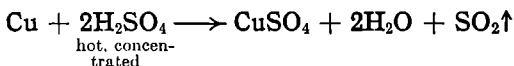
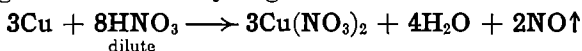


FIG. 71. Electrolytic process for refining copper.

metal drop to the bottom of the tank as a sludge, and are later recovered. The anode copper obtained by this process is 99.95% pure.

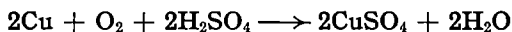
**Properties.**—Copper is a heavy metal with a characteristic reddish color. Next to gold and silver, it is the most malleable and ductile of the common metals, and is surpassed as a conductor of electricity by silver only. Copper oxidizes slowly in the air, forming a protective coating that prevents the corrosion from penetrating. Bronze statues and copper roofs soon turn green, owing to the formation of basic copper carbonate,  $\text{Cu}(\text{OH})_2 \cdot \text{CuCO}_3$ , which results from the action of the moisture, oxygen, and carbon dioxide present in the air.

Copper is below hydrogen in the replacement series of metals; thus, it cannot displace the hydrogen of acids. Hydrochloric acid and dilute sulfuric acid, therefore, have little effect on copper. Nitric acid and hot, concentrated sulfuric acid, both of which are oxidizing agents, attack copper; this reaction results in the evolution of gases other than hydrogen:



**Uses.**—Next to iron, copper is our most useful metal. Over one million tons of the latter are produced annually, of which about one-third is employed in the manufacture of all types of electrical apparatus, such as electric motors and dynamos; telephone, telegraph, cable, and power transmission lines; radio equipment; and the electric circuits of automobiles. For these purposes, the metal must be of exceptional purity, for the presence of even minute quantities of other elements reduces the conductivity of copper enormously. Copper is also widely used for boilers, vacuum kettles, leader pipes, and roofing. In copper-plating, a sheet of the pure metal is used as an anode, and the object to be plated is used as a cathode. The same method is also employed in making plates used in printing and illustrating books. The words on this page were printed with a copper electrotype. Considerable copper is used in making alloys, among the most important of which are brass, bronze, and gun metal. (The composition of these and other copper alloys is given in the Appendix, page 323.)

**Copper Compounds.**—Copper forms two series of salts: *cuprous*, in which copper has a valence of 1; and *cupric*, in which it has a valence of 2. These are illustrated by cuprous oxide ( $\text{Cu}_2\text{O}$ ), and cupric oxide ( $\text{CuO}$ ). *Copper sulfate* ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ), sometimes called *blue vitriol* or *bluestone* in the crystalline form, is the most important copper salt. It is prepared by the action of air and warm, dilute sulfuric acid on copper:



Copper sulfate crystals are somewhat efflorescent. In water solution, the salt hydrolyzes, giving an acid reaction. It is used in electric batteries, as a mordant in dyeing, in copper plating, in electrotyping, in refining copper, in purifying water, and in making fungicides. One well-known fungicide is *Bordeaux mixture*, consisting of solutions of calcium hydroxide and copper sulfate. This mixture is sprayed on vines and trees to kill tiny parasitic growths which live on the plants.

## QUESTIONS

1. Explain why man learned to make copper implements before iron implements.
2. Name two ores of copper, and indicate the general method used in their metallurgy.
3. (a) Draw a diagram illustrating the electrolytic refining of copper. (b) Show how the Electron Theory is useful in explaining the process.
4. When the Statue of Liberty was first erected, it had a metallic luster. The statue is now green. Explain.
5. Write the equation for the action of dilute nitric acid on copper.
6. In copper-plating, what would you use as (a) the anode, (b) the cathode, (c) the electrolyte? Explain the action.
7. (a) State the chief use of copper. (b) Explain why extremely pure copper is needed for this purpose.
8. Explain why a copper salt should not be put into an iron container.
9. A solution of cupric chloride turns blue litmus red. Explain this action.
10. How could you prove that a given substance is copper sulfate?

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

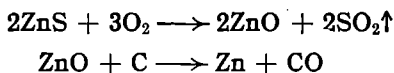
- | A                        | B                            |
|--------------------------|------------------------------|
| ( ) Chalcopyrite         | 1. Copper and tin            |
| ( ) A protective coating | 2. A fungicide               |
| ( ) Cuprous chloride     | 3. Blue vitriol              |
| ( ) Electrotype          | 4. Dissolves copper          |
| ( ) Gun metal            | 5. $\text{CuFS}_2$           |
| ( ) Blister copper       | 6. Copper, tin, and zinc     |
| ( ) Bordeaux mixture     | 7. For printing books        |
| ( ) Bronze               | 8. Basic copper carbonate    |
| ( ) Nitric acid          | 9. Cathode                   |
| ( ) Brass                | 10. $\text{CuCl}$            |
|                          | 11. Contains gold and silver |
|                          | 12. Copper and zinc          |

## CHAPTER 23

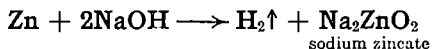
### OTHER METALS

**Importance.**—Of the ninety-two elements which are the building blocks of all matter, more than seventy are metals. Of these, the great majority are so rare and of such limited commercial use (if any) that they are scarcely ever mentioned, save in specialized textbooks. We have already encountered a half-dozen metals—sodium, potassium, calcium, iron, aluminum, and copper—which, both in themselves and in their compounds, are of great practical importance. We shall now take up briefly a number of other metals which are far less abundant than those already treated, but, in a variety of ways, have demonstrated their usefulness to man.

**Zinc.**—Zinc does not occur free in nature. Its most common ore is *zinc blende* ( $\text{ZnS}$ ). The first step in the metallurgy of zinc is to convert the sulfide to the oxide by roasting. Then the oxide is reduced with coke in small clay retorts. The heat applied to effect the reduction is sufficient to vaporize the zinc, which is then condensed to a liquid, run into molds, and allowed to harden into ingots, known as *speller*.



Zinc is a bluish-white metal which is readily converted into vapor form. When molten zinc is slowly poured from a great height into water, the falling drops of metal cool into irregular shapes. In this form it is known as *granulated zinc*. At ordinary temperatures, it reacts with the air, forming a protective coating. Zinc readily liberates hydrogen from most acids. It also reacts with a strong, hot solution of a base, forming a *zincate* and setting hydrogen free:



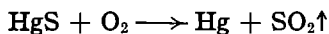
Zinc is used for roofing, lining tanks, and coating sheet iron (galvanizing) to prevent it from rusting. In the printing industry,

zinc is used to make plates for illustrations. (The illustrations in this book were reproduced by means of zinc plates.) The metal is also employed for making the containers of dry cells, and for preparing alloys, such as brass, bronze, and German silver. *Zinc oxide* ( $\text{ZnO}$ ), commonly known as zinc white, is used to make zinc ointment and serves as a white pigment in paints. *Zinc sulfate* ( $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ), known as white vitriol, is used in making solutions for electric batteries.

**Magnesium.**—Magnesium occurs in several minerals, the most important of which are *dolomite* ( $\text{CaCO}_3 \cdot \text{MgCO}_3$ ) and *carnallite* ( $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ ). The metal is prepared by the electrolysis of fused carnallite. Magnesium is a light, silvery-white metal which is easily tarnished by moist air. In the ribbon or powder form, it burns with a brilliant white light, forming white magnesium oxide ( $\text{MgO}$ ). The metal is used chiefly in making fireworks, flashlight powders, and light alloys, such as magnalium ( $\text{Mg} + \text{Al}$ ).

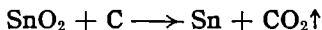
*Magnesium sulfate* ( $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ ), which is the common Epsom salts, and *magnesium citrate* are used medicinally. *Magnesium chloride* ( $\text{MgCl}_2$ ) is present in small amounts in table salt. Being deliquescent, it readily absorbs moisture from the air, thus causing the salt to "cake" when the weather is damp.

**Mercury.**—Mercury has the unique distinction of being the only metal that is liquid at ordinary temperatures. Its principal ore is *cinnabar* (mercuric sulfide,  $\text{HgS}$ ). When this ore is roasted, the sulfur is converted to sulfur dioxide, liberating the mercury, which is driven off as a vapor and is condensed:



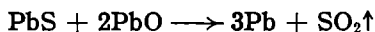
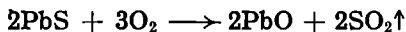
Mercury is a heavy, silvery-white metal which is inert in air. Because of its characteristic properties, it is particularly well-suited for use in thermometers, barometers, and other scientific instruments. The alloys of mercury are known as *amalgams*, the most important of which are silver amalgam, employed in dentistry, and gold amalgam, formed during the metallurgy of gold. Mercury forms two series of compounds: *mercurous*, in which the metal is univalent; and *mercuric*, in which it is divalent. *Mercurous chloride* (*calomel*,  $\text{HgCl}$ ) is a white, insoluble powder used for medical purposes. *Mercuric chloride* ( $\text{HgCl}_2$ ), commonly known as bichloride of mercury or corrosive sublimate, is used as an antiseptic.

**Tin.**—The principal ore of tin is *cassiterite* (stannic oxide,  $\text{SnO}_2$ ). The metal is obtained by reduction with carbon:



Tin is a silvery-white metal which is practically unaffected by air. Being inert to ordinary corrosive agents, tin finds its widest use as a protective material for other metals. Tinware is made by first cleaning sheet iron with dilute sulfuric acid, and then dipping the metal into a bath of molten tin. The ordinary tin can is the most common example of this type of protection. Block tin, which is tin in practically pure form, is employed in making pipes for dispensing carbonated beverages, and tin foil for wrapping articles of food. Tin is a constituent of many important alloys, some of the best known being bronze, pewter, solder, Babbitt metal, and type metal.

**Lead.**—The chief ore of lead is *galena* (lead sulfide,  $\text{PbS}$ ). The metal is extracted by first roasting the ore until two-thirds of the sulfide is converted to the oxide. The heating is then continued at a higher temperature with no air present, thus causing the remaining sulfide to reduce the oxide previously formed:



Lead, the heaviest of the common metals, is bluish-white in color, and is unusually soft and malleable. When freshly cut, it has a bright luster, but soon takes on a dull hue, owing to the formation of a self-protective coating. Lead is readily attacked by nitric acid and acetic acid, but is only slightly affected by hydrochloric acid or dilute sulfuric acid. Its principal uses are in making lead plates for storage batteries, and white lead for paints. Lead is also used in the manufacture of water pipes, alloys and roofing materials; and in sheathing underground cables and lining tanks.

*White lead*, the most important compound of lead, is a basic carbonate with the formula  $\text{Pb}(\text{OH})_2 \cdot 2\text{PbCO}_3$ . In the Dutch process of preparing this compound, a series of reactions takes place slowly, involving acetic acid, lead, and carbon dioxide produced by the fermentation of tanbark. The process takes from three to four months. The product is a heavy, white powder which is valuable as a base in the manufacture of paint. A complete

paint consists of the following: a *base*, usually white lead; a *vehicle*, usually linseed oil; a *thinner*, such as turpentine; a *chemical dryer*, as lead oxide; and frequently a *pigment*, to give the paint color.

**Nickel.**—Nickel is a hard, silvery-white metal which is not oxidized by dry air. For this reason it is used extensively for protecting other metals. Nickel-plating is carried out by an electrolytic process. The article to be plated (cathode) and a bar of pure nickel (anode) are suspended in a solution of some nickel salt. Nickel is of great practical importance because of its many useful alloys, the more common of which are listed in the Appendix.

**Bismuth.**—Bismuth occurs in both the native and the combined state. It is a lustrous, brittle metal with a rather low melting point. Its chief use is in the making of alloys that melt easily. One of the best known is *Wood's metal*, which melts at 60°C. This alloy is used in automatic sprinklers, electric fuses, and safety plugs for preventing boiler explosions.

**Antimony.**—Antimony is an amphoteric element, that is, it has certain definite properties which belong to metals, but at times it may act like a non-metal. It is a silvery, brittle solid, having a crystalline structure. Antimony expands when it solidifies, a property which makes it indispensable in the preparation of type metal, which is an alloy of lead, tin, and antimony. *Babbitt metal* (Sb, Pb, Sn, Cu) is used for machine bearings because it helps to reduce friction.

**Arsenic.**—Arsenic is another element with amphoteric properties. It is a steel-gray, crystalline solid which burns in air, forming arsenic trioxide. The addition of a small amount of arsenic to the lead used in making shot causes the metal to harden quickly. The poisonous nature of all compounds of arsenic renders them valuable in making insecticides.

**Manganese.**—Manganese occurs as *pyrolusite* ( $MnO_2$ ), from which it is extracted by the thermit process (page 236). It is a hard, steel-like metal which is tarnished by moist air. Its chief use is in making the iron alloys ferromanganese, spiegeleisen, and manganese steel. An important compound of manganese is potassium permanganate ( $KMnO_4$ ), which is used extensively in the laboratory as an oxidizing agent.

**Chromium.**—Chromium is obtained from its oxide ore by the thermit process. Like manganese, it is a hard metal resembling

steel in appearance. Chromium is used to make certain special steel alloys. Recently, methods have been perfected for electroplating with chromium. Objects, such as automobile headlights and radiators, that are thus plated present a brilliantly lustrous surface which is unusually resistant to corrosion. Stainless steel, containing from 11 to 15% chromium, is widely used for knife blades. Chromium forms several series of salts, in which the element possesses valences of from 2 to 6.

**Silver.**—Silver, the most common of the precious metals, occurs in the free state, and combined with sulfur as *silver glance* ( $\text{Ag}_2\text{S}$ ). Frequently, the silver ore contains gold, copper, lead, and other metals. These metals are separated from the silver in the refining processes, which are both chemical and electrolytic in nature.

Silver is a soft, white metal having a high luster. It is the most ductile and malleable metal next to gold, and surpasses all metals as a conductor of heat and electricity. Silver is inert to oxygen. It tarnishes readily because of the action of sulfur compounds which are present in the air and in certain foods, such as eggs and mustard. The black deposit of silver sulfide ( $\text{Ag}_2\text{S}$ ) may be conveniently removed by placing the article in an aluminum pan together with a hot solution of washing soda. A weak electric current is generated in this simple voltaic cell, and the hydrogen which is evolved in the reaction reduces the silver sulfide to metallic silver.

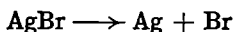
Silver is used as a coating on silver-plated ware, and in coins and jewelry. So-called "solid" silver contains a certain percentage of copper to increase its hardness. A silver twenty-five cent piece contains 90% silver and 10% copper, and is said to be 900 fine (900 parts of silver in 1000). Sterling silver is 925 fine.

The most important compound of silver is silver nitrate (*lunar caustic*,  $\text{AgNO}_3$ ). It is used to prepare other silver compounds, and to cauterize wounds. The silver halides ( $\text{AgCl}$ ,  $\text{AgBr}$ , and  $\text{AgI}$ ) are all decomposed by the action of light. Enormous quantities of these compounds, particularly the bromide, are used each year in photography.

**Photography.**—A little over a century ago the first photograph involving the use of a silver salt was made. Today the industry has grown to giant proportions, embracing the commercial, amateur, and motion-picture fields, and consumes 150 tons of silver annually. The essential steps in the production of a finished picture may be summarized as follows:



1. *Exposing*.—The film or plate, consisting of a strip of cellulose acetate or a sheet of glass, is first coated with an emulsion of silver bromide (AgBr) and gelatin. This sensitized medium is placed in the back of the camera. Light rays from the object being photographed pass through the lens, forming an image on the film. Wherever the light rays strike, decomposition of the silver bromide takes place, the degree of decomposition at any one spot varying with the intensity of the light. The reaction may be represented as follows:



2. *Developing*.—After the film has been exposed, it bears no visible evidence of the above decomposition. It is now dipped into a chemical solution called a *developer*. This further reduces the silver salt which has been exposed to the light, precipitating black, finely divided silver. This results in the formation of a visible image. The portions of the film that have not been exposed to light are unaffected by the developer.

3. *Fixing*.—In order to make the image permanent, the film is treated with some chemical, usually sodium thiosulfate (*hypo*,  $\text{Na}_2\text{S}_2\text{O}_3$ ), which dissolves away all the unexposed silver bromide. The film is then washed thoroughly. The final product is called a *negative*, because it is an accurate picture of the object photographed with all the shades of light and dark reversed.

4. *Printing*.—To obtain a positive, with the areas of light and dark in their proper relationship, the negative is placed on print paper, which has a sensitized coating similar to the original film. A strong light is then directed on the negative. The silver salt on the paper is thus subjected to varying degrees of decomposition, depending on how much light each portion of the negative allows to pass through. The print is then developed and fixed as above, yielding a faithful image of the object photographed. Frequently the finished product is further improved by a process of *toning*. The print is immersed in a solution of a gold, copper, or platinum salt, which causes the silver of the image to be replaced by the particular metal being used. The result is a print of pleasing color and quality.

**Gold**.—Gold is found largely in the free state in the form of dust or nuggets, and in gold-bearing quartz. One method of extraction is to pulverize the ore and allow the gold to amalgamate with

mercury. The latter is then driven off by heat. Another method is to treat the ore with a solution of sodium cyanide, which dissolves the gold. The metal is then extracted either by electrolysis or by replacement with zinc. Gold is also obtained as a by-product, together with silver, in the refining of copper and lead.

Gold is soft, heavy, bright-yellow in color, and is the most malleable and ductile of all the metals. It is inert to air and to all common acids. Aqua regia (concentrated  $\text{HNO}_3$  and  $\text{HCl}$ ) dissolves the metal, forming gold chloride ( $\text{AuCl}_3$ ). Gold, to be used for jewelry or coinage, is alloyed with copper to give it the necessary hardness. The expression "24 carats fine" denotes pure gold. Hence, 18 carat gold means  $18/24$  or 75% pure gold. Jewelry is made chiefly from 14 to 18 carat gold.

**Platinum.**—Platinum occurs alloyed with other rare metals, such as osmium and iridium. About 90% of the world's supply comes from the Ural Mountains in Russia. It is a white metal with a permanent luster,<sup>1</sup> and is harder, heavier, and more expensive than gold. Platinum has a very high melting point, and is unaffected by common reagents, except aqua regia, although it is attacked by fused alkalis. Because of its inertness and high melting point, platinum is useful for contact points in electrical apparatus, and for making crucibles, wires, and similar articles to be used in the chemical laboratory. In the spongy or colloidal form, it is used as a catalytic agent in the industrial preparation of many chemicals, notably sulfuric acid and nitric acid.

**Radium.**—Radium is the costliest substance known to man. It was first obtained by Professor and Madame Curie from pitchblende, an ore of uranium. The most striking fact about radium is that the element itself, as well as all its compounds, undergoes a constant disintegration, which is accompanied by the evolution of three kinds of rays:

1. *Alpha rays* are positively charged nuclei of the helium atom.
2. *Beta rays* are electrons.
3. *Gamma rays* are similar to X-rays, traveling at the same speed as light, and having enormous penetrative powers.

This disintegration is an example of *radioactivity*, also exhibited to a less marked degree by several other rare elements, notably uranium. The phenomenon is believed to result from the breaking down of a complex atom to form a simpler atom. The decomposition of uranium (atomic weight, 238) results, after several stages, in

the formation of radium (atomic weight, 226), which, in turn undergoes disintegration, the final product being lead (atomic weight, 206). Exhaustive investigation of the entire field of radioactivity has done much to build up our modern concept of the structure of the atom.

In addition to its importance in theoretical chemistry, radium has some very practical uses. Its rays render certain substances luminous, hence minute traces of radium are used in paints for watch dials, electric push buttons, etc. These radiations also kill bacteria and other microscopic forms of life. Under carefully controlled conditions, radium rays have been found useful in checking the growth of cancers and tumors of the skin.

### QUESTIONS

1. (a) Starting with the sulfide ore, describe the metallurgy of zinc. (b) Write equations for the reactions involved.

2. (a) Write the equation for (1) burning magnesium in air; (2) dissolving the resulting product in water. (b) What is the effect of this solution on litmus paper? Explain.

3. Give the chemical name, formula, and uses of (a) calomel, (b) corrosive sublimate.

4. (a) What is *tinplate*? (b) How is it made? (c) Account for the rusting of tin cans that have been scratched.

5. (a) Describe the Dutch process of making white lead. (b) Compare zinc white and white lead as paint bases.

6. Give the composition and one use of each of the following alloys: (a) German silver, (b) nichrome, (c) monel metal, (d) perm-alloy.

7. (a) Name the alloy used in automatic sprinklers. (b) Give a reason for the use of this alloy.

8. Antimony alloys expand when they solidify. Name one such alloy and explain how its use depends on this property.

9. Explain why arsenic is used in the manufacture of lead shot.

10. What is the chief use of potassium permanganate in the laboratory?

11. Give three reasons why chromium-plating has come into such widespread use.

12. Name the steps involved in making a finished photograph. Explain each step.

13. Explain why aqua regia is able to dissolve gold.

14. Give two illustrations of the use of platinum as a catalytic agent.

15. Explain how the study of radium and radioactivity has made possible the modern concept of atomic structure.

## COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. Galvanized iron is sheet iron covered with a coating of ..... in order to .....
2. Epsom salts is crystallized ....., the formula of which is .....
3. An amalgam is an alloy of ..... and some other metal.
4. Tin is classed as an amphoteric element because, in its compounds, it behaves either as a ..... or as a .....
5. Storage battery plates are made of .....
6. In the hydrogenation of oils, ..... is used as a catalyst.
7. The most characteristic property of alloys of bismuth is .....
8. Babbitt metal contains the three metals ..... It is used for .....
9. Arsenic is valuable in the preparation of insecticides because .....
10. Spiegeleisen is an alloy of iron and ..... It is used in the ..... process for making steel. Its purpose is to .....
11. Chromium is obtained from its oxide ore by the ..... process.
12. Pure silver is never used for coinage because .....
13. Gold occurs in nature chiefly in the ..... state because .....
14. Platinum has a ..... melting point; hence, a platinum ..... is a useful piece of apparatus in the laboratory.
15. Radium is always found associated with ores of uranium because .....

## MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

- | A                      | B                                   |
|------------------------|-------------------------------------|
| ( ) Galvanized iron    | 1. Zinc in ingot form               |
| ( ) Mercury            | 2. Causes salt to "cake"            |
| ( ) Lunar caustic      | 3. Alloy of Cu, Sn, and Zn          |
| ( ) Galena             | 4. 24 carat                         |
| ( ) Spelter            | 5. To dissolve unexposed AgBr       |
| ( ) Silver             | 6. $ZnSO_4 \cdot 7H_2O$             |
| ( ) Bronze             | 7. Gives off gamma rays             |
| ( ) Fixing             | 8. Chief lead ore                   |
| ( ) White vitriol      | 9. Zinc forms a protecting coat     |
| ( ) Magnesium chloride | 10. $AgNO_3$                        |
|                        | 11. The best conductor of heat      |
|                        | 12. Liquid at ordinary temperatures |

## CHAPTER 24

### FOODS

**Importance.**—Man can get along without many things, but two substances are vitally essential, namely, air and food. Deprive him of either, and life becomes impossible. The science of nutrition deals with such questions as the composition of foodstuffs, the function of each food element in the body, the proper balance of the various types of food constituting a healthful diet, and the work of the organs of digestion and assimilation. Nutrition is one of the most important branches of applied chemistry.

**Nutrients.**—Food is any substance which, when taken into the body, furnishes heat energy by oxidizing, or builds and repairs worn-out tissues. Our food is composed of individual chemical substances known as nutrients. These are five in number: *carbohydrates*, *fats*, *proteins*, *mineral matter*, and *vitamins*. Each nutrient has a specific function in the body.

**Carbohydrates.**—As indicated in Chapter 16, carbohydrates are compounds of carbon, hydrogen, and oxygen, the last two elements being present in the ratio of 2:1, as in water. Starch and the various sugars are the two classes of foods that make up the carbohydrates. They are the chief energy producers of the body. Grains, bread, and potatoes furnish most of our starch, while cane sugar, candy, fruits, and jellies provide for our sugar requirements.

**Fats.**—Fats and oils are, like carbohydrates, compounds of carbon, hydrogen and oxygen, and are likewise energy producers. Fats are stored in the body as reserve foods, being oxidized only in emergencies, when the supply of carbohydrates is nearly exhausted. Butter, lard, cream, various meats, and all the vegetable oils are rich in fat.

**Proteins.**—Proteins are extremely complex substances that contain nitrogen along with carbon, hydrogen, and oxygen. Their primary work in the body is to build tissues, and repair the cells that have been broken down in the normal functioning of the organism. In extreme cases, proteins are oxidized by the body to

furnish energy. This process constitutes starvation and, unless discontinued, will result in death. Proteins are found in such foods as egg white, fish, lean meat, casein of milk, gluten of wheat, beans, and peas.

**Mineral Matter.**—The body requires a small quantity of inorganic salts in order to maintain proper health. Calcium and phosphorus are essential elements in the building of bones and teeth. Iron is important as a constituent of the red blood corpuscles, which transport oxygen to the body cells. Iodine is valuable in maintaining the thyroid gland in normal health. Other essential mineral elements are sulfur, chlorine, magnesium, sodium, and potassium. Egg yolk, milk, green vegetables, fruits, and whole wheat are important sources of mineral matter in our diet.

**Water.**—Water, while it does not furnish energy or build tissues, is an indispensable element in our diet. Its functions are to act as a solvent, to help digest and distribute the nutrients, and to aid in the elimination of waste products from the body.

**Vitamins.**—Up to a comparatively few years ago it was believed that the nutrients mentioned above were all that a healthful diet need contain. At about the turn of the century, experimenters began to realize that foods contain minute quantities of some additional substances, without which the diet is seriously deficient. These mysterious ingredients, whose exact chemical nature is still very little understood, have been termed *vitamins*. Thousands of diet experiments on test animals have resulted in the identification of at least seven such substances and in an accurate determination of their function in maintaining health. The table on page 255 gives the important facts about the better known vitamins.

**Food Tests.**—In the study of foods, we are able to recognize the presence of the various nutrients by the following tests:

1. **Carbohydrates.**—To test a food for *starch*, it is heated with water, then cooled, and treated with a few drops of iodine solution. The resulting color is blue-black. To test for *glucose* or *fructose*, the food is heated with Fehling's solution. A brick-red color results. *Sucrose* or cane sugar does not give this reaction. If, however, the sucrose is first heated with water and a little hydrochloric acid, it is converted to a mixture of glucose and fructose (process of *inversion*). Then, after the solution has been rendered slightly alkaline, it is tested in the usual way with Fehling's solution.

VITAMINS	FUNCTION	FOODS RICH IN VITAMINS
A	Promotes growth. Prevents certain eye diseases.	Milk, butter, cod liver oil, eggs, leafy vegetables.
B	Prevents beri-beri or poly- neuritis.	Peas, tomatoes, asparagus, yeast.
C	Prevents scurvy.	Lemons, oranges, tomatoes.
D	Prevents rickets.	Cod liver oil, food exposed to ultra-violet light.
E	Prevents sterility.	Lettuce, wheat kernels.
F	Prevents nervousness, loss of appetite, paralysis.	Lettuce, liver.
G	Prevents pellagra.	Milk, eggs, meat, yeast.

2. *Proteins*.—The presence of protein in food is shown by treating the food with concentrated nitric acid. A yellow color results. If the food is removed from the acid, washed, and then treated with ammonium hydroxide, it turns orange in color.

3. *Fats*.—To prove the presence of fats or oils, the food is placed on a sheet of glazed paper and warmed. A translucent area ("grease spot") results.

4. *Mineral Matter*.—The food is heated strongly in air (burned) until all volatile matter is driven off. The minerals remain in the form of a gray ash.

5. *Vitamins*.—The presence of vitamins in foods is determined by noting the effect of various diets on certain test animals such as white mice, guinea pigs and rabbits.

**Energy Value of Foods.**—We have seen that the heat and energy of the body result from the oxidation of food. It is therefore important to determine the relative food values of the various substances which make up our diet, on the basis of their capacity as fuels. The unit of heat measurement is the *Calorie*, which is the amount of heat needed to raise the temperature of 1000 grams of water one degree centigrade.

To measure the heat (or *calorific*) value of a food, a weighed quantity is burned in a bomb calorimeter, which is an air-tight metal container filled with oxygen, and surrounded by water. The

# We Are What We Eat

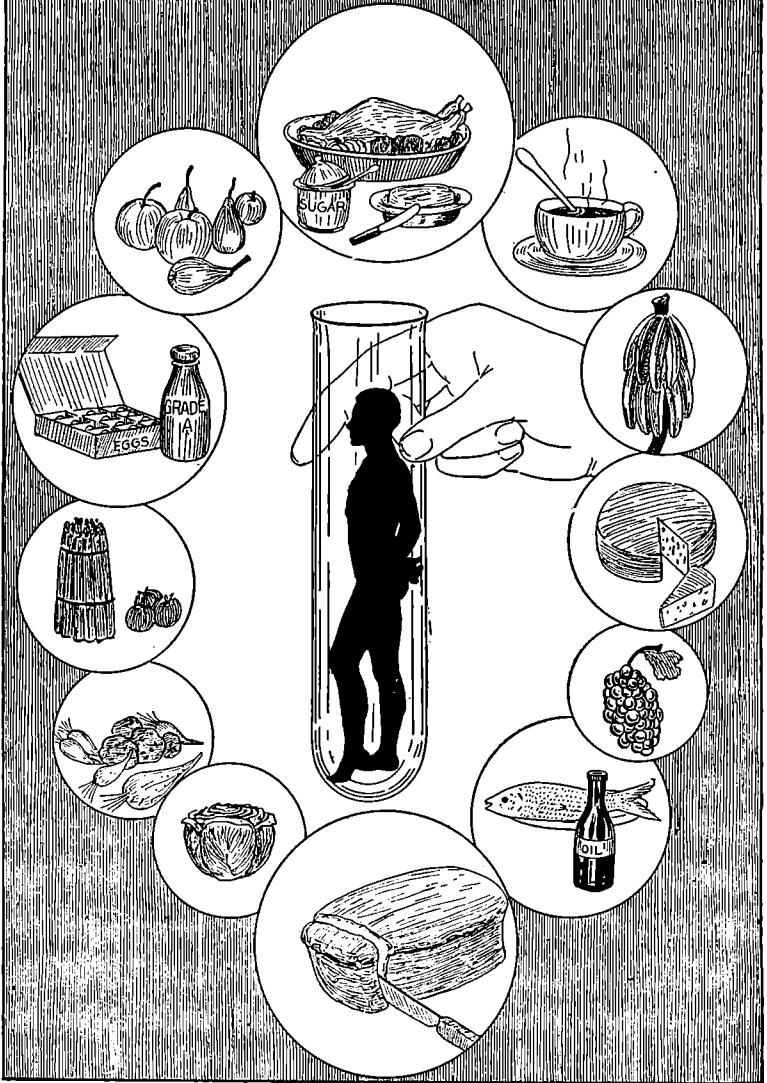


FIG. 72.



heat evolved by the burning food is absorbed by the water, and the increase in the temperature of the water is recorded. A simple calculation is then required to determine the fuel value of the food. Calorimeter tests show that, weight for weight, fats yield more than twice as much heat energy as either carbohydrates or proteins.

**Balancing the Diet.**—Much experimental work has been done in determining the kinds of food and the amount of each necessary to keep the body healthy. It is evident that many factors enter into the question of what a balanced diet should be, such as age, weight, occupation, climatic conditions, and other considerations. The average adult person requires enough food each day to furnish about 3000 Calories of heat energy. The nutrients should be in about the following ratio by weight: carbohydrates—3; fats—1; proteins—1. In addition, the diet must contain a sufficient supply of the minerals and vitamins which have been proved to be so vital to good health. Milk has often been called the “perfect food,” particularly for growing children, since, aside from the water content (87%), it has practically all the essential food elements in the proper proportion for health, and in such a form as to be easily digested and absorbed by the body.

**Food Preservation.**—Most foods spoil easily, owing to the work of minute organisms which get into the food, principally from the air, and cause *fermentation* and *putrefaction*. Various methods are employed to destroy these organisms, prevent them from entering food, or stop them from multiplying. Following are some of the more important procedures:

1. *Cooling.*—Food is kept in a cold place, which prevents the growth of the organisms already present.

2. *Sterilizing.*—Food is raised to a temperature sufficient to destroy the organisms, after which it is sealed in an air-tight container. This is usually done in the process of “canning” foods.

3. *Pasteurizing.*—Milk is pasteurized by heating it to about 65°C., maintaining it at that temperature for thirty minutes, then cooling it to near the freezing point. This destroys most of the bacteria present and checks the growth of those surviving.

4. *Drying.*—Some foods, such as fish and meat, are dried so that there is not enough moisture left for the growth of the micro-organisms.

5. *Use of Preservatives.*—The addition of large amounts of salt or sugar will keep foods from spoiling. Examples are salt fish and preserved fruits. Small amounts of certain chemicals, such as sodium benzoate, are employed to prevent the destructive organisms from multiplying, as in the case of some ketchups.

**Digestion.**—Before the food which we eat can be utilized by the body, it must be so changed as to render it *soluble*. Only in this form can it be absorbed by the blood stream and carried to various parts of the body, there to do its work of furnishing energy or building tissues. The digestive system consists of a chain of organs made up of the mouth, stomach, small intestine, and large intestine, together with certain auxiliary organs like the pancreas, liver, gall bladder, and spleen. Certain of the organs secrete digestive juices which contain complex chemical substances called *enzymes*. The function of these enzymes, which act very much like the catalytic agents previously studied, is to convert the insoluble nutrients present in food into soluble form. *Ptyalin*, occurring in the saliva of the mouth, changes starch to glucose. *Pepsin* in the stomach breaks down proteins into simpler nitrogenous compounds. In the small intestine, *pancreatic juice* and *intestinal juice*, both containing various enzymes, complete the digestion of starch, proteins and fats. Here, also, the process of assimilation takes place, by which the now soluble nutrients pass into the blood stream.

**Cells.**—The cell is the unit of the living organism. Man is composed of millions of cells, and in each of them a multitude of chemical functions are constantly taking place. The sum total of these functions is called "life." The cell absorbs food and oxygen from the blood, complex processes occur within it, and waste products are given back to the blood to be transported to various parts of the body, and there eliminated. The chemist understands very little, as yet, of the chemical nature of *protoplasm*, the living material of the cell, or of the processes by which it is synthesized and by which it is ultimately decomposed.

## QUESTIONS

1. (a) Define *food*. (b) On the basis of your definition, is water a food? Explain.
2. Name five nutrients present in food, and mention two functions of each in the body.
3. Name two foods that contain rather high percentages of each nutrient (ten foods in all).
4. Can a man live solely on bread and water? Explain.
5. Describe the tests which you would make to determine the presence of four nutrients in a cheese sandwich.
6. "An apple a day keeps the doctor away." Is there any truth in this maxim? Explain.
7. Explain the close association of the two foods in such food combinations as bacon and eggs, bread and butter, cereal and cream.
8. (a) Define *Calorie*. (b) Describe the method used to measure the heat value of one gram of butter.
9. (a) What are *vitamins*? (b) Explain the important contribution which the discovery of vitamins has made to the science of nutrition.
10. Name two digestive enzymes, indicate in which organ of the human body each is secreted, and explain the work of each in the process of digestion.

## COMPLETION TEST

*Supply the word or words required to complete each of the following statements.*

1. Carbohydrates and fats are composed of the three elements . . . . . Proteins contain, in addition, the element . . . . .
2. The mineral matter in food is often called "ash" because . . . . .
3. The two mineral elements that aid in the building of bones and teeth are . . . . .
4. The chief mineral element present in the red blood corpuscles is . . . . .
5. The temperature of the human body is higher than that of the surrounding air because . . . . .
6. The function of water in the human body is . . . . .
7. The proper ratio by weight of nutrients in a balanced diet for the average adult should be: carbohydrates . . . . .; fats . . . . .; proteins . . . . .
8. Milk is regarded as an ideal food because . . . . .
9. Vitamin . . . . ., found abundantly in . . . . ., is effective in preventing scurvy, a deficiency disease.
10. Before our body can make use of the foods we eat, they must be converted into a . . . . . form. This process is called . . . . .

## MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

## A

- ( ) Source of proteins
- ( ) Iodine
- ( ) A food preservative
- ( ) Fats
- ( ) Living substance of the cell
- ( ) Bomb calorimeter
- ( ) Pancreatic juice
- ( ) A deficiency disease
- ( ) Present in saliva
- ( ) Putrefaction

## B

- 1. Pellagra
- 2. To measure fuel values
- 3. Renders nutrients soluble
- 4. Stored as reserve food
- 5. Sodium benzoate
- 6. Caused by microorganisms
- 7. Nitric acid
- 8. Egg white
- 9. Vitamin C
- 10. Protoplasm
- 11. Essential for the thyroid gland
- 12. Ptyalin

## CHAPTER 25

### TEXTILES

**Importance.**—Since the remote periods of prehistoric time, man has clothed himself, both for protection against the elements, and for personal adornment. The art of making clothing has advanced a long way since primitive man fashioned rude garments of furs, skins, leaves and grasses. The modern textile industry, which is concerned with the weaving of various fibers into cloth, is based on exact science. The chemical nature of the fibers, their physical properties, their qualities of durability and strength, their ability to absorb and retain dyes, are all important factors in the manufacture of clothing. In recent years, various artificial fibers have been perfected which are, in many respects, superior to nature's products, and which stand as a mighty monument to the achievements of the textile chemist.

**Textile Fibers.**—Although there are hundreds of fabrics which may be used in making clothing, the basic fibers employed are very few in number (Fig. 73). Practically all of them come from vegetable and animal sources, as shown in the following table:

FIBER	ORIGIN
Cotton Linen	Vegetable
Wool Silk	Animal
Rayon	Artificial (vegetable source)
Asbestos	Mineral

**Cotton.**—Cotton is the most important and most extensively used textile fiber. Each year more than half of the world's cotton supply is furnished by the United States. The cotton fiber, under

the microscope, looks like a flat, spiral-shaped thread. Chemically, it is almost pure cellulose, having the formula  $(C_6H_{10}O_5)_y$ . It burns easily, giving off very little odor, and leaving almost no ash. Strong acids attack cotton, but alkalis have little effect. Chlorine bleaches cotton, but weakens the fiber at the same time. *Mercerized cotton* is made by stretching the fiber and dipping it in a strong solution of sodium hydroxide. It thus acquires a smoothness and luster resembling the qualities of silk. In addition, it is stronger than cotton, and can be dyed more easily.

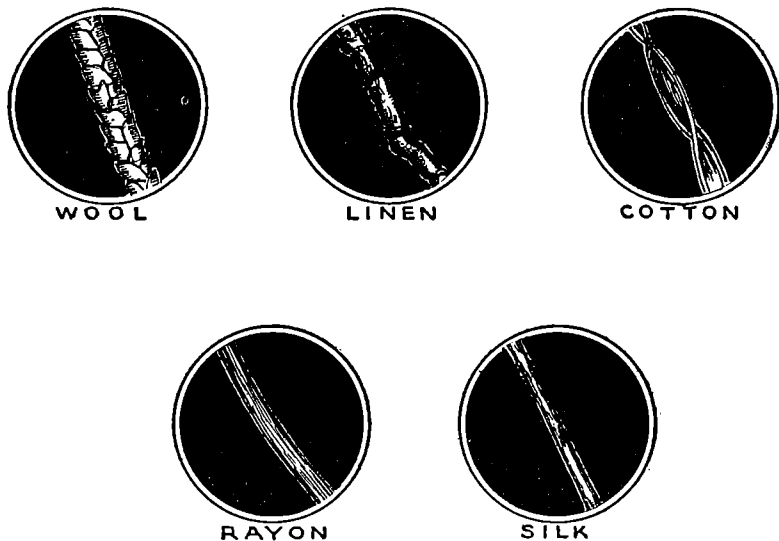


FIG. 73. Some common textile fibers.

**Linen.**—Linen is obtained from the flax plant. The linen fiber is stronger than cotton, and has a better luster, but does not take a dye as easily. Like cotton, linen is nearly pure cellulose, and also burns readily, with almost no odor, and very little ash. Linen is a better heat conductor than any other fiber, and is therefore ideal for summer garments, since it permits the body heat to escape rapidly. The linen fiber absorbs water with great ease, and for this reason is used in making towels.

**Wool.**—Several textile fibers are of animal origin; of these, by far the most important is wool, which we obtain from the sheep.

When examined under the microscope, the wool fiber is seen to be round, with short, overlapping segments that give the fiber a notched appearance. Chemically, wool is a protein compound which contains sulfur, thus resembling hair, feathers, and nails in composition. When ignited, it burns slowly, giving off a disagreeable odor, and leaving much ash. Nitric acid turns wool yellow, a reaction which is sometimes used as a test for wool. A warm solution of sodium hydroxide will dissolve wool completely, which is another way to distinguish it from the vegetable fibers.

Wool is very *hygroscopic*, being able to absorb up to half its own weight of water without showing any evidence of dampness. In addition, wool is a poor conductor of heat, owing to the insulating qualities of the air that is trapped in the meshes of the fibers. These two properties of wool make it ideal for cold-weather clothing.

**Silk.**—Silk, the strongest and most costly of the textile fibers, is, like wool, an animal product. The silkworm (which is really not a worm, but the caterpillar stage in the life cycle of a certain type of moth) expels from its glands a sticky fluid, which hardens into a fine thread. With this thread the silk worm spins its cocoon. The cocoons are first heated to kill the caterpillar, then the fibers are unwound and spun into thread.

The silk fiber is round, smooth and lustrous. Like wool, silk is a protein compound, but does not contain sulfur. Like wool, also, it burns slowly, yielding the same unpleasant odor. Silk is dissolved by concentrated hydrochloric and nitric acids. Salt makes silk tender, which accounts for the harmful effects of perspiration on delicate silk garments. Both wool and silk are dyed with ease.

**Rayon.**—Rayon, an artificial fiber made from cellulose, is the most important silk substitute. The cellulose is obtained chiefly from wood pulp or low-grade cotton products. Of the several methods for making rayon in use at the present time, the *viscose process* is the most important. In this process, the cellulose is first bleached, then mercerized by treatment with sodium hydroxide, and finally dissolved in carbon disulfide. The product is a thick, straw-colored liquid called *viscose*. This is squeezed through tiny holes to produce fibers (compare the action of the silk worm in spinning its cocoon). The fibers then pass into a chemical bath which hardens the viscose into fine threads. After treatment to eliminate the excess chemicals, the threads are dried, and are ready to be woven into cloth.

The rayon fiber, under the microscope, is a glossy, solid thread. Although resembling silk in appearance, it is not as strong, especially when wet. Since rayon is nearly pure cellulose, its chemical properties are very much like those of cotton. Thus, it burns readily, with little odor or ash, reacts like cotton to acids and bases, and takes dyes with difficulty.

**Asbestos.**—Asbestos is a complex magnesium silicate occurring as a fibrous mineral. It is mined in large quantities in Quebec, Canada. Asbestos is woven into cloth to be used in making fire-proof theater curtains and protective suits for firemen.

**Dyes.**—Coloring matter for textiles has been used for thousands of years. Up to comparatively recent times, all dyes were natural products, obtained from plant and animal sources, and from the ground. At present, practically all of our dyes are synthetic substances derived from coal tar, one of the products obtained in the destructive distillation of soft coal. These synthetic, or *aniline*, dyes are far better than the natural dyes because of their greater uniformity, their superior coloring powers, and their stability in the presence of certain destructive agents. In addition, they can be produced at a much lower cost than the natural substances.

A coloring material, in order to be a good dye, must have the following properties:

1. Its color must be acceptable.
2. It must have no harmful effects on the textile fibers being dyed.
3. It must be *fast*, that is, it should stick to the cloth during ordinary washing, and not "run" or "bleed."
4. It must not lose color too readily on exposure to light, air, or other fading agents.

**Dyeing Processes.**—In general there are three methods now in use for dyeing cloth:

1. *Direct Dyeing.*—Many dyes dissolve in water, and color a piece of cloth when it is merely dipped into the resulting solution. The animal fabrics, silk and wool, are particularly suited for direct dyeing. For instance, a piece of silk is colored light-yellow when boiled in a solution of picric acid. The action is believed to be colloidal in nature, the dye being adsorbed by the fibers. Very few dyes will color cotton directly, one of them being *congo red*.



2. *Mordant Dyeing*.—A large number of dyes cannot be used directly on the cloth, but must be employed in conjunction with an assisting material known as a mordant. This auxiliary substance acts as a binder, probably as a result of colloidal adsorption. The mixture of mordant and dye is called a "lake." Aluminum hydroxide and tannic acid are two commonly used mordants, while *alizarin* is an example of a mordant dye. Cotton, one of the most difficult textiles to dye, usually requires a mordant.

3. *Vat Dyeing*.—Some dyes, such as *indigo*, are applied by first impregnating the cloth with a colorless derivative, then oxidizing this derivative to form an insoluble colored substance. Such dyes are unusually fast to both washing and bleaching, and are used extensively for dyeing cotton goods.

### QUESTIONS

1. Name five textile fibers and state the origin of each.
2. Compare the chemical composition of an animal fiber with that of a vegetable fiber.
3. How would you distinguish between cotton and wool by (a) the burning test, (b) a chemical test?
4. Compare cotton and mercerized cotton as to (a) composition, (b) appearance, (c) strength, (d) ability to take a dye.
5. Woolen garments are warmer than cotton garments. Explain why this is so.
6. When ignited, wool and silk give off an odor of burning feathers, while cotton and linen do not. Explain the reason for this.
7. Compare rayon with silk as to (a) origin, (b) composition, (c) appearance, (d) strength, (e) cost, (f) ability to take a dye.
8. (a) What is *viscose*? (b) Outline the viscose process for making rayon.
9. Account for the name "coal-tar dyes."
10. Outline the process of vat dyeing, and explain the chemical action involved.

### COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. Cotton contains the elements . . . . ., while wool contains the elements . . . . .
2. When cotton is stretched and treated with sodium hydroxide solution, it is converted into . . . . .

3. Linen is superior to cotton as a material for towels because . . . . .
4. A textile fiber which can be dissolved by a warm sodium hydroxide solution is . . . . .
5. Cotton . . . . . (is, is not) readily attacked by strong acids.
6. Silk belongs to a class of organic compounds called
7. Three common cellulose fibers are . . . . .
8. Four requirements of a good dye are . . . . .
9. Aluminum hydroxide is used as a . . . . . in dyeing. Its function is to . . . . .
10. Wool and silk are . . . . . (more, less) easily dyed than cotton, linen and rayon.

### MATCHING TEST

*In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.*

- | A  | B                                 |
|--|-----------------------------------|
| ( ) Turns wool yellow                        | 1. A flat spiral thread           |
| ( ) Silk                                     | 2. Salt                           |
| ( ) A mineral fiber                          | 3. A good heat conductor          |
| ( ) Nearly pure cellulose                    | 4. A direct cotton dye            |
| ( ) Linen                                    | 5. Carbon disulfide               |
| ( ) Makes silk tender                        | 6. Asbestos                       |
| ( ) Cotton fiber                             | 7. A vat dye                      |
| ( ) Congo red                                | 8. Rayon                          |
| ( ) A hygroscopic fiber                      | 9. Nitric acid                    |
| ( ) A solvent used in making silk substitute | 10. A "lake"                      |
|  | 11. The most costly textile fiber |
|  | 12. Wool                          |

## CHAPTER 26

### COLLOIDAL STATE OF MATTER

**Importance.**—Colloid chemistry has been aptly described as “the chemistry of bubbles, drops, grains, filaments and films.” From the definition, we gather that this branch of our science deals with matter in a *finely divided condition*. While in this state of fine division, substances exhibit properties that are striking and unusual—properties that are rarely met with under any other conditions.

The applications of colloid chemistry in everyday life are so extremely varied as to be truly astonishing. The baking of bread and the gorgeous colors of sunset are both dependent on colloids. The brewing of beer and the formation of a river delta are made equally clear by a knowledge of colloid chemistry. Important industrial operations such as tanning, dyeing, metallurgy, and a hundred others are based on reactions in the colloidal state. Even life itself is a matter of colloids, for protoplasm, the substance of the living cell, is colloidal in nature, while many of the vital processes of the body are nothing more than colloidal phenomena.

**What Is a Colloid?**—In our study of solutions we observed the differences between a true solution and a suspension. The colloidal state lies midway between these two conditions. The determining factor in each case is the size of the particles of the substance dispersed or distributed throughout the second medium. In a *solution* the particles are extremely *small*; in a *suspension* they are relatively *large*; in a *colloid* they are *intermediate* in size.

Up to very recently, there was believed to be a sharp dividing line between substances that are *crystalloids*, that is, form true solutions, and substances that are *colloids*, and form colloidal suspensions. The modern concept is that any substance may be a crystalloid or a colloid, depending solely on the size of the dispersed particles. Hence, when we now speak of a colloid, we refer not to a particular substance, but to the colloidal state of matter.

**Varieties of Colloids.**—Although the common type of colloid is the dispersion of a solid in a liquid, other kinds are known to exist.

Since there are three states of matter, solid, liquid, and gas, there are nine possible combinations in which one is the dispersed, and one the dispersing, medium. Actually only eight are examples of true colloids. These are illustrated in the table below:

ILLUSTRATIONS OF COLLOIDAL MIXTURES

DISPERSING MEDIUM	DISPERSED SUBSTANCE		
	Solid	Liquid	Gas
Liquid	Starch paste Colloidal gold	Mayonnaise Milk	Foam Froth
Solid	Ruby glass Alloys	Jelly Glue	Bread Pumice
Gas	Smoke Fine dust	Fog Mist	(No colloidal state possible)

**Preparation of Colloids.**—Certain substances, such as gelatin, starch, and glue, are natural colloids. Other substances are converted into the colloidal state by reducing them to the correct condition of fineness. There are two general methods of doing this:

1. *Peptization* is the process of breaking down coarse material into finer particles. This may be accomplished (a) by the use of a special grinding apparatus called a *colloid mill*; (b) by shaking a liquid with an emulsifying agent, as in the making of mayonnaise; and (c) by underwater arcing between metallic electrodes, as in the preparation of colloidal gold.

2. *Coagulation* is the process of combining the molecules of a substance into particles of colloidal size. This is illustrated by the preparation of colloidal silver bromide, in which a solution of potassium bromide is treated with a large excess of silver nitrate.

**Properties of Colloids.**—The colloidal state of matter is characterized by a number of striking properties. Some of these are:

1. *Tyndall Effect.*—If a beam of strong light is projected sideways through a true solution, no illumination is visible within the solution. If the beam is allowed to pass through a colloidal suspension in a liquid which seems to be clear, the colloidal particles reflect the light, thus revealing the path of the beam.

2. *Size of Particles.*—The *ultramicroscope*, which is an ordinary

microscope used in conjunction with the Tyndall cone of light, as described above, enables us to ascertain the size of colloidal particles. The largest approaches in size the particle of solid matter in a suspension. The smallest is about one hundred times the size of a hydrogen atom.

3. *Brownian Movement*.—When viewed through an ultra-microscope, the particles of a colloidal dispersion are seen to be in rapid zigzag motion. This erratic movement is known to be produced by the bombardment of the relatively large and slowly moving particles by the smaller and swifter molecules of the surrounding liquid. This action helps to prevent the colloidal particles from settling.

4. *Color*.—Many colloidal suspensions bear striking colors, which have very little relation to the nature of the dispersed substance. Thus colloidal gold is variously colored blue, green, red, or violet, depending on the size of the colloidal particle. Mere traces of the colloidal substance are sometimes sufficient to yield brilliant colors. A tiny amount of selenium present in glass gives the bright-red color observed in the automobile tail-light.

5. *Electrical Nature*.—Colloidal particles are always electrically charged, some positively, and some negatively. The addition of a substance whose particles bear a charge which is opposite to that of the colloidal particles will cause the latter to precipitate. For example, colloidal gold, which is charged negatively, is precipitated by the positively charged ions of a salt. The degree of precipitation is greater the higher the number of positive charges carried by the metallic ion. Thus the  $Al^{+++}$  ions of  $AlCl_3$  are far more effective in precipitating colloidal gold than are the  $Na^+$  ions of  $NaCl$ . A practical application of the electrical precipitation of colloids is the *Cottrell Process* for preventing smoke. Gaseous products from furnaces are passed through a pipe containing electrodes bearing strong charges. By this means the solid matter is precipitated from the gases, thus eliminating the smoke nuisance, as well as recovering valuable products which otherwise would have been allowed to escape into the air. For example, the process is used very profitably in reclaiming gold, silver, and other metals from smelter smokes.

6. *Dialysis*.—Since colloidal particles are relatively large, they will not go through openings which easily permit the passage of dissolved crystalloid molecules. If a mixture of a colloid and a crystalloid is put into a parchment bag or other animal membrane,

and the latter is placed in water, the dissolved molecules will escape through the paper, leaving the colloidal particles behind. This furnishes an easy method for separating the two.

**Emulsions and Protective Colloids.**—When two non-miscible liquids are shaken together, an emulsion is produced, which is a suspension of one liquid in the other. An example is a mixture of kerosene and water. If this is allowed to stand, the oil will rise to the top. However, the addition of a small amount of soap to the mixture serves to prevent this separation, since the soap forms a thin film about each colloidal particle of oil. In this action, soap is called a *protective colloid*, or an *emulsifying agent*. Similarly, milk, which is largely an emulsion of butter fat in water, does not separate owing to the protective action of another colloid present, namely, *lactalbumin*.

**Adsorption.**—By reason of the fine division in which colloidal particles exist, the total surface presented by these particles is enormous. This may be illustrated by imagining a cube of matter with a certain amount of outer surface exposed. By cutting the cube in half, we expose two fresh surfaces. A further halving produces additional surfaces. By thus continuing the cutting process until the particles are of colloidal size, we have multiplied the total surface tremendously without any increase of the total amount of matter. Owing to this great enlargement of surface, colloids are able to condense large quantities of gases and other substances, since such condensation is strictly a surface phenomenon. This action is called *adsorption*. The removal of poison gases by charcoal, and of the impurities present in sugar by bone-black are examples of adsorption by colloidal substances. Many catalytic actions, notably that of colloidal platinum in the manufacture of sulfuric acid by the contact process, are also examples of adsorption.

**Applications of Colloid Chemistry.**—We have already mentioned some of the practical applications of colloids and colloid chemistry. Following are a number of other fields in which this type of substance is of outstanding importance.

**Photography.**—The photographic film is a mixture of a colloidal silver salt with gelatin. The effect of the exposure to light is the adsorption of some metallic silver by the colloidal salt, thus forming the latent image that is subsequently brought out in the process of developing.

*Clothing.*—All textile fibers are colloidal substances. By a process of adsorption, they are able to hold water and other materials such as dyes. Many of these dyes are colloidal in nature, as are the mordants which help to make the coloring material stick to the cloth.

*Tanning.*—In the preparation of leather, tannic acid and chromium hydroxide are generally employed to coagulate the gelatin in the hide.

*Cement, Mortar and Plaster.*—All these materials harden, owing to the action of colloidal substances. The colloids coagulate into larger masses, which subsequently harden by the absorption of water of crystallization.

*Agriculture.*—The size of the particles of soil has a very direct bearing on its ability to grow crops. Particles of colloidal size appear to be best, owing probably to their greater powers of adsorption. The humus (organic matter) present in the soil must also be of colloidal division if the plant is to make use of it. It has been said that a soil with no colloids present is a sterile soil.

## QUESTIONS

1. (a) Describe two general methods used to convert matter into the colloidal condition. (b) Give an example to illustrate each method.
2. Name three properties exhibited by matter in the colloidal state.
3. What is one explanation for the fact that colloidal particles do not settle on standing?
4. A strong beam of light is passed through a solution of ferric chloride, and then through a colloidal suspension of ferric hydroxide. Explain the difference in the effects produced.
5. Fogs have been successfully dispelled by spraying electrically charged sand from an airplane. Explain.
6. (a) What is a "protective colloid"? (b) Mention two protective colloids.
7. (a) Distinguish between *adsorption* and *absorption*. (b) Which is a colloidal phenomenon? (c) Give two examples of each.
8. Explain the formation of the Mississippi River delta from the standpoint of colloid chemistry.
9. How does the cleansing action of soap depend on colloidal behavior?
10. Explain how colloidal action enters into the following: (a) the curdling of milk; (b) the elimination of the smoke nuisance; (c) the setting of cement; (d) the dyeing of cloth; (e) the refining of sugar.

## . COMPLETION TEST

Supply the word or words required to complete each of the following statements.

1. Colloidal particles are intermediate in size between those present in a . . . . . and in a . . . . .
2. The Brownian movement results from collisions between . . . . . and . . . . .
3. The process of separating a colloid from a crystalloid is called . . . . .
4. Glue and starch are . . . . . (natural, artificial) colloids.
5. The addition of an electrolyte will often precipitate a colloid because . . . . .
6. Finely divided platinum is a . . . . . (more, less) powerful catalytic agent than a solid piece of the metal because . . . . .
7. Charcoal . . . . . gases. Two applications of this are . . . . .
8. Gelatin behaves like a . . . . . colloid in the manufacture of ice cream.
9. The ingredients of milk that form a colloidal suspension are . . . . .
10. Ordinary filtration will remove . . . . . from a liquid but will not remove . . . . . or . . . . . particles.

## MATCHING TEST

In the parenthesis next to each item in column A, write the number of the item in column B which is most closely associated with it.

- | A                                    | B                                       |
|--------------------------------------|---|
| ( ) Molecules unite to form colloids | 1. To study Brownian movement           |
| ( ) Colloid mill                     | 2. Electrical precipitation of colloids |
| ( ) Soap                             | 3. Electric arcing                      |
| ( ) Crystalloid                      | 4. An emulsion                          |
| ( ) To make colloidal gold           | 5. Gas dispersed in a liquid            |
| ( ) Cottrell Process                 | 6. Coagulation                          |
| ( ) Colloidal selenium               | 7. Occurs only on surfaces              |
| ( ) Adsorption                       | 8. Passes through parchment             |
| ( ) Foam                             | 9. In baking bread                      |
| ( ) Ultramicroscope                  | 10. Colors glass bright red             |
|                                      | 11. A protective colloid                |
|                                      | 12. A device for preparing colloids     |



## CHAPTER 27

### FRONTIERS OF CHEMISTRY

**Chemical Research.**—The work of the chemist is never done. In commercial and university laboratories all over the world, men are ceaselessly engaged in searching for new facts, perfecting new processes, and developing new products. No theory concerning atomic structure or chemical behavior is so firmly entrenched that it cannot be altered, enlarged, or, if need be, completely scrapped in the light of newer discoveries. No industrial process is so perfect that it cannot be improved by the use of better materials and more practical techniques. The paragraphs that follow outline briefly some of the more important recent developments in both abstract chemistry and practical technology.

**Structure of the Atomic Nucleus.**—The concept of planetary electrons revolving about a central core is an easy one to visualize, and a satisfying theory in the light of the chemical properties of the atom. The structure of the core or nucleus, however, has always been very much of a mystery. During the last few years the atomic nucleus has been subjected to vigorous bombardment with a variety of projectiles, resulting in the identification of four (possibly five) elementary particles which were hitherto unknown. In addition to the now familiar proton and electron, we have the following:

1. The *neutron* is an electrically neutral particle, made up of a proton and an electron that are closely united to each other. It has a unit mass, and is believed to result from the collapse of a hydrogen atom, during which the single planetary electron has been drawn into the nucleus. Being electrically neutral, the neutron is not subject to the attraction or repulsion of other particles, and therefore possesses enormous penetrative power. For this reason it is admirably suited for use as a "projectile" to bombard the nucleus. Latest research seems to throw doubt on the presence of free electrons in the atomic nucleus. Most authorities now hold that the latter consists of protons and neutrons. When subjected to some unusual disturbance, a neutron is believed to break down into its constituent proton and electron.

2. The *positron* is exactly like the electron in all respects, save that it carries a positive electric charge. The proton is believed by many scientists to be composed of a neutron and a positron. A study of cosmic rays—those mysterious streams of energy that come to us from interstellar space—reveals that they are probably made up largely of positron particles. The life of the positron is extremely short. A collision with an electron results in the transformation of both particles into a *photon*, or bundle of radiant energy.

3. The *neutrino* (or *little neutron*) is an elementary particle whose existence has not yet been completely established. It is believed to resemble the neutron, except that its mass is much smaller. According to the latest theory, the beta rays emitted by a radioactive substance consist of electrons, together with accompanying neutrinos. The presence of such a particle would account for a slight discrepancy in mass sometimes noted when the electrons are traveling at a slower speed than normal.

4. The *deuteron* is the nucleus of deuterium, the heavy isotope of hydrogen discovered by Urey (see below). It is believed to be made up of a proton and a neutron.

5. In May, 1937, the discovery of a *new elementary particle* was announced almost simultaneously by two groups of scientists working independently, one in California and the other in Massachusetts. The new particle, as yet unnamed, has a mass somewhere between that of the proton and the electron, but exhibits far greater penetrative power than the latter.

**Transmutation of Elements.**—The modern physical chemist is an alchemist in the literal sense of the word, for he has succeeded in converting one element into another. By subjecting the atomic nucleus to bombardment with neutrons or protons moving at high speed under enormous electrical voltages, he has been able to disrupt the core of the atom and produce one or more lighter elements. Thus, sodium has been converted into neon, nitrogen into carbon, and aluminum into magnesium. Moreover, the reverse, or *building-up*, process has been demonstrated in the conversion of beryllium (atomic weight, 9) into carbon (atomic weight, 12). In these transformations, the amount of energy required is tremendous, and the quantity of the new element thus produced has been almost infinitesimally small. Yet these results, now of no practical importance, point in the direction of breath-taking achievements still to come.

**The Energy within the Atom.**—Radioactive elements, such as radium and uranium, are constantly giving off energy that is known to result from the disruption of the atomic nucleus. The vast storehouse of energy contained within the atom has been estimated to be many thousand times greater than all of man's present sources of energy, as represented by coal deposits, oil, water power, etc. Will science succeed in unlocking the storehouse? Much brilliant experimental work has already been done in this field, but it appears that, as yet, the surface of the mighty problem has scarcely been scratched. It may be, however, that in the not too distant future our daily power requirements for heat, light, and mechanical operations will be furnished by the energy derived from the disruption of a few atomic nuclei.

**Heavy Hydrogen.**—In 1932, Dr. Harold C. Urey discovered an isotope of hydrogen with a mass of 2, which is present in ordinary hydrogen to the extent of 1 part in 5000. He named this substance *deuterium*. The nucleus of heavy hydrogen, called a *deuteron* (formerly *deuton*), is now regarded as one of the important elementary particles present in the nuclear structure of the atom. In 1934, still another isotope of hydrogen with a mass of 3 was isolated. This has been named *tritium*. It is present in ordinary hydrogen to the minute extent of 1 part in 1,000,000,000.

**Heavy Water.**—A new hydrogen means a new water, and the liquid known as *deuterium oxide* ( $D_2O$ ) is indeed a remarkable substance. Heavy water differs slightly from ordinary water in all of its physical and chemical properties. In its biological effects, it opens up vast new fields of investigation, since all life processes are so closely associated with water. Much research is now going on to determine the effect of heavy water on plants and animals. So far has the work on this new compound progressed, that a factory has recently been put into operation specifically to manufacture heavy water for commercial purposes. The capacity of the plant is about four grams of pure heavy water per week, and the product sells for \$80 per gram.

**Advances in Chemical Industry.**—During the last decade, the industrial chemist has developed hundreds of products and processes which have contributed much to our material comfort and welfare. A new soap, an improved method of electro-plating, a better fertilizer, a process for utilizing waste products from farm and factory, an improved safety glass, a new cellulose derivative—

all of these are advances in chemical technology which we read about almost daily. So numerous and varied are these developments, that to attempt to keep abreast of them is in itself a prodigious task. We can note here very briefly only a few of the representative achievements as an indication of how the chemist is changing the pattern of our daily lives.

**Synthetic Rubber.**—For many years American chemists have struggled with the problem of providing a source of rubber which would render this country (the largest consumer of rubber) independent of foreign sources. Several varieties of artificial rubber have now been perfected. The most recent, and perhaps the most promising, development in the field of man-made rubber is the substance *Neoprene*, which is made from acetylene by *polymerization* (union of simple molecules to form complex structures). Neoprene looks like rubber, acts like rubber, and can be used wherever rubber is used. Indeed, for many purposes Neoprene is actually better than natural rubber, since it is more resistant to attack by gasoline, oils, and oxygen from the air.

**Plastics.**—Plastics include a wide variety of products, both natural and artificial, which may be molded under pressure into any desired form. Perhaps the best known, although by no means the newest, of the synthetic plastics is *bakelite*, made by the reaction between phenol and formaldehyde. More recently, considerable progress has been made in the development of plastics from casein (the nitrogenous solid in milk and cheese) and from albumen. Plastics find a very extensive market in the manufacture of electrical insulators, switches, toilet articles, fountain pens, toys, phonograph records, and hundreds of other common articles.

**Metals.**—Of the ninety-two chemical elements, over seventy are metals. Of these, all except about twenty have hitherto been regarded as not much more than laboratory curiosities. Of late, however, intensive research has been directed toward finding uses for these fifty-odd metals that have been "buried in textbooks" for so many years. It is to the great credit of the metallurgical chemist that he has succeeded in putting many of them to work. The list is a long one, and only a few isolated instances can be mentioned here.

**Beryllium.**—Beryllium is the lightest of all rigid metals. It has been alloyed with many other metals such as iron, copper,

and silver, thereby yielding materials of great lightness, hardness, and resistance to corrosion.

*Cerium*.—Cerium is alloyed with iron to produce a brittle, pyrophoric (*i.e.*, spark-producing) metal. The property of giving off copious sparks when the metal is rubbed with a file makes it useful in the manufacture of pocket lighters and other ignition devices.

*Cesium*.—Cesium, the heaviest of the alkali metals, has a relatively unstable atomic structure, since the single electron in its outermost ring is so far from the nucleus that it is readily set free. Hence the metal is being used for the sensitive coating within the photo-electric cell.

*Iridium*.—Iridium is an extremely hard and resistant metal. It is alloyed with platinum to make fountain pen points, chemical laboratoryware, and standard weights.

*Lithium*.—Lithium, the lightest of the metallic elements, has become valuable in removing impurities from alloys. The addition of a very small amount of lithium has been found to improve iron and copper used for casting.

*Osmium*.—Osmium, the heaviest metal known, finds its principal use in the manufacture of the hard, corrosion-resistant alloys used for tipping gold fountain pens.

*Tantalum*.—Tantalum is an extremely hard metal with a practically perfect resistance to corrosion by acid. It has found extensive use in making a large variety of chemical machinery.

*Zirconium*.—Zirconium finds its chief use in the form of its oxide, *zircite*, which is an excellent refractory, and is employed for lining furnaces.

**The March of Chemistry.**—So overwhelming is the flood of inventions, devices, products, improvements and processes coming from the world's chemical laboratories each year, that merely to list them would require a good-sized volume. Following is a collection of brief items selected from the crowded pages of chemistry's recent achievements:

A satisfactory *synthetic wool* is now being made from the casein of milk. *Vitamin B*, the vitamin that prevents beri-beri, has been produced artificially by chemical methods. The first plant in America to manufacture *ethyl alcohol* for blending with gasoline as a motor fuel has started operations in Kansas. The production of 99.92% pure iron on a commercial scale is now possible, thus

opening up a new field of usefulness for the metal. A *removable printing ink* which contains no lampblack, and therefore can be easily bleached from newsprint paper, has been developed. A new type of *window glass* is now on the market which is transparent to light but holds back 70% of the heat. An *improved explosive*, which is both smokeless and flashless, has been produced by American chemists. A *safe refrigerant*, non-poisonous and non-inflammable, is now made by treating a chlorinated hydrocarbon, such as carbon tetrachloride, with fluorine. A *cellulose sponge* has been developed which holds four times as much water as does the natural sponge, does not sink, and in various other respects is definitely superior to the natural product. A *soapless soap*, made by treating a fatty alcohol with sulfuric acid, is now in use, and is found to work equally well in hard water and in soft water.

**What of Tomorrow?**—The foregoing paragraphs are, of necessity, exceedingly sketchy in character, yet they should be helpful in conveying a vivid picture of this drama of chemical growth and progress. In thousands of laboratories and factories all over the world, our science continues its daily march to greater successes. With all that has been done to advance chemistry to its present high position of achievement, we see ourselves today standing on the threshold of newer wonders and still more astonishing accomplishments in the fields of abstract science and of practical application.

## CHAPTER 28

### CHEMICAL CALCULATIONS

**Importance.**—Chemical arithmetic is a fundamental branch of our science from both the theoretical and the practical standpoints. The chemist, studying the structure of matter and the behavior of atoms and molecules, utilizes weights, volumes, and general measurements that must be made with a high degree of accuracy. The chemist in a commercial laboratory is equally concerned with exact measurements and calculations. He analyzes raw materials, determines the amounts required for a particular reaction, and computes the probable yield of the desired product. Chemical arithmetic is an essential factor in helping to produce a material of the highest possible quality at the lowest possible cost.

The arithmetical problems that are encountered in the study of elementary chemistry may be conveniently grouped into four classes.

#### I. CORRECTION OF GAS VOLUMES

**Need for Correction.**—The volume of a gas may vary between wide limits, depending on the temperature and the pressure. It is therefore essential, in order that our results may be comparable, to measure the volumes of gases under identical conditions, or (more practicably) to *correct* the volumes for a given set of conditions. The ideal set of conditions that have been adopted are known as *standard conditions*. These are: for *temperature*,  $0^{\circ}\text{C}$ . (the freezing point of water); for *pressure*,  $760\text{ mm}$ . (the height of the mercury column in the barometer at sea level).

**Charles' Law.**—A gas expands when it is heated and contracts when it is cooled. Charles, a French chemist, discovered that when a gas, under constant pressure, is heated from  $0^{\circ}\text{C}$ . to  $1^{\circ}\text{C}$ ., it expands  $1/273$  of its volume; and, when cooled from  $0^{\circ}\text{C}$ . to  $-1^{\circ}\text{C}$ ., it contracts  $1/273$  of its volume. He discovered, further, that this gain or loss in volume is the same for each change of

one degree in temperature. He reasoned that, if a gas at 0°C. contracted by 1/273 of its volume for each degree of cooling, it would have zero volume at -273°C. Actually, all gases are converted into liquids before such a low temperature is reached.

Charles designated the temperature of -273°C. as the *absolute zero*, from which he built up another scale of temperature measurement known as the *absolute scale*. Centigrade readings are converted into absolute readings by merely adding 273 algebraically. Charles summarized the relationship between the volume and temperature of a gas in the law which bears his name:

*If the pressure remains constant, the volume of a gas varies directly as the absolute temperature.*

**Application of Charles' Law.**—To determine the new volume of a gas undergoing a change of temperature, first convert the given centigrade temperatures to the absolute readings and then form a *direct* proportion between the volumes and the absolute temperatures:

$$\frac{\text{new volume}}{\text{old volume}} = \frac{\text{new temperature (absolute)}}{\text{old temperature (absolute)}}$$

The method may be simplified by merely multiplying the old volume by a fraction composed of the two absolute readings. An increase of temperature will cause the gas to expand, hence the fraction must be greater than 1. A decrease of temperature will cause the gas to contract, hence the fraction must be less than 1.

**Type Problem.**—The volume of a gas measured at 70°C. is 300 cc. Find its volume at (a) 0°C.; (b) 100°C.

*Solution*

Convert the given centigrade temperature (C.) to the absolute temperature (A.) by adding 273°.

$$70^{\circ}\text{C.} = 70^{\circ} + 273^{\circ} = 343^{\circ}\text{A.}$$

$$0^{\circ}\text{C.} = 0^{\circ} + 273^{\circ} = 273^{\circ}\text{A.}$$

$$100^{\circ}\text{C.} = 100^{\circ} + 273^{\circ} = 373^{\circ}\text{A.}$$

(a) Since the temperature decreases from 70°C. to 0°C., the volume must also decrease. Thus, we have:

$$\text{New volume} = 300 \times \frac{273}{343} = 238.8 \text{ cc.}$$



(b) Since the temperature increases from 70°C. to 100°C., the volume must also increase. Thus, we have:

$$\text{New volume} = 300 \times \frac{373}{343} = 326.2 \text{ cc.}$$

### PROBLEMS

1. A quantity of gas occupies a volume of 75 cc. at 30°C. Find its volume at 0°C. **Ans.** 67.6 cc.

2. When measured at 50° C., a quantity of oxygen has a volume of 15 liters. Find its volume if the temperature is lowered to -10°C. **Ans.** 12.2 liters

3. A certain quantity of gas occupies a volume of 2000 cc. when measured at 100°C. What volume will it occupy at 10°C.? **Ans.** 1517.4 cc.

4. A quantity of nitrogen has a volume of 10 liters at 20°C. Calculate its volume at (a) 40°C.; (b) -20°C. **Ans.** (a) 10.68 liters  
(b) 8.63 liters

5. A gas collected at standard temperature has a volume of 100 cc. Find its volume at 20°C. **Ans.** 107.3 cc.

6. A quantity of hydrogen at 22°C. has a volume of 90 cc. What volume would it occupy at standard temperature? **Ans.** 83.28 cc.

7. With the pressure remaining constant, a liter of gas whose initial temperature is 20°C. is heated until it has expanded to 2 liters. Find the new temperature. **Ans.** 313°C.

8. To what temperature must a gas at standard conditions be heated in order to increase its volume by 50%? **Ans.** 136.5°C.

9. A quantity of gas has a volume of 300 cc. at 136.5°C. What volume will it have at 136.5°A.? **Ans.** 100 cc.

10. A toy balloon holding 2700 cc. of hydrogen at 20°C. shrinks to nine-tenths of its size when taken into a cold-storage room. What is the temperature of the room? **Ans.** -9.3°C.

**Boyle's Law.**—The volume of a gas decreases when the pressure on it is increased, and increases when the pressure is decreased. If a given quantity of a gas has a volume of 100 cc. under a certain pressure, and the pressure is doubled (the temperature remaining unchanged), the volume of the gas is reduced to 50 cc. If, instead, the pressure is halved, the gas expands to a volume of 200 cc.

The seventeenth-century English scientist, Boyle, summarized this relationship in the law which bears his name:

*If the temperature remains the same, the volume of a gas varies inversely as the pressure on it.*

**Application of Boyle's Law.**—To determine the new volume of a gas under a change of pressure, form an *inverse* proportion between the volumes and the pressures:

$$\frac{\text{new volume}}{\text{old volume}} = \frac{\text{old pressure}}{\text{new pressure}}$$

The method may be simplified by merely multiplying the old volume by a fraction composed of the two pressures involved. A decrease of pressure will cause the gas to expand, hence the fraction must be greater than 1. An increase of pressure will cause it to contract, hence the fraction must be less than 1.

**Type Problem.**—The volume of a gas measured at 740 mm. pressure is 200 cc. Calculate the volume (a) at standard pressure (760 mm.); (b) at 610 mm. pressure.

*Solution*

(a) Since the pressure increases from 740 mm. to 760 mm., the volume must decrease. Thus, we have:

$$\text{New volume} = 200 \times \frac{740}{760} = 194.7 \text{ cc.}$$

(b) Since the pressure decreases from 740 mm. to 610 mm., the volume must increase. Thus, we have:

$$\text{New volume} = 200 \times \frac{740}{610} = 242.6 \text{ cc.}$$

### PROBLEMS

1. The volume of a gas is 100 cc. at 780 mm. pressure. Calculate its volume at 680 mm. pressure. **Ans.** 114.7 cc.
2. A quantity of gas measures 10 liters at standard pressure. What would its volume be at 900 mm. pressure? **Ans.** 8.44 liters
3. The volume of a gas is 5 liters at 740 mm. pressure. Find its volume at standard pressure. **Ans.** 4.87 liters

4. Find the volume of a gas at 795 mm. pressure if its volume at standard pressure is 1000 liters. **Ans.** 956 liters

5. A quantity of gas measures 50 cc. at 3 atmospheres of pressure. Calculate its volume at normal atmospheric pressure. **Ans.** 150 cc.

6. The volume of a gas measured under 1 atmosphere of pressure is 100 cc. Find the volume if the pressure is increased to 10 atmospheres. **Ans.** 10 cc.

7. A certain quantity of oxygen is compressed to five-sixths of its original volume. The original pressure was 750 mm. What is the new pressure? **Ans.** 900 mm.

8. A certain quantity of chlorine has a volume of 400 cc. at 740 mm. pressure. What volume would it occupy if the pressure were increased to 770 mm? **Ans.** 384.4 cc.

9. A balloon containing 20,000 cu. ft. of hydrogen at sea level (barometric pressure = 30 in.) rises to a height where the barometer registers 27 in. Calculate the new volume of the gas. **Ans.** 22,222 cu. ft.

10. An automobile tire contains 700 cu. in. of air under 75 lb. pressure. If this air were allowed to escape, what volume would it occupy at ordinary air pressure (15 lb.)? **Ans.** 3500 cu. in.

#### Correction for Changes in Both Temperature and Pressure.—

In order to correct the volume of a gas when changes in both the temperature and the pressure are involved, the original volume is multiplied by two fractions, the first to correct for the change in temperature, the other to correct for the change in pressure.

**Type Problem.**—The volume of a certain gas, when measured at 70°C. and 690 mm. pressure, is 100 cc. Calculate the volume of this gas at standard conditions.

*Solution*

$$70^{\circ}\text{C.} = 70^{\circ} + 273^{\circ} = 343^{\circ}\text{A.}$$

$$0^{\circ}\text{C.} = 273^{\circ}\text{A.}$$

Since the temperature decreases from 70°C. to 0°C., the volume must decrease. Thus, the first multiplying fraction must be less than 1. Since the pressure increases from 690 mm. to 760 mm., the volume must decrease. Thus, the second fraction must also be less than 1. Hence, we have:

$$\text{New volume} = 100 \times \frac{273}{343} \times \frac{690}{760} = \mathbf{72.3 \text{ cc.}}$$

*Solution by Formula*

A simple method involving mere substitution in a formula may be employed to solve this kind of problem. The formula is:

$$\frac{p \times v}{t + 273} = \frac{P \times V}{T + 273}$$

where  $p$ ,  $v$ , and  $t$  are the *old*, and  $P$ ,  $V$ , and  $T$  are the *new* pressures, volumes, and temperatures, respectively. Thus, applying the formula to the above problem, we have:

	OLD	NEW
<i>Volume</i>	$v = 100$ cc.	$V = ?$
<i>Temperature</i>	$t = 70^\circ\text{C}$ .	$T = 0^\circ\text{C}$ .
<i>Pressure</i>	$p = 690$ mm.	$P = 760$ mm.

Substituting these values in the formula, we have:

$$\frac{690 \times 100}{70 + 273} = \frac{760 \times V}{0 + 273}$$

$$V = \frac{690 \times 100 \times 273}{343 \times 760} = 72.3 \text{ cc.}$$

## PROBLEMS

1. A quantity of hydrogen at standard conditions has a volume of 100 cc. Find the volume occupied by the gas at  $-20^\circ\text{C}$ . and 700 mm. pressure. **Ans.** 100.6 cc.
2. A certain gas occupies a volume of 13.2 liters at a temperature of  $40^\circ\text{C}$ . and 727 mm. pressure. Find what its volume would be at standard conditions. **Ans.** 11 liters
3. Given 120 cc. of oxygen at  $25^\circ\text{C}$ . and 770 mm. pressure. Calculate the volume at standard conditions. **Ans.** 111.4 cc.
4. A quantity of nitrogen occupies a volume of 40 cc. at  $50^\circ\text{C}$ . and 740 mm. pressure. Find the volume of the gas at standard conditions. **Ans.** 32.9 cc.
5. A certain gas has a volume of 100 cc. when measured at standard conditions. What would its volume be when measured at  $-20^\circ\text{C}$ . and 700 mm. pressure? **Ans.** 100.6 cc.

6. Given 430 cc. of a gas at 25°C. and 780 mm. pressure. Correct this volume for standard conditions. **Ans.** 404.3 cc.

7. A quantity of gas has a volume of 1 liter at 20°C. and 755 mm. pressure. What will the volume be at 150°C. and 644 mm. pressure? **Ans.** 1.69 liters

8. A quantity of air at a temperature of 50°C. and under a pressure of 1500 mm. is found to have a volume of 120 cc. Calculate the volume under standard conditions. **Ans.** 200.2 cc.

9. A balloon holds 7000 cu. ft. of helium when filled at 20°C. and 30 in. pressure. If the balloon ascends to a height at which the thermometer registers -10°C. and the barometer registers 15 in., what is the new volume? (Assume that the balloon fabric can stretch.) **Ans.** 12,567 cu. ft.

10. Given 1 liter of a gas at a temperature of 20°C. and a pressure of 5 atmospheres. Find the volume of the gas at standard conditions. **Ans.** 4.66 liters.

**Gay-Lussac's Law.**—In the synthesis of water it is known that two volumes of hydrogen unite with one volume of oxygen to form two volumes of steam. Gay-Lussac, a French chemist, first noted the simple volume relationship in this reaction, as well as in similar reactions in which gases are involved. He expressed these facts in the law which bears his name:

*When gases react, the volumes of these gases, and the volumes of the products, if gaseous, are in a ratio to each other that may be expressed by small whole numbers.*

The law may be illustrated by the following cases:

1 vol. hydrogen + 1 vol. chlorine  $\longrightarrow$  2 vols. hydrogen chloride

2 vols. hydrogen + 1 vol. oxygen  $\longrightarrow$  2 vols. steam

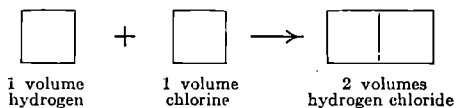
3 vols. hydrogen + 1 vol. nitrogen  $\longrightarrow$  2 vols. ammonia

1 vol. nitrogen + 1 vol. oxygen  $\longrightarrow$  2 vols. nitric oxide

**Avogadro's Law.**—To explain this uniformity, as well as the simple relationship expressed in the other gas laws, Avogadro, an Italian scientist, formulated the following hypothesis, now regarded as a law:

*Equal volumes of all gases, when measured under the same conditions of temperature and pressure, contain the same number of molecules.*

**Number of Atoms in a Molecule of Hydrogen.**—By applying the Laws of Gay-Lussac and Avogadro we may demonstrate that hydrogen (as well as most gaseous elements) has two atoms in its molecule. We know first from experiment that:



Let us make the assumption that one volume of hydrogen contains 100 molecules. It follows therefore, from Avogadro's Law, that one volume of chlorine also contains 100 molecules, and that one volume of hydrogen chloride contains 200 molecules. Hence we have:



Each of these 200 molecules of hydrogen chloride must contain *at least* one atom of hydrogen, making 200 atoms of hydrogen in all. These 200 atoms originate from the 100 molecules of hydrogen with which we started. Hence, *one molecule of hydrogen must contain at least two atoms*, thus making the formula of hydrogen  $H_2$ .

In the above reasoning, the term "at least two" is employed. The hydrogen molecule might conceivably contain more than two atoms but no evidence has ever been found to show that the molecule of hydrogen in any reaction splits up into three, four, or more parts. Therefore, we may conclude that there are two atoms only.

A similar process of reasoning leads to the conclusion that the molecules of chlorine, oxygen, nitrogen and most common gaseous elements contain two atoms each. Ozone is known to have three atoms of oxygen to the molecule. Helium, neon, argon, and the other inert gases of the atmosphere have been proved to be monatomic. Likewise, the metals mercury and zinc, when in vapor form, have only one atom to the molecule.

## II. MOLECULAR WEIGHT

**Molecular Weight from the Formula.**—The molecular weight of an element or compound is the sum of the weights of all the atoms in its molecule. When the formula of a substance is given, the molecular weight is determined simply by adding together the atomic weights of its elements, each weight being taken as many times as the subscript number indicates.

**Type Problem.**—Calculate the molecular weight of (a)  $\text{CaSO}_4$ ; (b)  $\text{Cu}(\text{NO}_3)_2$ ; (c)  $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ .

*Solution*

$$(a) \quad \begin{array}{ccc} \text{Ca} & \text{S} & \text{O}_4 \\ 40 & + 32 & + 16 \times 4 = 136 \end{array}$$

$$(b) \quad \begin{array}{ccc} \text{Cu} & & (\text{NO}_3)_2 \\ 63.6 & + 2(14 & + 16 \times 3) = 187.6 \end{array}$$

$$(c) \quad \begin{array}{cccc} \text{K} & \text{Al} & (\text{SO}_4)_2 & 12\text{H}_2\text{O} \\ 39 & + 27 & + 2(32 & + 16 \times 4) + 12(2 & + 16) = 474 \end{array}$$

## PROBLEMS

(NOTE.—To solve these and subsequent problems, refer to the table of atomic weights on the inside of the back cover.)

*Determine the molecular weights of the following compounds:*

1. $\text{H}_2\text{O}$	Ans. 18	6. $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Ans. 286
2. $\text{KCl}$	Ans. 74.5	7. $\text{C}_6\text{H}_{10}\text{O}_5$	Ans. 162
3. $\text{H}_3\text{PO}_4$	Ans. 98	8. $\text{C}_3\text{H}_5\text{OH}$	Ans. 58
4. $\text{H} \cdot \text{C}_2\text{H}_3\text{O}_2$	Ans. 60	9. $\text{KNaC}_4\text{H}_4\text{O}_6$	Ans. 210
5. $\text{Al}_2(\text{SO}_4)_3$	Ans. 342	10. $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	Ans. 246

**Molecular Weight of a Gas.**—In the case of a gas, the method outlined above for determining the molecular weight cannot always be applied. The reason is that, even if we know the *simplest* formula of a gaseous compound, it may not be the *true* formula. For instance, hydrogen fluoride is shown by analysis to have an equal number of hydrogen and fluorine atoms in its molecule. Hence its formula might be  $\text{HF}$ ,  $\text{H}_2\text{F}_2$ ,  $\text{H}_3\text{F}_3$ , etc.

Actually, the formula of hydrogen fluoride is  $\text{H}_2\text{F}_2$ . We must therefore use other means of finding the molecular weight of a gaseous element or compound. Three common methods are employed, based on (1) vapor density, (2) weight of one liter, and (3) specific gravity.

1. *Vapor Density*.—The vapor density of a gas, which can be determined by experiment, is the ratio of the weight of the gas to the weight of an equal volume of hydrogen:

$$\text{V.D.} = \frac{\text{Weight of a given volume of the gas}}{\text{Weight of an equal volume of hydrogen}}$$

According to Avogadro's Law, the given volumes of the gas and of hydrogen contain the same number of molecules. Thus, we may rewrite the relationship as follows:

$$\begin{aligned} \text{V.D.} &= \frac{\text{Weight of one molecule of the gas}}{\text{Weight of one molecule of hydrogen}} \\ &= \frac{\text{Molecular weight of the gas}}{\text{Molecular weight of hydrogen}} \\ &= \frac{\text{Molecular weight of the gas}}{2} \end{aligned}$$

Hence:

$$\text{Molecular weight of the gas} = 2 \times \text{V.D.}$$

**Type Problem.**—The vapor density of hydrogen chloride is 18.25. Find its molecular weight.

*Solution*

$$\text{Molecular weight} = 2 \times 18.25 = 36.5$$

2. *Weight of One Liter of a Gas*.—Experiment has shown that, under standard conditions, the molecular weight of any gas in grams (*gram-molecular weight*) occupies 22.4 liters (*gram-molecular volume*). Hence:

$$\text{Molecular weight of a gas} = 22.4 \times \text{weight of one liter}$$

**Type Problem.**—The weight of one liter of carbon monoxide is 1.25 gms. Find its molecular weight.

*Solution*

$$\text{Molecular weight} = 22.4 \times 1.25 = 28$$



3. *Specific Gravity.*—The specific gravity of a gas is the ratio of the weight of the gas to the weight of an equal volume of air:

$$\begin{aligned}\text{Specific gravity} &= \frac{\text{Weight of a given volume of the gas}}{\text{Weight of an equal volume of air}} \\ &= \frac{\text{Weight of 22.4 liters of the gas}}{\text{Weight of 22.4 liters of air}}\end{aligned}$$

Weight of 22.4 liters of the gas = Gram-molecular weight of gas  
Weight of 22.4 liters of air = 28.9 gms.

Hence:

$$\text{Specific gravity} = \frac{\text{Gram-molecular weight of the gas}}{28.9}$$

or

Molecular weight of the gas = 28.9 × specific gravity

**Type Problem.**—The specific gravity of methane is 0.554. Find its molecular weight.

*Solution*

$$\text{Molecular weight} = 28.9 \times 0.554 = 16$$

### PROBLEMS

- Acetylene has a vapor density of 13. Find its molecular weight. Ans. 26
- Sulfur dioxide has a vapor density of 32. Find its molecular weight. Ans. 64
- One liter of nitric oxide at standard conditions weighs 1.34 gms. Find its molecular weight. Ans. 30
- The weight of 250 cc. of chlorine at standard conditions is 3.79 gms. Find its molecular weight. Ans. 70.84
- A quantity of hydrogen sulfide weighing 0.076 gm. has a volume of 50 cc. when measured at standard conditions. Find its molecular weight. Ans. 34
- Ten liters of a certain gas weigh 11.6 gms. at standard conditions. What is its molecular weight? Ans. 26
- Three liters of a certain gas weigh 7 gms. Find its molecular weight. Ans. 52.2

8. The specific gravity of bromine vapor is 5.524. Find its molecular weight. **Ans.** 159.6

9. The specific gravity of ammonia is 0.596. Find its molecular weight. **Ans.** 17

10. The specific gravity of helium is 0.137. Find its molecular weight. **Ans.** 4

**Weight of One Liter of a Gas.**—If the formula of a gas is known the weight of a liter may be determined in two ways:

1. *First Method.*—Obtain the molecular weight of the gas from its formula. Since the molecular weight, in grams, of a gas at standard conditions occupies 22.4 liters (gram-molecular volume), we have:

$$\text{Weight of a liter of gas} = \frac{\text{Gram-molecular weight of gas}}{22.4}$$

2. *Second Method.*—Obtain the molecular weight from the formula as before. We know that the molecular weight of a gas is equal to twice its vapor density. Therefore, the vapor density of a gas is one-half its molecular weight. By definition the vapor density of a gas is the ratio of the weight of the gas to the weight of an equal volume of hydrogen. Since a liter of hydrogen at standard conditions weighs 0.09 grams, we have:

$$\text{Weight of a liter of gas} = \text{Vapor density of gas} \times 0.09 \text{ gm.}$$

**Type Problem.**—What is the weight of 1 liter of ammonia at standard conditions?

*Solution 1*

$$\text{Molecular weight of NH}_3 = 17$$

$$\text{Weight of 1 liter} = 17 \div 22.4 = \mathbf{0.76 \text{ gm.}}$$

*Solution 2*

$$\text{Molecular weight of NH}_3 = 17$$

$$\text{Vapor density} = 17 \div 2 = 8.5$$

$$\text{Weight of 1 liter} = 8.5 \times 0.09 = \mathbf{0.765 \text{ gm.}}$$

## PROBLEMS

1. Calculate the weight of 1 liter of each of the following gases at standard conditions:

SO <sub>2</sub>	N <sub>2</sub> O	Cl <sub>2</sub>
C <sub>2</sub> H <sub>2</sub>	C <sub>2</sub> H <sub>6</sub>	CH <sub>4</sub>
HCl	CO	N <sub>2</sub>

(NOTE.—Verify your answers by referring to table in the Appendix.)

2. A certain gas has a vapor density of 19.9. Find the weight of 50 liters of the gas. Ans. 89.6 gms.

3. The weight of 100 cc. of a gas is 0.0178 gm. Find its molecular weight. Ans. 4

4. The vapor density of a gas is 17. Find the weight of 100 liters of the gas. Ans. 153 gms.

5. The vapor density of air is 14.44. Find the weight of 1 liter of air. Ans. 1.29 gms.

## III. PERCENTAGE COMPOSITION AND FORMULA

**Percentage Composition from Formula.**—The per cent by weight of each element present in a compound is obtained by dividing the atomic weight of each element by the molecular weight of the compound, and then multiplying each quotient thus obtained by 100.

**Type Problem 1.**—Find the percentage composition of sulfuric acid (H<sub>2</sub>SO<sub>4</sub>).

*Solution*

$$\text{H}_2 \quad \text{S} \quad \text{O}_4$$

$$1 \times 2 + 32 + 16 \times 4 = 98 = \text{Molecular weight of H}_2\text{SO}_4$$

$$\text{Per cent of H} = \frac{2}{98} \times 100 = 2.0\%$$

$$\text{Per cent of S} = \frac{32}{98} \times 100 = 32.7\%$$

$$\text{Per cent of O} = \frac{64}{98} \times 100 = 65.3\%$$

$$\text{Total} = 100.0\%$$

**Type Problem 2.**—Find the per cent of water in crystallized copper sulfate ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ).

*Solution*

$$\begin{array}{cccc} \text{Cu} & \text{S} & \text{O}_4 & \cdot & 5\text{H}_2\text{O} \\ 63.6 & + & 32 & + & 16 \times 4 & + & 5(2 + 16) = 249.6 = \text{Molecular weight} \\ & & & & & & \text{of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}. \end{array}$$

$$\text{Per cent of H}_2\text{O} = \frac{90}{249.6} \times 100 = \mathbf{36.06\%}$$

### PROBLEMS

1. What is the percentage composition of potassium chlorate ( $\text{KClO}_3$ )?  
**Ans.** K = 31.8%; Cl = 29.0%; O = 39.2%
2. Calculate the percentage composition of copper oxide ( $\text{CuO}$ ).  
**Ans.** Cu = 79.9%; O = 20.1%
3. Find the percentage composition of ammonia ( $\text{NH}_3$ ).  
**Ans.** N = 82.4%; H = 17.6%
4. Calculate the per cent of oxygen in (a)  $\text{HgO}$ , (b)  $\text{H}_2\text{O}$ , (c)  $\text{MnO}_2$ .  
**Ans.** (a) 7.4%; (b) 88.9%; (c) 36.8%
5. What weight of iron can be obtained from 100 tons of  $\text{Fe}_2\text{O}_3$ ?  
**Ans.** 70 tons
6. Find the per cent of water in crystallized sodium carbonate ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ).  
**Ans.** 62.9%
7. Find the percentage composition of phosphoric acid ( $\text{H}_3\text{PO}_4$ ).  
**Ans.** H = 3.1%; P = 31.6%; O = 65.3%
8. Calculate the weight of copper contained in 100 gms. of copper nitrate [ $\text{Cu}(\text{NO}_3)_2$ ].  
**Ans.** 33.9 gms.
9. Find the percentage composition of ethylene ( $\text{C}_2\text{H}_4$ ).  
**Ans.** C = 85.7%; H = 14.3%
10. An ore of zinc contains 70%  $\text{ZnCO}_3$ . Calculate the amount of zinc in 1 ton of this ore.  
**Ans.** 728 lbs.

**Simplest Formula from Percentage Composition.**—If we are given the percentage composition of a compound, we may calculate the relative number of atoms of each element in a molecule of the compound. Thus, we are enabled to write the *simplest* formula of the compound.

**Type Problem.**—A certain compound has the following percentage composition: barium, 58.8%; sulfur, 13.75%; oxygen, 27.45%. Find its simplest formula.

*Solution*

ELEMENT	PERCENTAGE COMPOSITION	DIVIDE BY ATOMIC WEIGHT	QUOTIENT	DIVIDE BY HIGHEST COMMON FACTOR	SUBSCRIPT OF EACH ELEMENT
Ba	58.80	÷ 137	= 0.429	÷ 0.429	= 1
S	13.75	÷ 32	= 0.429	÷ 0.429	= 1
O	27.45	÷ 16	= 1.716	÷ 0.429	= 4

Since the final quotients, 1, 1, and 4, are merely relative figures representing the smallest number of atoms of barium, sulfur and oxygen that may be present in a molecule of barium sulfate, the simplest formula of the compound is  $\text{BaSO}_4$ . The *true* formula may be represented by  $(\text{BaSO}_4)_x$ .

### PROBLEMS

1. A compound has the following composition: carbon, 75%; hydrogen, 25%. Find its simplest formula. **Ans.**  $\text{CH}_4$
2. Determine the simplest formula of a 'compound' that contains 70.13% silver, 9.09% nitrogen, and 20.77% oxygen. **Ans.**  $\text{AgNO}_2$
3. Calculate the simplest formula of a compound having the following composition: calcium, 29.4%; sulfur, 23.56%; oxygen, 47.04%. **Ans.**  $\text{CaSO}_4$
4. A certain compound contains 40% calcium, 12% carbon, and 48% oxygen. Find its simplest formula. **Ans.**  $\text{CaCO}_3$
5. The percentage composition of a substance is: sodium, 32.40%; sulfur, 22.54%; oxygen, 45.07%. Find its simplest formula. **Ans.**  $\text{Na}_2\text{SO}_4$
6. The mineral cryolite has the following percentage composition: sodium, 37.9%; aluminum, 12.9%; fluorine, 54.2%. What is its simplest formula? **Ans.**  $\text{Na}_3\text{AlF}_6$
7. Find the simplest formula of a compound, 1 gm. of which contains 0.2557 gm. of magnesium and 0.7443 gm. of chlorine. **Ans.**  $\text{MgCl}_2$

8. A certain compound has the following percentage composition: sodium, 32.4%; oxygen, 45.1%; phosphorus, 21.8%; hydrogen, 0.7%. Find its simplest formula. **Ans.**  $\text{Na}_2\text{HPO}_4$

9. Find the simplest formula of the compound formed by the union of 70 gms. of nitrogen and 15 gms. of hydrogen. **Ans.**  $\text{NH}_3$

10. What is the simplest formula of the compound formed by the union of 68.5 gms. of barium and 16 gms. of oxygen. **Ans.**  $\text{BaO}_2$

**True Formula of a Compound.**—To determine the *true* formula of a compound, we must know not only its percentage composition, but also its molecular weight. In the case of a gaseous compound, or one that may be easily volatilized, the molecular weight is readily obtained from the experimental determination of the vapor density or the weight of a liter of the gas at standard conditions.

We first determine the simplest formula, as explained above. From this we calculate the apparent molecular weight. The *true* molecular weight will be found to be either the same as, or a multiple of, this apparent molecular weight. If it is the same, then the simplest formula is also the true formula. If it is a multiple, then this multiple represents the number by which the subscript of each element in the simplest formula must be multiplied to give the true formula.

**Type Problem 1.**—A certain compound consists of 92.3% carbon and 7.7% hydrogen. Its vapor density is 39. What is its true formula?

*Solution*

ELEMENT	PERCENTAGE COMPOSITION	DIVIDE BY ATOMIC WEIGHT	QUOTIENT	DIVIDE BY HIGHEST COMMON FACTOR	SUBSCRIPT OF EACH ELEMENT
C	92.3	÷ 12	= 7.7	÷ 7.7	= 1
H	7.7	÷ 1	= 7.7	÷ 7.7	= 1

The *simplest* formula is, therefore, CH.

The *true* molecular weight =  $2 \times \text{V.D.} = 2 \times 39 = 78$ .

The *apparent* molecular weight (from formula CH) = 13.

$$78 \div 13 = 6$$

Hence the *true* formula of the compound is  $\text{C}_6\text{H}_6$ .

**Type Problem 2.**—The percentage composition of a substance is: carbon, 39.95%; hydrogen, 6.69%; oxygen, 53.36%. At standard conditions, 100 cc. of its vapor weigh 0.134 gm. Determine its true formula.

*Solution*

ELEMENT	PERCENTAGE COMPOSITION	DIVIDE BY ATOMIC WEIGHT	QUOTIENT	DIVIDE BY HIGHEST COMMON FACTOR	SUBSCRIPT OF EACH ELEMENT
C	39.95	÷ 12	= 3.33	÷ 3.33	= 1
H	6.69	÷ 1	= 6.69	÷ 3.33	= 2
O	53.36	÷ 16	= 3.33	÷ 3.33	= 1

The *simplest* formula is, therefore,  $\text{CH}_2\text{O}$ .

Since 100 cc. of the gas weigh 0.134 gm., then 1 liter weighs 1.34 gms.

The *true* molecular weight =  $22.4 \times 1.34 = 30$ .

The *apparent* molecular weight (from formula  $\text{CH}_2\text{O}$ ) = 30.

Hence, the *true* formula of the compound is  $\text{CH}_2\text{O}$ .

### PROBLEMS

1. The percentage composition of a gas is: nitrogen 46.6%; oxygen, 53.4%. At standard conditions 100 cc. of this gas weigh 0.134 gm. Determine its true formula. **Ans.** NO

2. A gas consists of 92.3% carbon and 7.7% hydrogen. Its vapor density is 13. Find its true formula. **Ans.**  $\text{C}_2\text{H}_2$

3. A certain gas has the following percentage composition: carbon, 42.86%; oxygen, 57.14%. One liter of the gas weighs 1.25 gms. What is its true formula? **Ans.** CO

4. A gas consists of 85.72% carbon and 14.28% hydrogen. At standard conditions 100 cc. of this gas weigh 0.125 gm. Find the true formula of the gas. **Ans.**  $\text{C}_2\text{H}_4$

5. A gas consists of 27.27% carbon and 72.73% oxygen. Its vapor density is 22. What is its true formula? **Ans.**  $\text{CO}_2$

6. A compound consists of 37.6% carbon; 12.5% hydrogen; and 49.9% oxygen. At standard conditions 100 cc. of its vapor weigh 0.143 gm. Find its true formula. **Ans.**  $\text{CH}_4\text{O}$

7. A certain gas has the following composition: carbon, 46.2%; nitrogen, 53.9%. One liter of this gas at standard conditions weighs 2.32 gms. Find its true formula. **Ans.**  $C_2N_2$

8. A compound consists of 2.74% hydrogen and 97.26% chlorine. Its vapor density is 18.25. What is its true formula? **Ans.**  $HCl$

9. A gas has the following composition: sulfur, 94.12%; hydrogen, 5.88%. At standard conditions 200 cc. of this gas weigh 0.306 gm. Find its true formula. **Ans.**  $H_2S$

10. The percentage composition of a gas is: fluorine, 95%; hydrogen, 5%. Its vapor density is 20. What is its true formula? **Ans.**  $H_2F_2$

#### IV. PROBLEMS INVOLVING EQUATIONS

**Straight-Weight.**—A chemical equation is more than a *qualitative* expression, showing what substances react, and what products are formed. It is also a *quantitative* statement, indicating exact weights of the substances taking part in the reaction. If we assume that a given reaction goes to completion, and if we know the weight of one of the substances involved (either a reactant or a product), we may calculate the weight of any other substance concerned in the reaction.

**Type Problem.**—What weight of oxygen can be obtained by heating 100 gms. of potassium chlorate?

*Solution*

1. Write the balanced equation.	$2KClO_3 \longrightarrow 2KCl + 3O_2$
2. Find the total molecular weights of the substances under consideration.	$K = 39$ $Cl = 35.5 \quad O_2 = 32$ $O_3 = 48 \quad 3O_2 = 96$  $KClO_3 = 122.5$ $2KClO_3 = 245$
3. Write the known and required weights above, and the calculated molecular weights below, the formulas of these substances in the equation.	$100 \text{ gms.} \quad \quad \quad x \text{ gms.}$ $2KClO_3 \longrightarrow 2KCl + 3O_2$ $245 \quad \quad \quad 96$
4. Form the proportion.	$\frac{100}{245} = \frac{x}{96}$
5. Solve for $x$ .	$245x = 9600$ $x = 39.18 \text{ gms. of oxygen}$



## PROBLEMS

1. Find the weight of calcium carbonate formed when 740 gms. of calcium hydroxide are treated with an excess of carbon dioxide.  
**Ans.** 1000 gms.
2. Sufficient sulfuric acid is added to 71 gms. of sodium hydroxide to effect complete neutralization. Calculate the weight of sodium sulfate formed.  
**Ans.** 126 gms.
3. Heat is applied to 50 gms. of crystallized sodium carbonate ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ) until the salt is completely anhydrous. Find the weight of the residue.  
**Ans.** 18.5 gms.
4. Find the weight of hydrogen produced by the action of hydrochloric acid on 150 gms. of zinc.  
**Ans.** 4.6 gms.
5. How many grams of potassium hydroxide are needed to neutralize 100 gms. of sulfuric acid?  
**Ans.** 114.2 gms.
6. What is the maximum weight of iron that can be extracted from 100 tons of ore that contains 85% ferric oxide?  
**Ans.** 59.5 tons
7. Manganese dioxide is treated with an excess of hydrochloric acid, yielding 100 gms. of chlorine. Find the weight of manganese dioxide used.  
**Ans.** 122.5 gms.
8. How many grams of zinc oxide are formed by burning 650 gms. of zinc in an excess of oxygen?  
**Ans.** 810 gms.
9. Find the weight of water that combines with 10 gms. of anhydrous copper sulfate, forming hydrated copper sulfate, the formula of which is  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .  
**Ans.** 5.7 gms.
10. One gram of silver is dissolved in nitric acid, forming silver nitrate. Find the weight of sodium chloride needed to convert all the silver in solution into silver chloride.  
**Ans.** 0.54 gm.

**Weight-Volume and Volume-Weight.**—If we know the weight of one of the substances taking part in a reaction, we may calculate the volume of any gaseous substance either consumed or evolved. Conversely, if we are given the volume of one of the substances, we may calculate the weight of some other substance in the reaction. This is done by forming a proportion to express the relationship between the substances under consideration, as in the case of the straight-weight problems. However, when we are concerned with volumes, we substitute 22.4 liters for each gram-molecular weight of the gaseous substance.

**Type Problem 1 (Weight-Volume).**—How many liters of carbon dioxide can be obtained by heating 500 gms. of calcium carbonate?

*Solution*

1. Write the balanced equation.	$\text{CaCO}_3 \longrightarrow \text{CaO} + \text{CO}_2$
2. Find the molecular weight of the substance.	$\text{Ca} = 40$ $\text{C} = 12$ $\text{O}_3 = 48$
3. Write the known weight above the given substance and the molecular weight below. Write the required volume above the gas, and, below it, 22.4 liters for <i>each molecule</i> of the gas.	$\begin{array}{rcc} & & x \\ & & \text{liters} \\ 500 \text{ gms.} & \longrightarrow & \text{CaO} + \text{CO}_2 \\ 100 & & 22.4 \\ & & \text{liters} \end{array}$
4. Form the proportion.	$\frac{500}{100} = \frac{x}{22.4}$
5. Solve for $x$ .	$100x = 11,200$ $x = 112$ liters of carbon dioxide

**Type Problem 2 (Volume-Weight).**—Find the weight of copper sulfide formed by passing 5 liters of hydrogen sulfide into a solution of copper chloride.

*Solution*

1. Write the balanced equation.	$\text{CuCl}_2 + \text{H}_2\text{S} \longrightarrow \text{CuS} + 2\text{HCl}$
2. Find the molecular weight of the required substance.	$\text{Cu} = 63.6$ $\text{S} = 32$
3. Write the known volume above the given gas, and, below it, 22.4 liters for <i>each molecule</i> of the gas. Write the required weight above the substance and the molecular weight below.	$\begin{array}{rcc} & 5 \text{ liters} & x \text{ gms.} \\ & & \\ 5 & \longrightarrow & \text{CuS} + 2\text{HCl} \\ 22.4 & & 95.6 \\ & & \text{liters} \end{array}$
4. Form the proportion.	$\frac{5}{22.4} = \frac{x}{95.6}$
5. Solve for $x$ .	$22.4x = 478$ $x = 21.3$ gms. of copper sulfide

## PROBLEMS

1. Find the volume of hydrogen evolved when 100 gms. of zinc are dissolved in hydrochloric acid. **Ans.** 34.46 liters
2. What volume of oxygen can be obtained by heating 98 gms. of potassium chlorate? **Ans.** 26.9 liters
3. Find the volume of chlorine that can be obtained by the action of manganese dioxide on 584 gms. of hydrogen chloride. **Ans.** 89.6 liters
4. What volume of hydrogen sulfide can be obtained by the action of hydrochloric acid on 200 gms. of ferrous sulfide? **Ans.** 50.9 liters
5. How many liters of carbon dioxide are formed by the complete burning of 800 gms. of hard coal containing 90% carbon? **Ans.** 1344 liters
6. Find the weight of zinc required to liberate 150 liters of hydrogen by reacting with an excess of sulfuric acid. **Ans.** 435.3 gms.
7. Calculate the weight of sodium chloride required to produce 6 liters of hydrogen chloride by reacting with concentrated sulfuric acid. **Ans.** 15.7 gms.
8. What weight of water must be electrolyzed to obtain 5 liters of oxygen? **Ans.** 8.04 gms.
9. What weight of sodium chloride would be produced if 10 liters of chlorine reacted completely with sodium? **Ans.** 52.2 gms.
10. Find the weight of potassium chlorate needed to produce 100 liters of oxygen. **Ans.** 364.6 gms.

**Straight-Volume.**—Avogadro's Law states that equal volumes of gases under the same conditions of temperature and pressure contain the same number of molecules. It follows, therefore, that the same number of molecules, under like conditions, occupy equal volumes. Hence, if the equation for a reaction involving gases represents these gases in *molecular* form, then we may conclude that the relation between the coefficients is identical with the volume relation between the gases. In solving straight-volume problems, we are not concerned with molecular weights, but merely with these coefficients, since they stand for the volumes of the respective gases.

**Type Problem.**—How many liters of oxygen are needed to burn completely 50 liters of acetylene?

*Solution*

1. Write the balanced equation.	$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$
2. Write the known and required volumes above the gases involved. Write the relative volumes below the gases, as indicated by the coefficients in the equation.	$\begin{array}{ccccccc} 50 & & x & & & & \\ \text{liters} & & \text{liters} & & & & \\ 2C_2H_2 + 5O_2 & \longrightarrow & 4CO_2 + 2H_2O & & & & \\ 2 \text{ vol.} & & 5 \text{ vol.} & & & & \end{array}$
3. Form the proportion.	$\frac{50}{2} = \frac{x}{5}$
4. Solve for $x$ .	$\begin{array}{l} 2x = 250 \\ x = 125 \text{ liters of oxygen} \end{array}$

NOTE.—Frequently a problem of this type calls for the volume of air required to burn completely a given volume of a gas. If we regard the air as one-fifth oxygen, we may obtain the volume of air required by multiplying the calculated volume of oxygen by 5. In the above type problem, the answer, in terms of air, would be  $125 \times 5$ , or 625 liters of air.

### PROBLEMS

1. What volume of oxygen is required for the complete combustion of 200 cu. ft. of hydrogen? **Ans.** 100 cu. ft.
2. What volume of sulfur dioxide can be oxidized to sulfur trioxide by 100 cu. ft. of air? (Assume that air is one-fifth oxygen by volume.) **Ans.** 40 cu. ft.
3. What volume of acetylene can be completely burned by the use of 300 liters of oxygen? **Ans.** 120 liters
4. How many liters of carbon dioxide are formed by the complete combustion of 50 liters of methane ( $CH_4$ )? **Ans.** 50 liters
5. How many liters of hydrogen are required to convert 30 liters of nitrogen into ammonia? **Ans.** 90 liters

6. Find the volume of oxygen necessary for the complete combustion of 100 cu. ft. of water gas containing 40% carbon monoxide and 60% hydrogen by volume.     **Ans.** 50 liters

7. Find the volume of carbon dioxide resulting from the complete combustion of 75 liters of carbon monoxide.

**Ans.** 75 liters

8. How many liters of sulfur dioxide can be obtained by the complete oxidation of 50 liters of hydrogen sulfide?

**Ans.** 50 liters

9. A sample of gasoline is half hexane ( $C_6H_{14}$ ) and half heptane ( $C_7H_{16}$ ). Find the volume of air required to burn completely 100 cu. ft. of the vapor obtained from this gasoline.

**Ans.** 5125 cu. ft.

10. (a) What volume of air (one-fifth oxygen) must be used to form 100 cu. ft. of  $NO_2$  by oxidation of NO? (b) How much NO must be used?

**Ans.** (a) 250 cu. ft.; (b) 100 cu. ft.

## CHAPTER 29

### CHEMISTRY IN WARFARE

#### EXPLOSIVES

**What Is An Explosive?**—Practically every explosive, from the earliest gunpowder of the ancient Chinese to the deadly T.N.T. of the modern shell or aerial bomb, contains some compound of nitrogen. Explosives derive their destructive force from the fact that nitrogen compounds, as a class, are very unstable. If such a compound is exposed to heat or subjected to a physical shock, an extremely rapid chemical change takes place. This results in the evolution of large quantities of gas. The great heat produced by the chemical action tends to increase further the volume of the gaseous products. The energy of this violent expansion is what we call an explosion.

In general, there are two ways in which explosives are used in warfare. The energy of an explosion may drive a projectile, such as a shell, through a gun; in this case, the explosive is called a *propellent*. An explosive may also be designed to shatter violently the container in which it is packed. Such an explosive is known as a *bursting charge*.

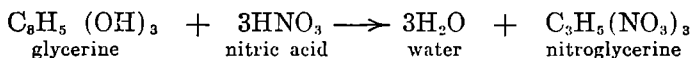
**Gunpowder.**—The ancient Chinese are credited with the invention of explosive mixtures that may be regarded as the first types of gunpowder ever used by man. These mixtures were used in rockets and other varieties of firecrackers. Gunpowder, as we know it today, is said to have been invented by Roger Bacon, a Franciscan monk, in the thirteenth century. This explosive, first employed in warfare at the Battle of Crécy in 1346, continued to be used as a propellent in military weapons up to the beginning of the twentieth century.

Gunpowder is made by mixing three parts of potassium nitrate (saltpeter), one part of powdered charcoal, and one part of powdered sulfur. When gunpowder is ignited, several complicated reactions take place, resulting in the liberation of free nitrogen, carbon dioxide, sulfur dioxide, and other gases. The amount of oxygen contained in the potassium nitrate is insufficient to bring about the complete burning of the combustible ingredients that are present in the gunpowder. Consequently, an explosion of gun-

powder always results in a large quantity of black smoke, made up of unburnt carbon, together with certain solid products of combustion.

**Smokeless Powders.**—Gunpowder is no longer used as a propellant for military purposes. It has been supplanted by more efficient explosives known as *smokeless powders*. There are a number of different types of smokeless powders now being employed in all kinds of guns, from pistols to the huge 16-inch artillery pieces that guard our coasts. Most of these explosives are not powders at all, but consist of relatively large grains, cylindrical in shape and containing many perforations.

**Nitroglycerine.**—One of the most important steps in the development of smokeless powder was the invention of the explosive nitroglycerine in 1846. This is an oily liquid which is made by allowing a cold mixture of concentrated nitric acid and sulfuric acid to act on glycerine; the latter is an abundant by-product in the manufacture of soap. The reaction is:



The function of the sulfuric acid is to act as a dehydrating agent—that is, to remove the water which is formed in the reaction.

**Dynamite.**—Nitroglycerine is extremely unstable and therefore explodes very easily. A substance of this type, which must be handled with great care, is not practical for military or industrial use as an explosive. The Swedish chemist Alfred Nobel succeeded in converting nitroglycerine into a form in which it can be used safely and conveniently. This product, which Nobel called *dynamite*, is made by soaking nitroglycerine into some absorbent material, such as sawdust or infusorial earth. In this solid form, the explosive can be shaped into sticks, and is much less sensitive to shock than the original liquid. Dynamite finds its greatest use as a blasting material in such peace-time activities as mining and building construction.

**Nitrocellulose.**—One of the most important of the smokeless powders is *nitrocellulose* or *guncotton*, which is made by the action of concentrated nitric acid on cellulose. Although cotton is the chief source of cellulose used for this purpose, purified wood pulp has also been employed with some success. As in the manufacture of nitroglycerine, the nitration (*i.e.*, the addition

of nitro groups to a compound) is brought about by a mixture of nitric and sulfuric acids. The function of the latter substance is to remove the water formed during the process.

Nitrocellulose looks very much like ordinary cotton, although it feels somewhat coarse to the touch. When a loose wad of it is ignited, it burns readily, but with no explosive violence. When, however, it is set off in a confined area (such as a gun) by means of a special percussion cap, it explodes with great force, yielding gaseous products that consist mainly of nitrogen, water vapor, carbon monoxide and carbon dioxide. The explosion takes place much more rapidly, and therefore much more forcefully, than in the case of gunpowder. The reason for this is that nitrocellulose is made up of organic compounds in which the oxygen supply (consisting of nitro or  $\text{NO}_2$  groups) is closely united with the combustible carbon and hydrogen in the rest of the molecule. When such a compound explodes, the atoms of fuel are so intimately tied up with the atoms of oxygen that burning takes place throughout the entire charge at almost the same instant. The evolution of gases, whose expansion supplies the energy of the explosion, is correspondingly rapid. The same is true of all the other smokeless powders. In gunpowder, on the other hand, the combustible materials are merely mixed mechanically with the substance that supplies the oxygen (potassium nitrate). Since it takes an appreciable time for the molecules of oxygen to "seek out" and unite with the molecules of fuel, the explosion is comparatively slow.

In this country, nitrocellulose is not used by itself for military purposes, but is mixed with acetone, or with a combination of alcohol and ether. This yields a plastic mass, which may be molded into rods of various lengths and thicknesses for use in different types of guns.

In England, nitrocellulose is mixed with nitroglycerine and petroleum jelly, producing a powerful solid explosive known as *cordite*.

**High Explosives.**—The term *high explosive* is usually applied to a nitrogen compound (or mixture of nitrogen compounds) which explodes, or *detonates*, with far greater speed and force than the smokeless powders mentioned above. High explosives are not used as propellents, but rather as bursting charges in shells, aerial bombs, torpedoes, mines and hand grenades. When the charge in a shell or bomb is set off, each cubic inch of the explosive is converted with indescribable rapidity into about 1000



cubic inches of gas, measured at normal temperature. The great heat generated by the explosion increases the potential gas volume to at least ten times that figure. Thus, at the instant of detonation, the charge seeks to expand to about 10,000 times its original volume, and creates within the container a pressure which has been estimated at 500 tons per square inch. This explains the terrific destructive power of high explosives. Their power is so great that if an attempt were made to use them as propellents, they would literally tear the gun or cannon apart.

**Trinitrotoluene (T.N.T.).**—Perhaps the most important and widely used high explosive for filling shells and bombs is trinitrotoluene. Many military experts regard T.N.T. as practically an ideal explosive. It is made by the action of nitric acid on toluene (or toluol), which is one of the products derived from coal tar. T.N.T. is a yellowish solid having a low melting point. It is remarkably stable, even under rough handling. When T.N.T. is detonated by the use of a suitable percussion cap, it explodes with terrific violence, evolving a heavy black smoke.

**Picric Acid.**—A bursting charge almost as effective as T.N.T. is picric acid, also called *Lyddite*. The substance has been known for almost three centuries, and for many years it was employed as a yellow dye. Picric acid is made by the action of nitric and sulfuric acids on phenol, which is another product obtained from coal tar.

**Dunnite.**—Dunnite, or ammonium picrate, is a high explosive developed by the United States Army during the First World War. While it has the same explosive strength as picric acid, it is much less sensitive to friction and shock.

**Amatol.**—Ammonium nitrate is a high explosive which is never used alone but is mixed with some organic substance containing nitro groups. One of the most effective and widely used of these mixtures is amatol, which consists of ammonium nitrate and T.N.T. The mixture may contain from 50% to 80% of the latter substance.

## PROJECTILES

**Cartridges.**—The projectiles which are fired by side-arms (pistols), rifles and most machine guns are known as *cartridges*. A cartridge consists essentially of three parts. (1) The *percussion cap*, at the back of the cartridge, is a container holding fulminate

of mercury. This compound, formed by the interaction of mercury, nitric acid, and alcohol, is extremely unstable. Even a slight jar is often sufficient to produce a violent decomposition. When the firing pin of the gun strikes the percussion cap, it ignites the fulminate and explodes the powder charge. (2) The *powder charge* now used universally in cartridges is smokeless powder, which has been described above. (3) The smokeless powder serves as a propellant for the *bullet* proper, which is the head of the cartridge. This consists of a thin jacket of a copper-nickel alloy, which is filled with lead. The bullet is often pointed at the tip, but so-called armor-piercing bullets are blunt-nosed.

Note that cartridges do not contain a bursting charge. They are supposed to destroy as a result of the physical impact with which they strike and penetrate their target. One exception to this is *tracer bullets*, which are filled with a pyrotechnic mixture. This mixture begins to burn in the barrel of the gun, and leaves a visible trail as it passes through the air. Such bullets are also used to ignite inflammable targets.

**Shells.**—The first primitive cannon used in the Middle Ages fired large stones, up to 750 pounds in weight. Then came crude iron cannon balls, which developed by a long, gradual process to the modern high-explosive shells. Not until the twentieth century did the development of steel alloys make it possible to fashion artillery shells capable of carrying high explosive charges which are fairly large in relation to the weight of the container.

**Structure of Shells.**—As mentioned above, modern high explosives are remarkably stable to ordinary jolts and shocks. Some of these materials can even be melted without doing any damage. To set off the explosive, most shells are provided with a special device consisting of a detonating cap and a booster.

The *detonating cap*, also called *percussion cap* or *primer*, is located at the nose of the projectile. Like the corresponding part of the cartridge, it contains a small amount of mercury fulminate. When the shell hits its target, the detonating cap is struck a sharp blow by a plunger, which carries the striking or firing pin. This causes the mercury fulminate to explode, and the shock is transmitted to the remainder of the charge in the projectile.

The *booster* is an explosive material that steps up the force produced by the exploding mercury fulminate. The booster is usually contained in a tube which extends from the percussion cap down into the main bursting charge that fills the hollow

shell. The most commonly used booster is tetryl (trinitromethylnitramine). This compound would make an excellent bursting charge itself if it were not for its extreme sensitivity; there is always a danger that it will explode prematurely.

**The "Explosive Train."**—The series of actions which result in the detonation of a shell are known as the "explosive train." There are three main steps involved.

Every shell is fitted with a safety pin, which is so arranged that it prevents the striker from prematurely setting off the fulminate of mercury. This safety pin is removed, either by the centrifugal force of the shell's spinning motion, or by the shock which the shell receives when it is fired. The explosion of the propellant powder in the gun shoots the shell into space. When the nose of the shell strikes its target, the "explosive train" begins. *First*, the striker is driven into the fulminate of mercury, which explodes violently. *Second*, the shock of this explosion is transmitted to the tetryl booster charge packed in the tube immediately below the percussion cap. *Third*, the tetryl explodes with sufficient force to detonate the bursting charge.

**The Time Fuse.**—Not all shells explode as a result of the impact of hitting their target. Some shells are designed to explode automatically within a certain period of time after leaving the gun. Such shells are fitted with a time fuse. This device consists of three rings near the nose of the shell. The upper and lower rings, which are fixed, contain circular channels filled with gunpowder. The central ring, which is adjustable, contains an opening that connects the channels of the other two rings.

When the propellant charge is fired and the shell is shot out of the gun, the safety pin is automatically removed; the force of the explosion drives the striker into the charge of powder and starts it burning. The combustion takes place rather slowly along the winding channel of the timing rings. Since the central ring (the "delay ring") can be adjusted in advance to vary the length of the channel, the time that it takes for the burning to reach the third ring may be regulated with a fair degree of accuracy. When the combustion reaches the end of the train of powder in the bottom ring, it explodes the detonator. This fires the booster, and the latter finally sets off the main charge. The explosion may take place in mid-air.

A variation of the time fuse is the "delayed action" fuse. The purpose of this device is to allow the shell to penetrate the target

before exploding. The general structure is the same as that of the time fuse, except that the striker does not set off the train of powder until the shell hits the objective.

**Shells Used in Warfare Today.**—The high explosive shells employed in today's warfare are of several distinct types, each intended for a particular use.

1. *Demolition shells* are large and heavy. They contain a considerable amount of explosive, up to 60% of the total weight of the projectile. Demolition shells explode with tremendous force, which is directed down into the ground as well as upward and outward from the point of detonation. They are used to destroy buildings, forts, bridges and other large objects.

2. *Armor-piercing shells* are designed to penetrate the steel sides and decks of battleships and explode within. These shells weigh 1000 pounds or more, and are made of a special steel alloy. They carry a relatively small bursting charge.

3. *Fragmentation (or brisance) shells* are designed to scatter a large number of steel fragments over as wide an area as possible. The shell itself is comparatively light and fragile in construction, and the charge of explosive which it carries is relatively small. The bursting force of a fragmentation shell is for the most part in a horizontal direction. These shells are used mainly against concentrations of troops on the fighting line and on the march.

4. *Shrapnel shells* are a specialized type of fragmentation shells. The steel cylinder is filled with half-inch balls of lead-antimony alloy. These balls are held in place by some cementing material such as resin. The explosive charge is located in the base of the shell. At the nose of the cylinder is a time fuse, connected by a central tube to the bursting charge in the base. The fuse starts to function at the instant the shell is discharged from the gun. The main charge is set off when the shell has reached a desired point. The effect of the explosion is to spray the balls fan-wise over a wide area.

5. *Anti-aircraft shells*, which have become such an effective method of combating raiders from the sky, are modified shrapnel shells. The time interval, and consequently the height, at which the "AA shells" will explode may be adjusted before firing.

6. *Grenades* are important weapons for close fighting. The ordinary type of hand grenade, commonly known as a "pine-apple," is actually a small fragmentation shell. It is made of cast

iron, with successive grooves running up and down, as well as around. Just before the soldier throws the grenade, he removes a pin which holds down the safety handle. The soldier's hand, however, continues to hold this handle in place until the grenade is thrown. Then the handle snaps open and activates a fuse, which sets off the main charge. The grenade explodes about five seconds after the handle has been released. The explosion tears the iron case into about forty fragments, which are sprayed over an area up to 200 feet in width. For this reason, the "pineapple" is known officially as a *40-fragmentation bomb*.

7. *Chemical shells* include all types of incendiary, smoke and gas projectiles. In construction they are essentially the same as high explosive shells. The bursting charge, however, is relatively small in quantity, being sufficient merely to rupture the container and scatter the particular chemical agent which is enclosed.

### AERIAL BOMBS

**Aerial Warfare.**—The airplane has assumed a position of primary importance in modern warfare. Planes today are armed with machine guns and cannon, but their principal offensive weapon is the bomb. As you know, troops and fortifications are not the only targets of aerial bombs. They have also been dropped on cities in order to destroy industrial plants, power stations, railroad terminals, docks and other centers of production and transportation. Another purpose of bombing cities may be to disrupt essential civilian services and to cause a general breakdown of morale among the population.

**Types of Bombs.**—In general the bombs that are dropped from airplanes resemble closely the shells fired from guns, except that the outer case of a bomb is provided with fins to steady the missile in its downward journey. *Fragmentation bombs* vary in weight from 25 to 200 pounds. They are rarely employed in attacks on civilian centers. *Demolition bombs*, which are the type most commonly used in air assaults against cities, may range in weight from 100 pounds to 4000 pounds. Frequently they are equipped with delayed-action fuses which bring about the detonation some time after the bomb has struck; this period may be anywhere from a few minutes to 100 hours. *Aerial mines* are massive bodies, weighing between 1000 pounds and 5000 pounds, of which the bursting charge comprises 90%. An aerial mine is

equipped with a parachute, and descends at the rate of about 45 miles an hour. It detonates the instant it touches the ground, causing wholesale demolition over a wide area.

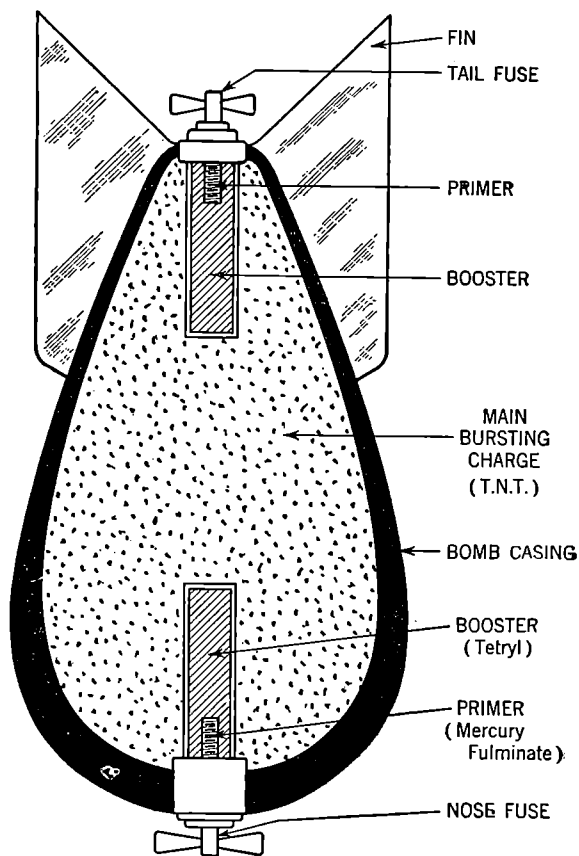
**Bomb Fuses.**—The firing mechanism of a bomb is essentially similar to that of the fuse in a shell but embodies certain additional refinements. The bomb fuse is usually equipped with a small wind vane or propeller which spins as the bomb falls, thus unlocking the safety catch. This action resembles the cocking of a rifle. In a demolition bomb, there are usually two fuses, one in the nose and the other in the tail; the purpose of this is to make sure that the bomb will be detonated, even if one of the fuses fails to work. The nose fuse works by impact, as in the shell. The tail fuse operates by the inertia of a firing pin plunger that continues to move forward when the bomb strikes a solid surface.

**Bombing Raids on Our Cities.**—In attempting to bomb one of our great coastal cities such as New York or San Francisco, the enemy would not be likely to use heavy demolition bombs. This does not seem practical because of the great distances which would have to be flown to and from the objective, and the limited number of such bombs that can be carried even by the largest type of bombing plane in use today. It is the belief of military experts that any attack on our great cities would be more in the nature of a "token" assault. The chief object of such a bombing foray would be not so much to cause really thorough and widespread destruction as to disrupt the normal industrial routine, to create fear and panic, and in general to undermine the morale of the populace. For such "token" bombings, the enemy would be likely to use some fragmentation bombs, together with a large number of small and medium sized demolition bombs, and above all, veritable showers of fire-producing chemicals.

**Incendiary Bombs.**—The modern fire bomb is a relatively small missile, weighing only about two pounds. An ordinary bomber can carry literally thousands of them, and can scatter them at random over a wide area. When we consider that out of the 800,000 buildings in New York City, only 12,000 are fire-proof, it may be seen how terribly destructive and demoralizing an incendiary raid can be.

The most successful type of incendiary (known in Europe as an *electron bomb*) consists of a thick-walled tube of magnesium alloy, filled with thermit, and supplied with a suitable fuse to

ignite the mixture when the bomb strikes. Frequently there is also a small charge of T.N.T. that explodes on impact and scatters the flaming materials about. Incendiary bombs are often packed in a container, holding twenty-five to thirty-six individual units.



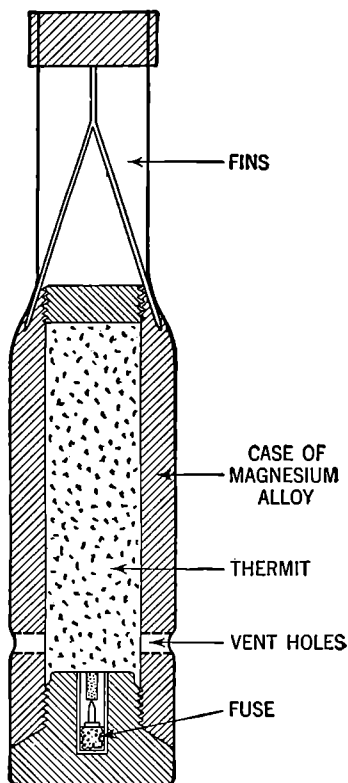
**A Demolition Bomb.**

This is sometimes called a "Molotov bread basket" or a "mother bomb." The container is released from the plane and bursts somewhere in its descent, thus scattering the bombs over a wide area.

**Combating Incendiaries.**—Thermit is a mixture of powdered aluminum and iron oxide. When ignited, it undergoes a chemical action that produces a temperature of nearly 5000° F. The magnesium catches fire and burns at a temperature of 2300° F. During the first minute of combustion, jets of flame are emitted

from the vent holes. The bomb continues to burn for approximately fifteen minutes, scattering molten iron and magnesium about. The metal case itself is finally consumed, so that no portion of the bomb is wasted.

The best method of combating an incendiary bomb is with a *fine spray of water, never with a solid stream*. The effect of the latter is to scatter white-hot iron and magnesium over a wider area. The spray, on the other hand, has the effect of hastening the combustion, so that the bomb burns itself out in only about five minutes, instead of fifteen. This is because the molten magnesium is able to extract oxygen from the spray of water, thus causing an acceleration of the burning. At the same time a fine spray serves to douse the surroundings, and thereby helps to confine the fire to a limited area. The ordinary chemical soda-acid fire extinguisher may also be used, provided some means is taken to convert the stream of liquid into a spray. The carbon tetrachlo-



An Incendiary Bomb.

ride (Pyrene) type of extinguisher should not be used against incendiaries. When this chemical strikes hot metal it is converted into a vapor similar to chloroform, which may have an anesthetic effect. There is also danger that the toxic gas phosgene may be generated.



**Other Types of Incendiary Bombs.**—The all-thermit incendiary bomb has also been used, although it is not as common as the thermit-magnesium type. Such a bomb cannot be easily extinguished, because it carries its own oxygen supply in the aluminum-iron oxide mixture. It must be allowed to burn itself out, while means are taken to localize the fire and protect the surrounding area. Another type of fire-bomb contains thermit and oil. The heat of the burning thermit quickly vaporizes the oil and ignites it with a tremendous burst of flame. The blazing oil fans out in all directions and rapidly spreads the fire. The best method of fighting such a blaze is with a spray of water applied over the surface, beginning at one end and gradually "wiping off" the flame. The spray is converted into steam, which has the effect of diluting the oxygen of the air, as well as cooling the oil below its kindling temperature.

The phosphorus bomb has been used in aerial warfare with some success. One extremely simple type (which has been facetiously termed a "calling card") consists of a thin layer of phosphorus held between two sheets of wet nitrocellulose. The "cards" are left damp until they are unloaded from the plane. As soon as the water evaporates, the phosphorus oxidizes and bursts spontaneously into flame. Another incendiary of this type consists of a shell containing a thermit igniter, a small charge of high explosive and a quantity of white phosphorus. The latter substance is scattered about when the bomb strikes. Water easily extinguishes such a phosphorus blaze. After this is accomplished, great care must be taken to clean up all pieces of unburnt phosphorus, even the tiniest particles, because they will burst into flame again as soon as they become dry.

**Effects of Bombs on Urban Centers.**—The bombing of London and other large urban centers has served as a vast and tragic "laboratory" for testing what happens when bombs are dropped in crowded cities. A large amount of data has been collected, and we can now predict with fair accuracy the probable extent of the damage when bombs of a given type fall within a city, either in residential sections or in areas where modern office buildings and other similar structures predominate.

When a demolition bomb explodes, the air pressure in the vicinity is increased tremendously. The air rushes away violently from the point of detonation, creating a partial vacuum. This partial vacuum lasts longer and frequently is more destructive

than the blast itself. A person caught within the area of reduced pressure may have the air drawn violently from his lungs. Windows may be shattered outward, and in many instances buildings may topple toward the spot where the bomb exploded.

If a medium-sized (300-pound) bomb should make a direct hit within a typical residential block in one of our cities, it might be expected to destroy three adjacent buildings; this would apply either to frame dwellings, or to brick apartment houses of the usual five- or six-story type. If the bomb struck in the street or in the rear yards of such a block, only adjacent walls would be demolished. A bomb falling in a suburban area, where the buildings are spaced relatively far apart, would probably wreck only one house, but would cause substantial damage to the walls of the two nearby houses.

**Bombs and Skyscrapers.**—The modern American office building, with its steel and concrete construction, can be expected to stand up under considerable punishment from air bombardment. Should a medium-sized bomb make a direct hit on the roof of such a building, it would probably penetrate three or four floors. The blasting effect of the bomb, when it exploded, would be dampened somewhat by the cubicle type of construction in such a building. The resulting damage should therefore extend over a floor area not greater than fifty feet in diameter. A bomb striking a glancing blow on the side of a skyscraper would probably demolish a considerable portion of the masonry and perhaps tear out and loosen some of the girders. However, it is unlikely that the steel framework of the structure as a whole would be seriously weakened.

It is not outside the realm of possibility that a bomb might score a direct hit upon the top of an elevator shaft and descend through that easy channel to the very interior of the building before it exploded. Such a freak occurrence would, of course, result in much more serious damage, although there is little danger that it would topple the building over, or wreck it completely. The sturdy steel construction of our American skyscrapers is believed to be more than a match for even the largest demolition bomb so far devised.

**CIVILIAN DEFENSE AGAINST AIR RAIDS**

**Air Defense, Active and Passive.**—The best way to defend a city against attack from the air is to prevent hostile planes from reaching the objective and dropping their deadly load. All measures designed to accomplish this result may be referred to as "active air defense." These active measures include destruction of enemy aircraft on the ground, bombardment of hostile airdromes and flying fields, and the use of interceptor and pursuit planes against approaching bombers. In the immediate vicinity of the city being protected, anti-aircraft artillery, searchlights, captive balloon barrages, and other similar means may be employed.

All too often, despite the strenuous efforts made by military defenders, enemy aircraft succeed in reaching their objective and start to drop bombs. Then it becomes imperative that suitable measures of "passive defense" be put into effect. It is the function of this phase of defense to minimize loss of life and injuries among civilians, to prevent unnecessary destruction of property, to check panic and hysteria, to care for the "bombed-out" population, and to bolster public morale. In this country, there is now in operation a vast civilian defense program, designed to achieve all the objectives mentioned above whenever air raids on our cities may occur. This program, which is still taking shape, is extremely intricate. It involves such elements as systems of air-raid alarms, organization of air-raid wardens and auxiliary firemen, plans for evacuating women, children and elderly persons, blackouts, bomb shelters, first aid, mobile kitchens, protection against toxic gases, rescue and repair personnel, etc.

**What to Do During an Air Raid.**—Everyone of us must play his part in the many-sided program of civilian defense. The best way to begin coöperating with this program is to become thoroughly familiar with what you should, and should *not* do, during an air raid. The recommendations and suggestions which have been issued by civilian defense authorities are summarized below. Some of these precautions may seem trivial and unnecessary to you. Remember, however, that this advice is based on scientific knowledge and actual experience gained during air raids in England and elsewhere. Thousands of lives may be saved if civilians will master these simple rules and will put them into effect during air raids.

1. *General Instructions.*—When the alarm sounds, it is important to act swiftly but calmly, and without panic. If you are indoors, stay there. If you are outdoors, get off the streets. Enter the nearest building but don't go into the subways. Most subways are too near the surface to be safe from direct hits. Furthermore, there is always danger from broken water conduits, sewage pipes and gas mains.

If you are driving in a car when an air raid starts, park it quickly and go inside a building. If you are in a bus or street car, it will stop immediately and you should go to the nearest building. If you are in a subway, the train will stop wherever it is, and you will probably have to remain there until the raid is over.

If you can't get to a building, lie flat and give the best possible protection to the back of your head.

Obeys the orders of police, firemen and air-raid wardens. They have been trained in the task of protecting you.

Students who are in school when an air raid occurs will be told what to do. They should obey all instructions swiftly, calmly, and without asking questions.

2. *During the Raid.*—Stay off the streets. There is danger not only from bombs and flying debris, but also from falling slugs of your own soldiers' anti-aircraft fire.

Stay away from windows. A big bomb will shatter the windows for 200 feet around into deadly splinters of flying glass.

Stay away from outside walls. Such walls may collapse even when the interior of the building is safe.

Don't use the telephone unless it is vitally important. The lines must be free for the communications of the defense officials.

Keep your radio turned on and listen for bulletins and other information.

If the raid is at night, make sure that all the lights in your home are turned out or effectively screened. Good blackout curtains can be made from black sateen, oilcloth, or any other opaque material. If you are caught unprepared, you can improvise blackout materials from blankets or quilts. But don't be caught unprepared!

3. *In Large Buildings.*—If you live or work in a large modern building, certain floors probably have been designated as shelters. Go to the nearest one immediately. Walk there; don't use the elevator. The safest floors are the middle ones—the floors between the three highest and the three lowest stories.

The inside rooms or corridors of these middle floors, where there is a minimum of glass, are the safest.

Basements of such buildings are relatively bomb-proof shelters, but may be dangerous in case of broken water or gas mains.

4. *In Small Buildings.*—Buildings up to five or six stories in height, especially old structures, are less safe. But with certain precautions, both in physical preparation and in the conduct of the occupants, they can be made reasonably secure against anything but a direct hit.

Select one interior room, in the strongest part of your building, as your family air-raid shelter.

During the raid, open the windows of this room, so as to minimize concussion breakage and the attending flying-glass menace. Also, the glass can be strengthened in advance by taping the pane, inside and out, or by covering the pane with a fine wire screen.

All gas jets in the house should be shut off during a raid. A bomb concussion may blow out the pilot light on a kitchen stove, causing escape of gas and a possible explosion, particularly if an incendiary bomb strikes the building.

The family air-raid shelter room should be equipped with a strong table. If the family crouches under the table, danger of falling ceilings and plaster is lessened.

Overstuffed chairs can be placed between the table and the window to guard against glass and bomb splinters.

The raid-shelter room should be stocked with drinking water, food, books and magazines, a flashlight, a portable radio, and other articles of this type. Medical supplies adequate to treat at least minor burns and cuts should be kept at hand.

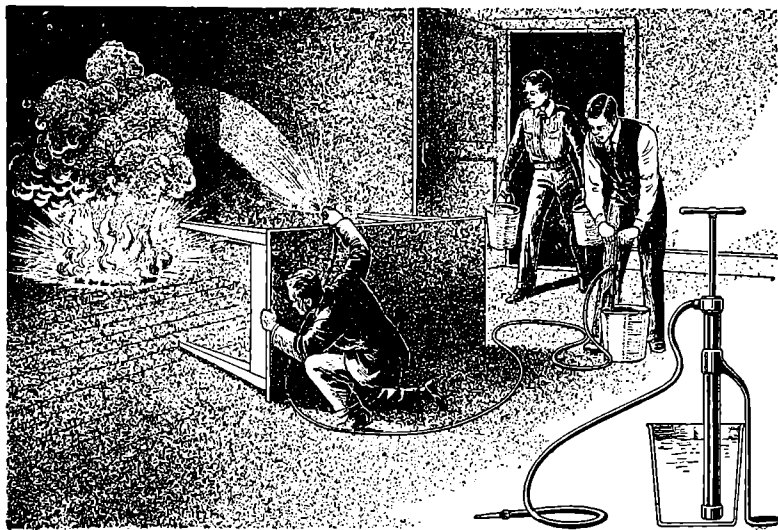
5. *Incendiary Bombs.*—Anybody can equip himself to combat incendiary bombs. Left alone, they burn about fifteen minutes with an intense heat, igniting everything they touch.

If a gentle water spray—not a hard direct jet—is played on an incendiary bomb, it will burn out in about five minutes. One of the most convenient ways of providing such a spray of water is by means of a stirrup pump. The drawing on page 18 shows how the pump is worked. You will notice that it is not unlike a bicycle pump in construction. The cylinder is placed in a container of water, which is pumped by moving the piston up and down. Three persons are needed to handle a stirrup pump properly: one to direct the spray, one to do the pumping, and one to replenish the supply of water. The person directing the spray

should either lie prone under a damp blanket, or stand behind some kind of protective shield; a chair or table can be used for this purpose in an emergency. It is also advisable to wear goggles and heavy leather gloves.

Still another way of handling an incendiary bomb is to scoop it up in a long-handled shovel and drop it in a bucket of sand. Then it can be set down in a place where it will burn itself out harmlessly.

Buckets of water and fire extinguishers should be kept on hand to combat any fires which may be started. If local condi-



Fighting an Incendiary Bomb with a Stirrup Pump.

tions of water supply permit, fill your bathtub beforehand. This may also come in handy as drinking water, or as a source of water for the stirrup pump.

Under war-time conditions, it is more important than ever to avoid the common fire hazards, such as accumulations of oily or paint-covered rags, faulty electrical connections, etc. Check your house at the earliest opportunity and eliminate all such hazards. Particular attention should be given to cleaning out all materials of an inflammable nature from attics, since incendiary bombs strike attics first.

## GAS WARFARE

**Types of Poison Gas.**—More than fifty toxic gases have been employed in modern warfare. They may be divided into several classes, according to their physiological effects. These are: (1) *lung irritants*, which produce injuries to the respiratory passages; (2) *lacrimators*, which bring on excessive tearing of the eyes; (3) *sternutators*, which produce violent spasms of sneezing; and (4) *vesicants*, which cause severe blistering of the skin.

**Lung Irritants.**—*Chlorine*, familiar to anyone who has ever worked in a chemical laboratory, is a greenish yellow gas with a disagreeable, pungent odor. An exposure of two minutes to strong concentrations of this gas is sufficient to cause serious injury or death. When the first chlorine attack was made in the First World War, there were 15,000 casualties, of which 5,000 were fatal. *Phosgene*, or carbonyl chloride, was brought out by German chemists early in the First World War. It has a pungent odor, very much like new-cut hay. It is more than ten times as poisonous as chlorine; one or two breaths of this gas at certain concentrations have been known to produce death in several hours. *Chlorpicrin* is a gas with a sweetish odor, resembling licorice. Besides being a lung irritant, this gas produces tearing and coughing in the victim, as well as severe spasms of nausea and vomiting.

**Lacrimators.**—Tear gases are usually compounds containing bromine, chlorine or iodine. *Bromacetone* was one of the first lacrimators used in warfare. *Chloracetophenone* has a powerful lacrimatory effect. Even one part of this vapor in five million parts of air is sufficient to produce severe tearing. Lacrimators are designed to penetrate gas masks. Their chief use on the battlefield is to make the soldier remove his mask, and thus leave himself exposed to the more toxic gases which may be sent over at the same time.

**Sternutators.**—All the sneezing gases contain arsenic, and are usually non-lethal in the concentrations employed in combat. *Adamsite* and *diphenylchlorarsine* are the best known members of this group of chemical agents. The sternutators, like the tear gases, are designed primarily to cause soldiers such intense discomfort that they will tear off their masks in an effort to obtain relief.

**Vesicants.**—The blistering agents are the most highly efficient of all combat gases used in the last war. *Mustard gas* (dichlorethyl sulfide) has rightly been termed the “king of the battle gases.” This compound is really a dark-brown liquid, slowly changing to a colorless gas with an odor like garlic or horseradish. It forms painful blisters and sores on the skin, resulting in permanent destruction of the tissues with which it comes in contact. When inhaled, it frequently brings on bronchial pneumonia. Mustard gas soaks into the ground, and the vapors may sometimes be given off for weeks. It easily penetrates all ordinary articles of clothing. To be effectively protected against this vesicant, a soldier must be properly masked and must be covered from head to foot with special impermeable materials.

*Lewisite*, a product of our own Chemical Warfare Service, is even more insidious and deadly than mustard gas. It is a dark-brown liquid having an odor like geraniums. Lewisite has all the vesicant and lung-irritant qualities of mustard gas. In addition, it is a powerful systemic poison—that is, it enters the blood stream, paralyzes the central nervous system, and quickly produces death. Lewisite was not perfected until near the very close of the last war, and so was never employed in actual combat.

**Poison Gases and the Present Conflict.**—Poison gases were first used during the last war, although this use violated an international agreement signed by all the combatant nations. During the two decades of peace that followed the Armistice of 1918, chemists of the great powers worked diligently on problems of gas warfare. The results of this research have been kept more or less secret, but its main objectives have become widely known. Some of the more important of these were: to develop a poison of the central nervous system which would bring about serious mental disturbances; to perfect the use of the more deadly fluorine in place of chlorine in toxic gases; to devise highly condensed types of lethal chemicals, so that even extremely low concentrations would prove fatal; to develop toxic substances capable of penetrating all known gas masks.

There have not been any official reports as yet of poison gases being used in the present conflict but the nations involved are constantly on the alert for possible gas attacks. Soldiers are provided with gas masks, and in some countries, particularly England, the civilian population has been similarly equipped, free of



charge. In the United States, measures have already been taken to protect urban centers against aerial gas attacks.

**The Gas Mask.**—The most effective protection for the eyes and lungs against toxic gases is provided by the gas mask. The civilian mask, which is designed for general use, is fairly simple in construction. It consists of a facepiece made of sheet rubber, and provided with a large window of cellulose acetate or some similar material that is both non-inflammable and transparent. The mask is held securely over the face by means of adjustable straps. A metal canister, which is attached directly to the facepiece, contains the substances for purifying the inhaled air. These substances consist of activated charcoal to absorb the poison gases, together with some filtering material to screen out toxic dust and smoke. The inhaled air is made to pass through the canister. A check valve prevents the exhaled air from returning the same way; instead, this air is forced out between the edge of the mask and the cheek of the person wearing it.

The military gas mask is a heavier and more durable type than the civilian mask. The canister, instead of being attached directly to the facepiece, is held in the canvas carrier, and is connected to the mask by a short length of flexible tubing. In addition, the contents of the canister are more highly activated and therefore retain their protective powers for a longer time.

For protection against vesicant gases, which attack any part of the body, special clothing must be worn, in addition to the standard mask. Oilskins and heavy rubber garments, including hood and boots, have been found to provide the best protection against such toxic agents.

## CHAPTER 30

### BRIEF BIOGRAPHIES

It is less than two centuries since the modern science of chemistry emerged from the crude superstitions and semi-magical practices of alchemy. The remarkable subsequent development of the science has been due to the efforts of thousands of individual workers, each of whom has labored to add his contribution—large or small—to the structure of chemical knowledge. Some of these enthusiastic, yet methodical, workers devoted their energies to discovering fundamental truths of a purely theoretical nature. Others were engineers and technicians who made important contributions in the field of practical devices and processes. All of them may be regarded as crusaders in the development of a science which not only has given us a greater understanding of the world in which we live, but has helped us to live those lives in greater comfort, security, and happiness.

Following are thumb-nail sketches of some of the outstanding personalities who have contributed to the growth of modern chemistry:

**ACHESON, EDWARD (1856–1934).**—American. Developed the electric furnace and, with the aid of this device, perfected processes for manufacturing carborundum, graphite, and numerous other products of high-temperature reactions.

**ARRHENIUS, SVANTE (1859–1927).**—Swedish. Formulated the Theory of Electrolytic Dissociation (Ionization Theory).

**ASTON, FRANCIS (1877– ).**—English. Has made important contributions in the field of isotopes.

**AVOGADRO, AMADEO (1776–1856).**—Italian. Formulated the hypothesis which bears his name, and which made possible the correlation between Gay-Lussac's Law of Combining Volumes and Dalton's Atomic Theory.

**BAEKELAND, LEO (1863– ).**—American. Invented *bakelite*, a synthetic resin or plastic.

**BECQUEREL, HENRI (1852–1908).**—French. Discovered the radioactivity of uranium ores and paved the way for the isolation of radium.

- BERZELIUS, JOHANN** (1779–1848).—Swedish. Introduced our modern system of chemical symbols.
- BLACK, JOSEPH** (1728–1799).—Scottish. The first to prepare pure carbon dioxide (by heating magnesium carbonate).
- BOHR, NILS** (1885– ).—Danish. Formulated the concept of planetary electrons revolving in orbits about the nucleus of the atom.
- BOYLE, ROBERT** (1627–1691).—English. Discovered the relation between the volume of a gas and the pressure on it, as expressed in the law which bears his name.
- BUNSEN, ROBERT** (1811–1899).—German. Invented the Bunsen burner. Perfected the spectroscope and introduced methods of spectrum analysis.
- CASTNER, HAMILTON** (1857–1899).—American. Invented the Castner Process for producing sodium hydroxide by the electrolysis of sodium chloride solution.
- CAVENDISH, HENRY** (1731–1810).—English. First to identify hydrogen as an element and to demonstrate that water is composed of hydrogen and oxygen.
- CHARLES, JACQUES** (1746–1823).—French. Discovered the relation between the volume of a gas and its temperature, as expressed in the law which bears his name.
- CURIE, MARIE** (1867–1934).—Polish. Collaborated with her husband, Pierre, in the discovery and isolation of radium, and made other notable contributions to the study of radioactivity.
- DALTON, JOHN** (1766–1844).—English. Formulated the Atomic Theory. Discovered the Laws of Definite and Multiple Proportions.
- DAVY, HUMPHRY** (1778–1829).—English. Perfected the process of electrolysis, and thus discovered sodium and potassium. Invented the miners' safety lamp.
- FARADAY, MICHAEL** (1791–1867).—English. Regarded by many as the most brilliant scientific mind of all times. Liquefied many gases, discovered benzene, and made many noteworthy contributions in the study of electrochemical and electromagnetic phenomena.
- FISCHER, EMIL** (1852–1919).—German. Synthesized a variety of complex organic compounds including glucose, fructose and other sugars, caffeine, and substances closely resembling proteins.
- FRASCH, HERMAN** (1852–1914).—American. Devised the Frasch Method of sulfur extraction, and perfected processes in the field of petroleum refining.
- GAY-LUSSAC, JOSEPH** (1778–1850).—French. Discovered the relation expressed in his Law of Combining Volumes.

- GRAHAM, THOMAS (1805-1869).—English. Regarded as the father of colloid chemistry.
- HABER, FRITZ (1868-1934).—German. Invented the Haber Process for the fixation of atmospheric nitrogen.
- HALL, CHARLES (1863-1914).—American. Invented the process for the extraction of aluminum from its ore by electrolysis, thus founding the American aluminum industry.
- HOPKINS, B. SMITH (1873- ).—American. Discovered the element *illinium*.
- KEKULÉ, FRIEDRICH (1829-1896).—German. Originated the ring method of representing the structure of benzene and its derivatives.
- LANGMUIR, IRVING (1881- ).—American. Winner of the Nobel prize in chemistry (1932) for contributing to the development of the Electron Theory.
- LAVOISIER, ANTOINE (1743-1794).—French. Known as the father of modern chemistry. First to explain the true nature of burning, and to state the Law of Conservation of Matter.
- LEWIS, GILBERT (1875- ).—American. Has made notable contributions to our knowledge of atomic structure.
- LIEBIG, JUSTUS (1803-1873).—German. Founded the modern science of agricultural chemistry.
- MCCOLLUM, ELMER (1879- ).—American. Has pioneered in the field of vitamin research.
- MENDELEJEFF, DMITRI (1834-1907).—Russian. Formulated the Periodic Law, and devised the first practical periodic table of the elements.
- MILLIKAN, ROBERT (1868- ).—American. Has made many important contributions to theoretical science, most noteworthy of which is the isolation and weighing of the electron.
- MOISSAN, HENRI (1852-1907).—French. Perfected the electric arc furnace and investigated many reactions at high temperatures. Made artificial diamonds and was the first to isolate fluorine.
- MORLEY, EDWARD (1838-1923).—American. Determined with great accuracy the relative combining weights of hydrogen and oxygen in the formation of water.
- MOSELEY, HENRY (1887-1915).—English. Studied the X-ray spectra of elements and thus discovered the Law of Atomic Numbers.
- NOBEL, ALFRED (1833-1896).—Swedish. Invented the process of converting nitroglycerine into dynamite. Founded the Nobel prize awards.
- PASTEUR, LOUIS (1822-1895).—French. Studied fermentation processes and the behavior of bacteria and other one-celled organisms.

- PERKIN, WILLIAM (1838-1907).—English. Synthesized the first dye from coal tar, and founded the aniline dye industry.
- PRIESTLEY, JOSEPH (1733-1804).—English. Discovered oxygen and several other gases.
- RAMSAY, WILLIAM (1852-1916).—Scottish. Collaborated with Rayleigh in the discovery of argon and other rare gases of the atmosphere.
- RICHARDS, THEODORE (1868-1928).—American. Winner of the Nobel prize in chemistry (1914) for determining with great accuracy the atomic weights of more than twenty elements.
- RÖNTGEN, WILHELM (1845-1923).—German. Discovered X-rays.
- RUTHERFORD, ERNEST (1871-1937).—English. The first to propose the modern concept of atomic structure. Made outstanding contributions in the field of radioactivity and transmutation of elements.
- SCHÉELE, KARL (1742-1786).—Swedish. The first to prepare oxygen. Discovered chlorine, barium, manganese, tungsten, and many organic compounds.
- SODDY, FREDERICK (1877- ).—English. Collaborated with Rutherford in developing the theory of radioactivity. Discovered isotopes.
- SOLVAY, ERNEST (1838-1922).—Belgian. Invented the Solvay Process for preparing sodium bicarbonate and sodium carbonate from common salt.
- THOMSON, JOSEPH (1856- ).—English. Discovered the electron. Has done outstanding work in the field of isotopes.
- UREY, HAROLD (1893- ).—American. Discovered isotopes of hydrogen and prepared "heavy water." Winner of the Nobel prize in chemistry (1934).
- VAN'T HOFF, JACOBUS (1852-1911).—Dutch. Helped to develop the modern theory of solutions. The first winner of the Nobel prize in chemistry.
- WÖHLER, FRIEDRICH (1800-1882).—German. Prepared urea from inorganic substances, and thus founded the modern concept of organic chemistry.

## WINNERS OF THE NOBEL PRIZE IN CHEMISTRY

YEAR	WINNER	NATIONALITY	FIELD OF ACHIEVEMENT
1901	J. H. Van't Hoff	Dutch	Theory of solution
1902	E. Fischer	German	Chemistry of sugars
1903	S. Arrhenius	Swedish	Ionization Theory
1904	W. Ramsay	English	Rare gases of the air
1905	A. von Baeyer	German	Synthesis of indigo
1906	H. Moissan	French	Discovery of fluorine
1907	E. Buchner	German	Enzymes and fermentation
1908	E. Rutherford	English	Radioactivity
1909	W. Ostwald	German	Fixation of nitrogen
1910	O. Wallach	German	Chemistry of terpenes
1911	Marie Curie	Polish	Discovery of radium
1912	{ V. Grignard	French	Organic synthesis
	{ P. Sabatier	French	Organic catalysis
1913	A. Werner	Swiss	Chemical structure
1914	T. W. Richards	American	Atomic weight determinations
1915	R. Willstätter	German	Chemistry of chlorophyll
1916	(Not awarded)		
1917	(Not awarded)		
1918	F. Haber	German	Synthesis of ammonia
1919	(Not awarded)		
1920	W. Nernst	German	Electrochemistry
1921	F. Soddy	English	Radioactivity
1922	F. W. Aston	English	Discovery of isotopes
1923	F. Pregl	Austrian	Microanalysis
1924	(Not awarded)		
1925	R. Zsigmondy	German	Ultramicroscope
1926	T. Svedberg	Swedish	Chemistry of colloids
1927	H. Wieland	German	Chemistry of bile acids
1928	A. Windaus	German	Chemistry of vitamins
1929	{ A. Harden	English	Fermentation and sugars
	{ von Euler-Chelpin	Swedish	Enzymes and vitamins
1930	H. Fischer	German	Chemistry of hemoglobin
1931	{ C. Bosch	German	Fixation of nitrogen
	{ F. Bergius	German	Hydrogenation of coal
1932	I. Langmuir	American	Electron theory
1933	(Not awarded)		
1934	H. C. Urey	American	Hydrogen isotope and "heavy water"
1935	{ F. Joliot	French	Transmutation of elements by atomic bombardment
	{ Irene Curie (Mmc. Joliot)	French	
1936	P. Debye	Dutch	Structure of the molecule
1937	{ W. N. Haworth	English	Carbohydrates and vitamin C Flavins and vitamins A and G
	{ P. Karrer	Swiss	

**ASSOCIATION TEST**

*Name the person most closely associated with each of the following achievements in chemistry:*

- |                                |  |
|--------------------------------|--|
| 1. Synthesis of ammonia        | 6. Development of spectrum analysis    |
| 2. Law of Multiple Proportions | 7. Introduction of modern symbols      |
| 3. Artificial diamonds         | 8. Vitamin research                    |
| 4. Heavy water                 | 9. Electrolytic extraction of aluminum |
| 5. Invention of bakelite       | 10. Extraction of sulfur               |

**MATCHING TEST**

*In the parenthesis next to each name in column A, write the number of the item in column B which is most closely associated with it.*

- | A               | B                        |
|-----------------|--------------------------|
| ( ) Graham      | 1. Planetary electrons   |
| ( ) Nobel       | 2. Atomic numbers        |
| ( ) Mendelejeff | 3. Isotopes              |
| ( ) Hopkins     | 4. Illinium              |
| ( ) Arrhenius   | 5. Weighing the electron |
| ( ) Moseley     | 6. Electrochemistry      |
| ( ) Bohr        | 7. Dynamite              |
| ( ) Millikan    | 8. Colloid chemistry     |
| ( ) Faraday     | 9. Discovery of radium   |
| ( ) Curie       | 10. Periodic table       |
|                 | 11. Coal-tar dyes        |
|                 | 12. Theory of ionization |

## APPENDIX

### IMPORTANT DEFINITIONS

- Absolute temperature.**—The centigrade temperature plus 273°. *Example:* 30°C. is equivalent to 303°A.
- Absolute zero.**—The temperature  $-273^{\circ}\text{C}$ ., at which, theoretically, there is an entire absence of heat.
- Acid.**—A hydrogen compound whose water solution contains hydrogen ( $\text{H}^+$ ) ions in sufficient number to turn blue litmus red. *Example:*  $\text{HCl}$ .
- Acid anhydride.**—A non-metallic oxide that unites with water to form an acid. *Example:*  $\text{SO}_2$ .
- Acid salt.**—A salt in which only part of the hydrogen of the acid has been replaced by a metal. *Example:*  $\text{NaHCO}_3$ .
- Adsorption.**—The condensation of a gas or other substance on the surface of a colloid. *Example:* removal of poison gases by charcoal.
- Alcohol.**—An organic compound consisting of a radical and one or more hydroxyl ( $\text{OH}$ ) groups. *Example:*  $\text{C}_2\text{H}_5\text{OH}$ .
- Aldehyde.**—An organic compound containing the  $\text{CHO}$  group. *Example:* formaldehyde ( $\text{H}\cdot\text{CHO}$ ).
- Alkali.**—A very soluble base whose water solution possesses strong basic properties. *Example:*  $\text{KOH}$ .
- Allotropic forms.**—Two or more forms of the same element differing mainly in physical properties, owing to a difference in their energy content. *Example:* Sulfur occurs in three forms, namely, rhombic, monoclinic, and amorphous.
- Alloy.**—A substance formed by mixing together two or more molten metals. *Example:* brass (composed of copper and zinc).
- Amalgam.**—An alloy of mercury with some other metal. *Example:* gold amalgam.
- Amorphous.**—Having no definite shape; non-crystalline. *Example:* starch.
- Amphoteric compound.**—A compound having acid or basic properties, depending upon the substance with which it reacts. *Example:*  $\text{Al}(\text{OH})_3$ .
- Amphoteric element.**—An element whose atom contains half or nearly half the maximum number of electrons in the outer ring, and which



can, therefore, act either as a metal or as a non-metal. *Example:* carbon.

**Analysis.**—The process of decomposing a substance into its elements or into simpler substances. *Example:* electrolysis of water, yielding hydrogen and oxygen.

**Anhydride.**—An oxide which combines with water to form an acid or a base. *Examples:*  $\text{CO}_2$ ,  $\text{MgO}$ .

**Anhydrous substance.**—A substance from which the water (either chemically combined or mechanically enclosed) has been removed. *Example:* anhydrous  $\text{Na}_2\text{CO}_3$ .

**Anion.**—An ion carrying a negative charge. It is attracted to the anode in electrolysis. *Example:* the chlorine ( $\text{Cl}^-$ ) ion.

**Anode.**—The electrode at which the current enters an electrolytic cell, and to which negative ions are attracted.

**Atom.**—The smallest particle of an element that can enter into a chemical reaction.

**Atomic number.**—The position number of an element in the list of all the elements. It is numerically equal to the number of planetary electrons in the atom. *Example:* The atomic number of sodium is 11.

**Atomic weight.**—The weight of an atom of an element compared to the weight of an atom of oxygen, which is established as 16.

**Base.**—A compound of a metallic element or radical with one or more hydroxyl groups; its water solution contains  $\text{OH}^-$  ions in sufficient number to turn litmus from red to blue. *Example:*  $\text{NaOH}$ .

**Basic anhydride.**—A metallic oxide that unites with water, forming a base. *Example:*  $\text{MgO}$ .

**Basic salt.**—A salt which contains one or more hydroxyl groups. *Example:*  $\text{Bi}(\text{OH})_2\text{NO}_3$ .

**Binary compound.**—A compound consisting of only two elements. *Examples:*  $\text{H}_2\text{O}$ ,  $\text{NaCl}$ .

**Calorie.**—The amount of heat required to raise the temperature of one gram of water one degree centigrade. A more convenient unit is the large Calorie (represented by a capital *C*), which is equivalent to 1000 small calories.

**Carbohydrate.**—A compound containing carbon, hydrogen, and oxygen, the last two elements being present in the ratio of 2:1, just as they are in water. *Example:* starch,  $\text{C}_6\text{H}_{10}\text{O}_5$ .

**Catalytic agent (catalyst).**—A substance which changes the speed of a chemical action without itself being permanently changed. *Example:*  $\text{MnO}_2$  in the preparation of oxygen from  $\text{KClO}_3$ .

**Cathode.**—The electrode at which the current leaves an electrolytic cell, and to which positive ions are attracted.

- Cation.**—An ion carrying a positive charge. It is attracted to the cathode in electrolysis. *Example:*  $H^+$ .
- Chemical change.**—A change which alters the composition of a substance, and forms one or more new substances. *Example:* rusting of iron.
- Chemistry.**—The science which deals with matter, its properties, and the changes which it undergoes.
- Colloid.**—A substance existing in such a finely divided condition that it remains permanently suspended in some supporting medium, and cannot be removed by ordinary filtration. *Example:* starch paste.
- Combining (equivalent) weight.**—The number of grams of an element that will combine with, or replace, one gram of hydrogen. *Example:* The combining weight of oxygen is 8.
- Combustion.**—Any chemical change that evolves heat and light. *Example:* burning of wood.
- Compound.**—A substance consisting of two or more elements chemically combined. *Example:*  $KClO_3$ .
- “Cracking”.**—The process of decomposing the molecules of complex hydrocarbons present in petroleum into simpler hydrocarbons, so as to obtain a greater yield of gasoline.
- Crystal.**—A solid of definite geometric shape. *Examples:* salt, sugar.
- Decomposition.**—The process of separating a compound into its elements or into simpler substances. *Example:* heating  $KClO_3$  to yield  $KCl$  and oxygen.
- Decrepitation.**—The crackling of certain crystals when heated, owing to the vaporization of water that is mechanically enclosed. *Example:* heated  $NaCl$  crystals.
- Dehydrating agent.**—A substance which, because it absorbs water readily, is used for drying other substances. *Example:* concentrated  $H_2SO_4$ .
- Deliquescent substance.**—A substance which absorbs sufficient water from the air to dissolve in it. *Examples:*  $NaOH$ ,  $CaCl_2$ .
- Density.**—The weight of a substance per unit volume. *Example:* The density of gold is 19.3 grams per cubic centimeter (gms: per cc.).
- Destructive distillation.**—The process of decomposing a complex organic substance by heating it in the absence of air and condensing the volatile products. *Example:* destructive distillation of soft coal.
- Deuterium.**—The isotope of hydrogen with a mass of 2, sometimes called *heavy hydrogen*.

- Deuteron.**—The nucleus of a deuterium atom, consisting of a proton and a neutron, and known to be one of the products of the decomposition of atomic nuclei.
- Developer.**—An organic reducing agent used in photography to make visible the latent image formed on the film by exposure to light rays. *Example:* pyrogallol.
- Dialysis.**—The process of separating a crystalloid from a colloid by the filtering action of a semi-permeable membrane such as parchment.
- Dibasic acid.**—An acid whose molecule has two replaceable hydrogen atoms. *Example:*  $\text{H}_2\text{SO}_4$ .
- Distillate.**—The liquid obtained by condensing the vapor during distillation.
- Distillation.**—The process of converting a liquid into a vapor by heat, and then condensing the vapor back to a liquid again, thus leaving behind non-volatile impurities.
- Double salt.**—A salt consisting of two metals combined with one acid radical, or one metal combined with two acid radicals. *Example:*  $\text{KAl}(\text{SO}_4)_2$ .
- Effervescence.**—The rapid escape of excess gas from a liquid. *Example:* the bubbling of soda water.
- Efflorescent substance.**—A substance which loses its water of crystallization on exposure to the air. *Example:*  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ .
- Electrochemical (replacement) series.**—A columnar arrangement of the metals which represents the order in which they replace each other from solution.
- Electrolysis.**—The process of decomposing a compound by means of the electric current. *Example:*  $2\text{H}_2\text{O} \longrightarrow 2\text{H}_2 + \text{O}_2$
- Electrolyte.**—A compound which, in water solution, conducts the electric current. *Examples:* soluble acids, bases, and salts.
- Electrolytic dissociation (ionization).**—The separation of a molecule of an electrolyte in solution into positive and negative ions. *Example:*  $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$ .
- Electron.**—A negatively charged particle which is one of the units of atomic structure.
- Element.**—A substance which scientists have not yet succeeded in decomposing into simpler substances by ordinary chemical means. *Examples:* hydrogen, carbon.
- Emulsion.**—A mixture of two non-miscible liquids in which separation takes place very slowly. *Example:* cream in milk.
- Energy.**—The capacity to do work. The various forms of energy are heat, light, mechanical, sound, electrical, and chemical.

**Enzyme.**—A complex substance that acts catalytically in hastening organic processes. *Example:* Pepsin promotes the digestion of protein in the stomach.

**Equilibrium.**—The point in a reversible reaction at which the speed of the reaction going to the right is exactly equal to that going to the left. *Example:*  $N_2 + 3H_2 \rightleftharpoons 2NH_3$ .

**Ester (ethereal salt).**—An organic compound formed by the reaction of an alcohol with an acid. *Example:* ethyl nitrate,  $C_2H_5NO_3$ .

**Explosive mixture.**—A mixture of gases in which, on ignition, combustion of the entire mass takes place at once. *Example:* hydrogen and oxygen.

**Fermentation.**—A chemical change brought about by the action of living organisms or enzymes.

**Filtrate.**—The clear liquid which runs through a filter in the process of filtration.

**Filtration.**—The process of running a liquid through a porous material, such as coarse paper, sand, etc., in order to remove suspended solid matter.

**Fixation of nitrogen.**—The process of converting atmospheric nitrogen into useful compounds such as ammonia and nitric acid.

**Fixer.**—A compound used in photography to dissolve unexposed silver bromide from the film. *Example:* sodium thiosulfate (hypo).

**Flame.**—A quantity of burning gas.

**Flux.**—A substance employed in metallurgy to promote fusion and to remove impurities. *Example:* limestone in the reduction of iron ore.

**Formula.**—A chemical expression made up of symbols, frequently with subscript numbers, and representing a molecule of a substance. *Example:*  $H_2SO_4$ .

**Fractional distillation.**—The process of distilling a mixture of two or more liquids and separating them according to their boiling points. *Example:* refining of petroleum.

**Fusion.**—The process of converting a solid into a liquid by heat.

**Gram-molecular volume.**—The volume (22.4 liters) occupied at standard conditions by the molecular weight (in grams) of a gas.

**Gram-molecular weight.**—The molecular weight of a substance expressed in grams. *Example:* The gram-molecular weight of nitrogen is 28 grams.

**Halogen ("salt-former").**—A member of the chlorine family of elements (chlorine, fluorine, bromine, and iodine).

**Hard water.**—Water which does not readily form a lather with soap, owing to the presence of dissolved salts of metals such as calcium and magnesium.

**Heat of formation.**—The number of calories of heat evolved or absorbed when one gram-molecular weight of a compound is formed from its elements. *Example:* The heat of formation of NaCl is 97,800 calories.

**Heat of neutralization.**—The heat evolved in the formation of one gram-molecular weight (18 grams) of water by the union of hydrogen ions and hydroxyl ions. The heat of neutralization is 13,700 calories.

**Heavy water.**—Water composed of two atoms of heavy hydrogen (*deuterium*) combined with an atom of oxygen; also known as *deuterium oxide*.

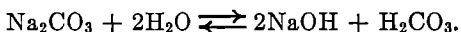
**Homologous series.**—A series of compounds (usually organic) in which the formulas of any two successive members vary by a common difference. *Example:* CH<sub>4</sub>, C<sub>2</sub>H<sub>6</sub>, C<sub>3</sub>H<sub>8</sub>, and C<sub>4</sub>H<sub>10</sub> are the first four compounds in the methane series of hydrocarbons. Thus the general formula for this group is given as C<sub>n</sub>H<sub>2n+2</sub>.

**Humidity (relative).**—The amount of moisture in the air, compared with the amount which the air could hold if it were saturated at the given temperature.

**Hydrate.**—A crystalline compound containing a definite quantity of water in combination. *Example:* Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O.

**Hydrocarbon.**—A compound consisting only of hydrogen and carbon. *Examples:* CH<sub>4</sub>, C<sub>2</sub>H<sub>2</sub>.

**Hydrolysis.**—The action of water on a salt to form an acid and a base, one of which is very slightly dissociated. *Example:*



**Hydroxyl.**—The OH radical.

**Indicator.**—A substance which changes color when in contact with an acid or a base. *Example:* litmus.

**Inert element.**—An element whose atom has a complete outer ring of electrons, and which is therefore chemically inactive. *Example:* helium.

**Inversion of sugar.**—The process of breaking up a complex sugar into simpler sugars. *Example:* changing sucrose to dextrose and levulose.

**Ion.**—An atom or a group of atoms carrying an electric charge. *Examples:* Na<sup>+</sup>, SO<sub>4</sub><sup>−−</sup>.

**Ionization.**—The dissociation of a substance into ions.

**Isomers.**—Two or more compounds with the same formula but with different properties, owing to a difference in the arrangement of the atoms in the molecule. *Example:* dextrose and levulose, both of which have the formula C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>.

**Isotopes.**—Two or more forms of an element differing slightly in atomic weight but having the same chemical properties. *Example:* Chlorine has three isotopes of atomic weights 35, 37, and 39.

**Kindling temperature.**—The lowest temperature at which a substance begins to burn.

**Law (chemical).**—A general truth established by exhaustive chemical experimentation. *Example:* Law of Definite Proportions.

**Matter.**—Anything that has weight and occupies space.

**Metal.**—An element which is a lender of electrons, forms a hydroxide, has a characteristic luster, and is usually a good conductor of electricity. *Examples:* sodium, copper, gold.

**Metallurgy.**—The process of extracting a metal from its ore.

**Metathesis (double decomposition).**—A chemical reaction between two compounds which results in the formation of two new compounds. *Example:*  $\text{NaCl} + \text{AgNO}_3 \longrightarrow \text{AgCl} + \text{NaNO}_3$ .

**Mineral.**—An inorganic substance found in the earth. *Example:* limestone,  $\text{CaCO}_3$ .

**Miscible liquids.**—Liquids which are mutually soluble. *Example:* glycerine and water.

**Mixture.**—A material made up of two or more substances that are not chemically combined. *Examples:* flour, air.

**Molar solution.**—A solution containing one gram molecular weight of the solute per liter. *Example:* A molar solution of  $\text{NaCl}$  contains 58.5 gms. of solute per liter.

**Molecular weight.**—The weight of a molecule of a compound compared to the weight of an atom of oxygen, taken as 16. It is the sum of the weights of all the atoms in the molecule. *Example:* The molecular weight of  $\text{H}_2\text{SO}_4 = 2 + 32 + 64 = 98$ .

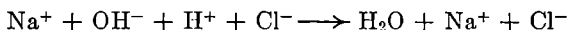
**Molecule.**—The smallest particle of a substance having the physical and chemical properties of that substance.

**Monobasic acid.**—An acid whose molecule has only one replaceable hydrogen atom. *Example:*  $\text{HCl}$ .

**Mordant.**—A substance used in dyeing to make the dye adhere to the cloth. *Example:* aluminum hydroxide.

**Nascent (atomic) state.**—The chemically active condition of an element at the moment that it is liberated from a compound.

**Neutralization.**—The reaction between an acid and a base, forming a salt and water; or the union of the hydrogen ( $\text{H}^+$ ) ions of an acid with the hydroxyl ( $\text{OH}^-$ ) ions of a base, forming undissociated water. *Example:*



**Neutron.**—An electrically neutral particle of unit mass, made up of a proton and an electron that are closely united to each other.

**Non-electrolyte.**—A compound whose water solution does not conduct the electric current. *Examples:* glycerine, sugar.

**Non-metal.**—An element which is a borrower of electrons, forms acids, does not have the characteristic metallic luster, and does not easily conduct electricity. *Examples:* chlorine, sulphur, nitrogen.

**Normal salt.**—A salt formed by the complete neutralization of an acid and a base, and which therefore does not contain replaceable hydrogen or hydroxyl groups. *Example:*  $\text{Na}_3\text{PO}_4$ .

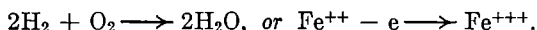
**Nucleus.**—The positively charged core or central portion of an atom, made up of electrons and protons, the latter always being in excess.

**Occlusion.**—The absorption of a gas by a solid. *Example:* absorption of hydrogen by platinum.

**Ore.**—A mineral from which an element, usually a metal, may be profitably extracted. *Example:* hematite ( $\text{Fe}_2\text{O}_3$ ), the chief ore of iron.

**Organic chemistry.**—The chemistry of the compounds of carbon.

**Oxidation.**—The union of a substance with oxygen; or the increase in the valence of a metal; or the loss of electrons. *Examples:*



**Oxide.**—A compound consisting of oxygen and another element. *Example:*  $\text{CaO}$ .

**Oxidizing agent.**—A substance which readily gives up oxygen to another substance; or a substance which can increase the valence of a metal. *Examples:*  $\text{H}_2\text{O}_2$ , or chlorine acting on  $\text{FeCl}_2$ .

**Physical change.**—A change which does not involve alteration in the composition of the substance. *Example:* tearing paper.

**Positron.**—A positively charged particle, of the same mass as an electron, which is present in the nuclear structure of the atom.

**Precipitate.**—An insoluble compound resulting from a chemical action between substances in solutions. *Example:* silver chloride formed in the reaction  $\text{KCl} + \text{AgNO}_3 \longrightarrow \text{AgCl}\downarrow + \text{KNO}_3$ .

**Property.**—A characteristic by which a substance may be identified. *Examples:* color, odor, taste.

**Proton.**—A positively charged particle which is one of the units of atomic structure.

**Radical.**—A group of elements which, in chemical reactions, behaves as a single element. *Examples:*  $\text{NH}_4$ ,  $\text{SO}_4$ ,  $\text{NO}_3$ .

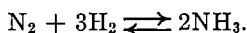
**Reaction.**—Any chemical change.

**Reducing agent.**—A substance which removes oxygen from another substance; or a substance which can decrease the valence of a metal.

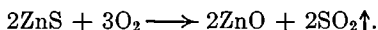
*Examples:*  $H_2$ , C, CO, or atomic hydrogen acting on  $FeCl_3$ .

**Reduction.**—The process of removing oxygen from a compound; or the decrease in the valence of a metal; or the gain of electrons.

**Reversible reaction.**—A chemical change that can proceed in either direction, depending upon the conditions. *Example:*



**Roasting.**—Heating an ore with an excess of air to convert the metallic compound present into an oxide. *Example:*



**Salt.**—A compound consisting of a metal or a metallic radical combined with a non-metal or an acid radical. *Examples:* NaCl,  $K_2SO_4$ ,  $NH_4Cl$ .

**Saponification.**—The reaction between a fat or an oil and an alkali. The process of soap-making.

**Saturated solution.**—A solution which contains all the solute it can normally dissolve at a given temperature and pressure.

**Slag.**—A product obtained in metallurgy, usually formed by the union of a flux with the impurities present in the ore. *Example:*  $CaSiO_3$  formed in the metallurgy of iron.

**Slow oxidation.**—The combining of a substance with oxygen, in which no light or noticeable heat is produced.

**Smelting.**—The processes involved in obtaining an element (usually a metal) from its ore by high temperature operations in a furnace.

**Soap.**—A metallic salt of a fatty acid. *Example:* sodium stearate.

**Solute.**—A dissolved substance. *Example:* salt in sea water.

**Solution.**—A uniform mixture of a solvent and a solute. *Example:* salt dissolved in water.

**Solvent.**—A liquid which dissolves a substance. *Example:* water.

**Specific gravity of a gas (air standard).**—The ratio between the weight of one liter of the gas and the weight of one liter of air under the same conditions of temperature and pressure. *Example:* The specific gravity of chlorine is 2.49.

**Specific gravity of a solid or a liquid.**—The ratio between the weight of the substance and the weight of an equal volume of water. *Example:* The specific gravity of mercury is 13.6.

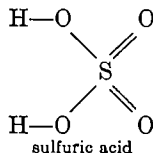
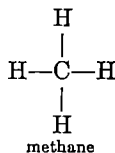
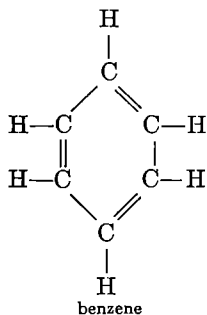
**Spontaneous combustion.**—Active burning caused by the gradual accumulation of heat that results from slow oxidation. *Example:* the combustion of oily rags in a closet.



**Standard conditions.**—A temperature of 0°C. and a pressure of 760 mm. of mercury.

**Strong acid (or base).**—An acid (or base) which is highly ionized in dilute water solution.

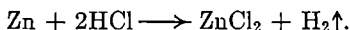
**Structural (graphic) formulas.**—A pictorial representation of the probable arrangement of the atoms in a molecule. *Examples:*



**Sublimation.**—The process of vaporizing a solid and then condensing the vapor back to a solid without having it pass through the liquid state. *Example:* sublimation of iodine.

**Substance.**—Any particular kind of matter, either element, compound, or mixture. *Examples:* nitrogen, salt, gasoline.

**Substitution (simple replacement).**—A chemical reaction in which one element takes the place of another in a compound. *Example:*



**Supersaturated solution.**—A solution which contains more solute than it can normally hold at the given temperature.

**Suspension.**—A mixture of a solid and a liquid in which the particles of the solid separate out on standing.

**Symbol.**—A letter or a pair of letters denoting one atom of an element. *Examples:* H, O, Ca, Fe.

**Synthesis.**—The combination of two or more substances, resulting in the formation of a more complex substance. *Example:*  $\text{C} + \text{O}_2 \longrightarrow \text{CO}_2$ .

**Tempering.**—The process of heating and cooling a metal in order to increase its hardness. *Example:* tempering of steel.

**Ternary compound.**—A compound consisting of three elements. *Example:*  $\text{K}_2\text{SO}_4$ .

**Tincture.**—A solution of some substance in alcohol. *Example:* tincture of iodine.

**Tribasic acid.**—An acid whose molecule has three replaceable hydrogen atoms. *Example:*  $\text{H}_3\text{PO}_4$ .

**Valence.**—The property of an element that determines the number of atoms of hydrogen with which its atom can combine or which it can replace. According to the Electron Theory, valence is the number which indicates how many electrons an element gains or loses in a chemical action. *Examples:* The valence of Na is 1; of Ca, 2.

**Vapor density.**—The ratio of the weight of a gas to the weight of an equal volume of hydrogen measured under the same conditions of temperature and pressure. *Example:* The V.D. of  $\text{CO}_2$  is 22.

**Vitamin.**—One of a group of complex substances, essential for health, which are present to a minute extent in foods. *Example:* Vitamin D, present in cod liver oil, prevents rickets.

**Volatile substance.**—A substance that vaporizes easily. *Example:* alcohol.

**Water of crystallization.**—The water that is chemically united with a substance when the latter crystallizes from solution. *Example:* the water in  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .

**Weak acid (or base).**—An acid (or base) which is only slightly ionized in dilute water solution.

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### GENERAL SOLUBILITY RULES

1. All nitrates, acetates, and chlorates are *soluble*.
2. All sodium, potassium, and ammonium salts (with very few exceptions) are *soluble*.
3. All chlorides, bromides, and iodides are *soluble*, except those of silver, lead, and mercury (-ous).
4. All sulfates are *soluble*, except those of barium, strontium, and lead. (Calcium sulfate is only slightly soluble.)
5. All oxides are *insoluble*, except those of sodium, potassium, calcium, barium, and strontium, all of which react with water to form soluble hydroxides.
6. All hydroxides are *insoluble*, except those of sodium, potassium, ammonium, calcium, barium, and strontium.
7. All carbonates, sulfides, phosphates, and silicates are *insoluble*, except those of sodium, potassium, and ammonium.

## CHEMICAL TESTS

## A—GASES

GAS	PROCEDURE	RESULTS
Ammonia.....	Note the odor. Test the gas with moistened red litmus paper.	Sharp, pungent odor. Litmus paper turns blue.
Carbon dioxide..	Pass the gas through limewater.	The limewater becomes milky.
Carbon monoxide....	Apply a flame to the gas. Pass the resulting product through limewater.	The gas burns with a pale-blue flame. The limewater becomes milky.
Chlorine.....	Note the color and odor.	Greenish-yellow color and strong, choking odor.
Hydrogen.....	Apply a flame (to the gas mixed with air).	Sharp explosion. The gas burns, forming water.
Hydrogen sulfide.....	Note the odor. Pass the gas over filter paper moistened with lead acetate solution.	Odor of rotten eggs. The filter paper turns black.
Nitric oxide.....	Expose the gas to the air.	Dark brown fumes of nitrogen peroxide are formed.
Nitrogen.....	Insert a lighted splint. Pass the gas through limewater.	The flame goes out. The gas has no odor and does not burn. The limewater remains clear.
Nitrogen peroxide.....	Note the color and odor. Test the solubility of the gas in water.	Brownish color and suffocating odor. The gas is fairly soluble.
Nitrous oxide...	Note the odor. Insert a glowing splint, also some barely ignited sulfur.	Sweetish odor. The splint bursts into flame, while the sulfur is extinguished.
Oxygen.....	Insert a glowing splint. Burn charcoal in the gas and test the product with limewater.	The splint bursts into flame. The limewater becomes milky.
Ozone.....	Note the odor. Pass the gas through a solution of potassium iodide and starch.	Pungent odor. The solution turns blue.
Sulfur dioxide...	Note the odor. Pass the gas through a solution of potassium permanganate.	Sharp choking odor. The solution is rendered colorless.

## B—ACID RADICALS

RADICAL	PROCEDURE	RESULTS
Acetate . . . . .	Add sulfuric acid to the substance and apply gentle heat.	Odor of vinegar.
Bromide . . . . .	To a solution of the salt, add chlorine water and chloroform. Shake vigorously.	The chloroform dissolves free bromine and becomes brownish-red.
Carbonate . . . . .	Add hydrochloric acid to the substance and pass the resulting gas through limewater.	The limewater turns milky.
Chloride . . . . .	To a solution of the substance, add a solution of silver nitrate.	A white curdy precipitate forms, which is insoluble in nitric acid and soluble in ammonium hydroxide.
Iodide . . . . .	Proceed as for the bromide test above.	The chloroform dissolves free iodine and becomes violet.
Nitrate . . . . .	To a water solution of the substance, add some freshly prepared ferrous sulfate solution. Then carefully add concentrated sulfuric acid, which forms a layer at the bottom.	A brown ring forms between the heavier sulfuric acid and the mixture above it.
Sulfate . . . . .	To a water solution of the substance, add some barium chloride solution.	A white precipitate, insoluble in hydrochloric acid, forms.
Sulfide . . . . .	To the substance (solid or solution), add hydrochloric acid and pass the resulting gas through a solution of lead acetate.	The gas formed has an odor of rotten eggs, and turns the solution black.
Sulfite . . . . .	Add hydrochloric acid to the substance and pass the resulting gas through a solution of potassium permanganate.	Effervescence. Sharp, choking odor. The solution becomes colorless.

## C—METALS

**1. Flame Tests.**—Carefully clean a platinum wire by dipping it into dilute nitric acid and heating it in the Bunsen flame. Repeat until the flame is colorless. Dip the heated wire into the substance being tested (either solid or solution), and then hold it in the hot outer part of the Bunsen flame.

COMPOUND OF	COLOR OF FLAME
Sodium	Yellow
Potassium	Violet
Lithium	Crimson
Calcium	Orange-red
Barium	Green
Strontium	Bright-red

NOTE.—In testing for a potassium salt, when a sodium salt is also known to be present, a cobalt glass must be used. This glass absorbs the yellow sodium light, but transmits the violet potassium light.

**2. Borax Bead Tests.**—Make a borax bead by heating some borax in a platinum wire loop. Dip the bead in the substance being tested, and heat it in the outer part of the Bunsen flame.

COMPOUND OF	COLOR OF BEAD
Cobalt	Blue
Manganese	Amethyst
Chromium	Bottle-green
Iron (ferric)	Yellow
Nickel	Brown

**3. Cobalt Nitrate Tests.**—Scoop out a small cavity in a plaster block, place in it some of the substance being tested, and heat strongly by means of a blow pipe. Moisten with a few drops of cobalt nitrate solution and reheat.

COMPOUND OF	COLOR OF SUBSTANCE
Aluminum	Blue
Zinc	Green
Magnesium	Pink

**4. Hydrogen Sulfide Tests.**—Pass hydrogen sulfide through a solution of the substance being tested.

COMPOUND OF	COLOR OF PRECIPITATE
Zinc	White
Antimony	Orange
Arsenic	Yellow (soluble in ammonium sulfide)
Cadmium	Yellow (soluble in concentrated nitric acid)
Copper } Lead } Silver }	Black (soluble in concentrated nitric acid)
	NOTE.—Further tests are required in these cases.
Mercury (-ic)	Black (soluble in aqua regia)
Mercury (-ous)	Black (insoluble in acids)

### 5. Other Tests

SUBSTANCE	PROCEDURE	RESULT
<b>Ferrous salt</b> . . . . .	To a solution of the salt, add a solution of potassium ferricyanide.	A dark-blue precipitate (Turnbull's blue) forms.
<b>Ferric salt</b> . . . . .	To a solution of the salt, add a solution of potassium ferrocyanide.	A dark-blue precipitate (Prussian blue) forms.
<b>Ammonium salt</b> .	To a solution of the salt, add a solution of sodium hydroxide and heat.	Odor of ammonia. The gas turns moist litmus paper blue.

### ACTIVITY (ELECTROCHEMICAL) SERIES

1. Lithium	7. Manganese	13. Nickel	19. Bismuth
2. Potassium	8. Zinc	14. Tin	20. Arsenic
3. Sodium	9. Chromium	15. Lead	21. Mercury
4. Calcium	10. Iron	16. Hydrogen	22. Silver
5. Magnesium	11. Cadmium	17. Copper	23. Platinum
6. Aluminum	12. Cobalt	18. Antimony	24. Gold

## ALLOYS

NOTE.—Because the compositions of certain alloys vary between rather wide limits, some of the figures in the following table may not agree with similar data from other sources.

CHIEF METAL	NAME	COMPOSITION (%)	SOME USES
Aluminum...	Duralumin	Al 95, Cu 4, Mg 1	Aircraft, kitchenware
	Magnalium	Al 90, Mg 10	Aircraft, balances
Bismuth....	Rose's metal	Bi 40, Pb 40, Sn 20	Boiler safety plugs
	Wood's metal	Bi 50, Pb 25, Sn 12.5, Cd 12.5	Automatic sprinklers, electric fuses
Copper.....	Aluminum bronze	Cu 90, Al 10	Scientific instruments
	Bell metal	Cu 75, Sn 25	Bells
	Brass	Cu 65, Zn 35	Electrical fixtures
	Bronze	Cu 82, Sn 16, Zn 2	Medals, statues
	Copper coin	Cu 95, Sn 4, Zn 1	U.S. coinage
	German silver	Cu 50, Ni 25, Zn 25	Electric rheostats
	Gun metal	Cu 90, Sn 10	Cannon
	Nickel coin	Cu 75, Ni 25	Coinage
Gold.....	Gold coin	Au 90, Cu 10	Coinage
	18-carat gold	Au 75, Cu 25	Jewelry
	White gold	Au 65, Ni 35	Jewelry
Iron.....	Chrome steel	Fe 97, Cr 3	Roller bearings, files
(NOTE: All iron alloys contain about 1% or less of carbon.)	Duriron	Fe 86, Si 14	Acid tanks
	Invar	Fe 64, Ni 36	Scientific instruments
	Manganese steel	Fe, 86, Mn 14	Safes, rock crushers
	Molybdenum steel	Fe 95, Mo 5	High-speed tools
	Nickel steel	Fe 94, Ni 6	Armor plate, bridges
	Platinite	Fe 58, Ni 42	"Lead-in" wires for electric bulbs
	Silicon steel	Fe 97, Si 3	Automobile springs
	Stainless steel	Fe 85, Cr 15	Kitchenware
	Tungsten steel	Fe 92, W 8	Cutting tools
Lead.....	Pewter	Pb 75, Sn 25	Ornaments, tableware
	Shot	Pb 99.5, As 0.5	Shot
	Solder	Pb 50, Sn 50	Joining metals
	Type metal	Pb 82, Sb 15, Sn 3	Printing type
Mercury....	Amalgams	Hg + Sn, Cu, Ag, or Au in varying proportions	Dental fillings, metallurgy
Nickel.....	Monel metal	Ni 60, Cu 34, Fe 6	Acid tanks
	Nichrome	Ni 60, Cr 15, Fe 25	Electric irons, toasters
	Permalloy	Ni 80, Fe 20	Electric transformers
Platinum....	Platinum	Pt 95, Ir 5	Jewelry
Silver.....	Silver coin	Ag 90, Cu 10	U.S. coinage
	Sterling	Ag 92.5, Cu 7.5	British coin, silverware
Tin.....	Babbitt metal	Sn 45, Pb 40, Sb 13, Cu 2	Machine bearings
	Britannia	Sn 90, Sb 8, Cu 2	Ornamental fixtures

## COMMON SUBSTANCES

COMMON NAME	CHEMICAL NAME	FORMULA OR SYMBOL
Agate	Silicon dioxide	$\text{SiO}_2$
Alcohol (grain)	Ethyl alcohol	$\text{C}_2\text{H}_5\text{OH}$
Alcohol (wood)	Methyl alcohol	$\text{CH}_3\text{OH}$
Alum	Potassium aluminum sulfate	$\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$
Alumina	Aluminum oxide	$\text{Al}_2\text{O}_3$
Ammonia water	Ammonium hydroxide	$\text{NH}_4\text{OH}$
Aqua fortis	Nitric acid	$\text{HNO}_3$
Aqua regia	Nitric and hydrochloric acids	$\text{HNO}_3 + 3\text{HCl}$
Arsenic, white	Arsenious oxide	$\text{As}_2\text{O}_3$
Baking powder	Sodium bicarbonate and a suitable salt that gives an acid reaction	$\text{NaHCO}_3 + \text{KHC}_4\text{H}_4\text{O}_6$
Baking soda	Sodium bicarbonate	$\text{NaHCO}_3$
Barite	Barium sulfate	$\text{BaSO}_4$
Benzol	Benzene	$\text{C}_6\text{H}_6$
Bleaching powder	Calcium chloride-hypochlorite	$\text{CaOCl}_2$
Blue vitriol	Copper sulfate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Boracic acid	Boric acid	$\text{H}_3\text{BO}_3$
Borax	Sodium tetraborate	$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$
Brimstone	Roll sulfur	S
Brine	Sodium chloride solution	$\text{NaCl} + \text{H}_2\text{O}$
Calomel	Mercurous chloride	$\text{HgCl}$
Cane sugar	Sucrose	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Carbolic acid	Phenol	$\text{C}_6\text{H}_5\text{OH}$
Carborundum	Silicon carbide	$\text{SiC}$
Caustic potash	Potassium hydroxide	$\text{KOH}$
Caustic soda	Sodium hydroxide	$\text{NaOH}$
Chalk	Calcium carbonate	$\text{CaCO}_3$
Chile saltpeter	Sodium nitrate	$\text{NaNO}_3$
Chloroform	Trichlormethane	$\text{CHCl}_3$
Choke damp	Carbon dioxide	$\text{CO}_2$
Chrome yellow	Lead chromate	$\text{PbCrO}_4$
Copperas	Ferrous sulfate	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
Corrosive sublimate	Mercuric chloride	$\text{HgCl}_2$
Corundum	Aluminum oxide	$\text{Al}_2\text{O}_3$
Cream of tartar	Potassium acid tartrate	$\text{KHC}_4\text{H}_4\text{O}_6$
Epsom salts	Magnesium sulfate	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Ether	Ethyl ether	$(\text{C}_2\text{H}_5)_2\text{O}$
Fire damp (marsh gas)	Methane	$\text{CH}_4$
Fluorspar	Calcium fluoride	$\text{CaF}_2$



## COMMON SUBSTANCES—Continued

COMMON NAME	CHEMICAL NAME	FORMULA OR SYMBOL
Fruit sugar	Fructose	$C_6H_{12}O_6$
Galena	Lead sulfide	PbS
Glauber's salts	Sodium sulfate	$Na_2SO_4 \cdot 10H_2O$
Glucose (grape sugar)	Dextrose	$C_6H_{12}O_6$
Glycerine	Glycerol	$C_3H_5(OH)_3$
Green vitriol	Ferrous sulfate	$FeSO_4 \cdot 7H_2O$
Gypsum	Calcium sulfate	$CaSO_4 \cdot 2H_2O$
Hematite	Ferric oxide	$Fe_2O_3$
Hypo	Sodium thiosulfate	$Na_2S_2O_3 \cdot 5H_2O$
Invert sugar	Dextrose and levulose	$C_6H_{12}O_6 + C_6H_{12}O_6$
Iodoform	Tri-iodo-methane	$CHI_3$
Iron pyrites	Iron disulfide	$FeS_2$
Laughing gas	Nitrous oxide	$N_2O$
Limestone	Calcium carbonate	$CaCO_3$
Limewater	Calcium hydroxide solution	$Ca(OH)_2$
Litharge	Lead monoxide	PbO
Lithopone	Zinc sulfide and barium sulfate	$ZnS + BaSO_4$
Lunar caustic	Silver nitrate	$AgNO_3$
Magnesia	Magnesium oxide	MgO
Magnetite	Iron tetroxide	$Fe_3O_4$
Marble	Calcium carbonate	$CaCO_3$
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$
Muriatic acid	Hydrochloric acid	HCl
Oil of vitriol	Sulfuric acid	$H_2SO_4$
Plaster of Paris	Calcium sulfate	$(CaSO_4)_2 \cdot H_2O$
Prussic acid	Hydrocyanic acid	HCN
Quartz	Silicon dioxide	$SiO_2$
Quicklime	Calcium oxide	CaO
Quicksilver	Mercury	Hg
Red lead	Lead tetroxide	$Pb_3O_4$
Rochelle salt	Potassium sodium tartrate	$KNaC_4H_4O_6$
Sal ammoniac	Ammonium chloride	$NH_4Cl$
Sal soda	Sodium carbonate	$Na_2CO_3 \cdot 10H_2O$
Salt (common)	Sodium chloride	NaCl
Salt peter	Potassium nitrate	$KNO_3$
Sand	Silicon dioxide	$SiO_2$
Silica	Silicon dioxide	$SiO_2$
Slaked lime	Calcium hydroxide	$Ca(OH)_2$
Soda ash	Anhydrous sodium carbonate	$Na_2CO_3$

COMMON SUBSTANCES—*Concluded*

COMMON NAME	CHEMICAL NAME	FORMULA OR SYMBOL
Sugar of lead	Lead acetate	$\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$
Vinegar	Dilute acetic acid	$\text{H} \cdot \text{C}_2\text{H}_3\text{O}_2$
Washing soda	Sodium carbonate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Water glass	Sodium silicate solution	$\text{Na}_2\text{SiO}_3 + \text{H}_2\text{O}$
White lead	Basic lead carbonate	$\text{Pb}(\text{OH})_2 \cdot 2\text{PbCO}_3$
White vitriol	Zinc sulfate	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
Whitewash	Calcium hydroxide suspension	$\text{Ca}(\text{OH})_2 + \text{H}_2\text{O}$
Whiting	Calcium carbonate	$\text{CaCO}_3$
Zinc white	Zinc oxide	$\text{ZnO}$

## PHYSICAL CONSTANTS OF COMMON GASES

NAME	FORMULA	MOLECULAR WEIGHT	VAPOR DENSITY ( <i>Hydrogen Standard</i> )	SPECIFIC GRAVITY ( <i>Air Standard</i> )	WEIGHT OF 1 LITER IN GRAMS ( <i>Standard Conditions</i> )
Acetylene.....	$\text{C}_2\text{H}_2$	26.0	13	0.906	1.17
Air.....			14.44	1.000	1.29
Ammonia.....	$\text{NH}_3$	17.0	8.5	0.596	0.77
Argon.....	A	39.9	19.9	1.378	1.78
Carbon dioxide.....	$\text{CO}_2$	44.0	22	1.529	1.98
Carbon monoxide.....	CO	28.0	14	0.967	1.25
Chlorine.....	$\text{Cl}_2$	71.0	35.5	2.491	3.20
Ethane.....	$\text{C}_2\text{H}_6$	30.0	15.0	1.049	1.35
Helium.....	He	4.0	2.0	0.137	0.178
Hydrogen.....	$\text{H}_2$	2.016	1.0	0.069	0.089
Hydrogen chloride.....	HCl	36.5	18.25	1.268	1.64
Hydrogen sulphide.....	$\text{H}_2\text{S}$	34.0	17.0	1.186	1.53
Methane.....	$\text{CH}_4$	16.0	8.0	0.554	0.717
Nitric oxide.....	NO	30.0	15.0	1.035	1.34
Nitrogen.....	$\text{N}_2$	28.0	14.0	0.967	1.25
Nitrous oxide.....	$\text{N}_2\text{O}$	44.0	22.0	1.530	1.98
Oxygen.....	$\text{O}_2$	32.0	16.0	1.105	1.429
Sulphur dioxide.....	$\text{SO}_2$	64.0	32.0	2.264	2.88

## THE NINETY-TWO ELEMENTS

(Arranged in order of atomic numbers)

ATOMIC NUMBER	ELEMENT	SYMBOL	ATOMIC WEIGHT	ATOMIC NUMBER	ELEMENT	SYMBOL	ATOMIC WEIGHT
1	Hydrogen	H	1.0078	47	Silver	Ag	107.88
2	Helium	He	4.002	48	Cadmium	Cd	112.41
3	Lithium	Li	6.940	49	Indium	In	114.8
4	Beryllium	Be	9.02	50	Tin	Sn	118.70
5	Boron	B	10.82	51	Antimony	Sb	121.76
6	Carbon	C	12.00	52	Tellurium	Te	127.61
7	Nitrogen	N	14.008	53	Iodine	I	126.92
8	Oxygen	O	16.0000	54	Xenon	Xe	131.3
9	Fluorine	F	19.00	55	Cesium	Cs	132.91
10	Neon	Ne	20.183	56	Barium	Ba	137.36
11	Sodium	Na	22.997	57	Lanthanum	La	138.90
12	Magnesium	Mg	24.32	58	Cerium	Ce	140.13
13	Aluminum	Al	26.97	59	Praseodymium	Pr	140.92
14	Silicon	Si	28.06	60	Neodymium	Nd	144.27
15	Phosphorus	P	31.02	61	Illinium	Il	146.0
16	Sulfur	S	32.06	62	Samarium	Sm	150.43
17	Chlorine	Cl	35.457	63	Europium	Eu	152.0
18	Argon	A	39.944	64	Gadolinium	Gd	157.3
19	Potassium	K	39.10	65	Terbium	Tb	159.2
20	Calcium	Ca	40.08	66	Dysprosium	Dy	162.46
21	Scandium	Sc	45.10	67	Holmium	Ho	163.5
22	Titanium	Ti	47.90	68	Erbium	Er	167.64
23	Vanadium	V	50.95	69	Thulium	Tm	169.4
24	Chromium	Cr	52.01	70	Ytterbium	Yb	173.5
25	Manganese	Mn	54.93	71	Lutecium	Lu	175.0
26	Iron	Fe	55.84	72	Hafnium	Hf	178.6
27	Cobalt	Co	58.94	73	Tantalum	Ta	181.4
28	Nickel	Ni	58.69	74	Tungsten	W	184.0
29	Copper	Cu	63.57	75	Rhenium	Re	186.31
30	Zinc	Zn	65.38	76	Osmium	Os	190.2
31	Gallium	Ga	69.72	77	Iridium	Ir	193.1
32	Germanium	Ge	72.60	78	Platinum	Pt	195.23
33	Arsenic	As	74.91	79	Gold	Au	197.2
34	Selenium	Se	78.96	80	Mercury	Hg	200.61
35	Bromine	Br	79.916	81	Thallium	Tl	204.39
36	Krypton	Kr	83.7	82	Lead	Pb	207.22
37	Rubidium	Rb	85.44	83	Bismuth	Bi	209.00
38	Strontium	Sr	87.63	84	Polonium	Po	210.0
39	Yttrium	Y	88.92	85	Alabamine	Ab	221.0
40	Zirconium	Zr	91.22	86	Radon	Rn	222.0
41	Columbium	Cb	92.91	87	Virginium	Vi	224.0
42	Molybdenum	Mo	96.0	88	Radium	Ra	225.97
43	Masurium	Ma	97.8	89	Actinium	Ac	227.0
44	Ruthenium	Ru	101.7	90	Thorium	Th	232.12
45	Rhodium	Rh	102.91	91	Protoactinium	Pa	231.04
46	Palladium	Pd	106.7	92	Uranium	U	238.14

## SOME HELPFUL BOOKS FOR FURTHER STUDY

- |                                       |   |
|---------------------------------------|---|
| ALEXANDER, J.                         | <i>Colloid Chemistry</i>                          |
| AMES, M. U. and JAFFE, B.             | <i>Laboratory and Workbook Units in Chemistry</i> |
| BENSON, H. K.                         | <i>Industrial Chemistry</i>                       |
| CALDWELL, O. W. and<br>SLOSSON, E. E. | <i>Science Remaking the World</i>                 |
| CHAMBERLAIN, J. S.                    | <i>Chemistry in Agriculture</i>                   |
| CLARK, B. L.                          | <i>Marvels of Modern Chemistry</i>                |
| CURIE, M.                             | <i>Pierre Curie</i>                               |
| DARROW, F. L.                         | <i>The Story of Chemistry</i>                     |
| DE KRUIF, P.                          | <i>Microbe Hunters</i>                            |
| DEMING, H. G.                         | <i>In the Realm of Carbon</i>                     |
| DUNCAN, R. K.                         | <i>The Chemistry of Commerce</i>                  |
| EDDINGTON, A. S.                      | <i>The Nature of the Physical World</i>           |
| FABRE, J. H. C.                       | <i>Wonder Book of Chemistry</i>                   |
| FINDLAY, A.                           | <i>Chemistry in the Service of Man</i>            |
| FOSTER, W.                            | <i>The Romance of Chemistry</i>                   |
| FREUND, I.                            | <i>The Study of Chemical Composition</i>          |
| GARBEDIAN, H. G.                      | <i>The March of Science</i>                       |
| HARROW, B.                            | <i>Eminent Chemists of Our Time</i>               |
| HAWK, P. B.                           | <i>What We Eat</i>                                |
| HAYNES, W.                            | <i>Chemical Economics</i>                         |
| HENDRICK, E.                          | <i>Everyman's Chemistry</i>                       |
| HOLMYARD, E. J.                       | <i>Makers of Chemistry</i>                        |
| HOWE, H. E.                           | <i>Chemistry in Industry</i>                      |
| JAFFE, B.                             | <i>Crucibles</i>                                  |
| —                                     | <i>Outposts of Science</i>                        |
| JENKS, T.                             | <i>Chemistry for Young People</i>                 |
| KENDALL, J.                           | <i>At Home among the Atoms</i>                    |
| LANGDON-DAVIES, J.                    | <i>Inside the Atom</i>                            |
| LEONARD, J. N.                        | <i>Crusaders of Chemistry</i>                     |
| MARTIN, G.                            | <i>Triumphs and Wonders of Modern Chemistry</i>   |
| MEADE, E. S.                          | <i>The Story of Gold</i>                          |
| PHILIP, J. C.                         | <i>The Romance of Modern Chemistry</i>            |
| RUSSELL, B.                           | <i>A. B. C. of Atoms</i>                          |
| SADTLER, S. P.                        | <i>Chemistry of Familiar Things</i>               |
| SLOSSON, E. E.                        | <i>Creative Chemistry</i>                         |
| SMITH, E. F.                          | <i>Chemistry in America</i>                       |
| STIEGLITZ, J. O.                      | <i>Chemistry in Medicine</i>                      |
| SURFACE, G. T.                        | <i>The Story of Iron</i>                          |
| THOMAS, L.                            | <i>Men of Danger</i>                              |
| TOWER, W. S.                          | <i>The Story of Oil</i>                           |
| VENABLE, F. P.                        | <i>Short History of Chemistry</i>                 |

# CHEMISTRY

Tuesday, June 20, 1944 — 1.15 to 4.15 p. m., only

*Answer all questions in part I and five questions from part II. Answers to the questions in part I should be written on the question paper as directed and handed in with the other answer paper. Answers should be numbered and lettered to correspond with the questions.*

## Part I

*Answer all questions in part I.*

Write on the line at the right of each statement the term which, if inserted in the blank, will make the statement true. [20]

1. A molecule of nitrogen is represented by the formula \_\_\_\_\_ 1\_\_\_\_\_
2. One of the products of the combustion of a solid was water. The solid contained the element \_\_\_\_\_ 2\_\_\_\_\_
3. An acid reacts with a carbonate to yield the gas \_\_\_\_\_ 3\_\_\_\_\_
4. Compounds that contain the radical  $\text{SO}_3$  are named \_\_\_\_\_ 4\_\_\_\_\_
5. Zinc carbonate, when roasted, yields the compound zinc \_\_\_\_\_ 5\_\_\_\_\_
6. The nitrogen compound obtained commercially by the destructive distillation of coal is \_\_\_\_\_ 6\_\_\_\_\_
7. In the reaction between hydrogen and ferric oxide, the \_\_\_\_\_ acts as the oxidizing agent. 7\_\_\_\_\_
8. Under standard conditions, the weight of 22.4 liters of carbon monoxide is \_\_\_\_\_ grams. 8\_\_\_\_\_
9. Water gas consists principally of carbon monoxide and \_\_\_\_\_ 9\_\_\_\_\_
10. When oxygen reacts with copper, the oxygen \_\_\_\_\_ electrons. 10\_\_\_\_\_
11. The reaction between steam and hot iron yields the gas \_\_\_\_\_ 11\_\_\_\_\_
12. An alloy of tin and lead is called \_\_\_\_\_ 12\_\_\_\_\_
13. One molecule of  $\text{Zn}_3(\text{PO}_4)_2$  consists of \_\_\_\_\_ atoms. 13\_\_\_\_\_
14. When lime mortar hardens, calcium hydroxide changes to calcium \_\_\_\_\_ 14\_\_\_\_\_
15. The formula for scandium (Sc) oxide is \_\_\_\_\_ 15\_\_\_\_\_
16.  $\text{Hg}^+ + 1$  electron yields \_\_\_\_\_ 16\_\_\_\_\_
17. The reaction between methyl alcohol and an acid yields water and a (an) \_\_\_\_\_ 17\_\_\_\_\_
18. The molecular weight of HCl is 36.5. A molar solution of this compound contains \_\_\_\_\_ grams of HCl per liter of solution. 18\_\_\_\_\_
19. Soda water does not taste so sour as lemon juice because the concentration of \_\_\_\_\_ ions in the soda water is lower. 19\_\_\_\_\_

## RECENT EXAMINATION PAPERS

20. In the electrolysis of a solution of cupric chloride, the cupric ions move toward the ..... electrode. 20.....

In *some* of the following statements the term in italics makes the statement incorrect. For each *incorrect* statement, write on the line at the right the term that must be substituted for the italicized term to make the statement correct. For each *correct* statement, write the word *true* on the line at the right. [11]

21. The compound represented by the formula  $C_6H_{10}O_5$  is a *carbohydrate*. 21.....

22. Hydrogen chloride is *less* soluble in water than is carbon dioxide. 22.....

23. Equal volumes of gases, at the same temperature and pressure, contain *the same* numbers of molecules. 23.....

24. An object that is to be plated is made the *cathode* of the plating circuit. 24.....

25. An atom which loses one or more electrons forms a *negative* ion. 25.....

26. When a *metallic* oxide reacts with water, a base is formed. 26.....

27. Most oxides are *soluble* in water. 27.....

28. Gases are *less* soluble in water at  $10^\circ$  C than in water at  $30^\circ$  C. 28.....

29. When calcium reacts with water, hydrogen and calcium *oxide* are formed. 29.....

30. *Dry* colored cloth is bleached when exposed to chlorine. 30.....

31. The pH of a 0.1 N solution of sodium carbonate is 11.6; the solution is *acidic*. 31.....

Write on the line at the right of *each* statement the *number* preceding the term that best completes the statement. [19]

32. The reaction between a strong acid and a weak base produces a salt that is (1)acidic (2)basic (3)neutral (4) variable 32

33. Kerosene is a mixture of (1)alcohols (2)carbohydrates (3)esters (4)hydrocarbons 33

34. Cotton fibers are nearly pure (1)cellulose (2)dextrose (3)nylon (4)starch 34.....

35. The formula of a chlorine substitution product of methane is (1) $CH_3Cl$  (2) $C_2H_4Cl_2$  (3) $C_3H_4Cl_4$  (4) $C_6H_5Cl$  35.....

36. Chlorine may be obtained from hydrochloric acid by (1) catalysis (2)dehydration (3)oxidation (4)reduction 36.....

37. When blue and red litmus are placed in a water solution of zinc chloride (1)neither changes color (2)both are bleached (3)the red litmus turns blue (4)the blue litmus turns red 37.....

38. The atomic number of an inert element is (1)37 (2) 53 (3)86 (4)88 38.....

## RECENT EXAMINATION PAPERS

39. A sulfur compound used as a bleaching agent is (1)  $\text{H}_2\text{SO}_4$   
(2)  $\text{H}_2\text{S}$  (3)  $\text{SO}_2$  (4)  $\text{SO}_3$  39.....
40. A water solution of ferrous sulfate will not react with a container made of (1) aluminum (2) magnesium (3) tin (4) zinc 40.....
41. Permanent hard water may contain the compound (1) calcium bicarbonate (2) magnesium chloride (3) potassium sulfate (4) sodium chloride 41.....
42. Ammonia is not collected by displacement of water because it (1) forms an acid in solution (2) dissolves in water (3) is easily decomposed (4) is less dense than air 42.....
43. Liquid air is an important source of (1) carbon dioxide (2) helium (3) hydrogen (4) oxygen 43.....
44. The gas produced when copper reacts with dilute nitric acid is (1) hydrogen (2) nitric oxide (3) nitrogen (4) nitrous oxide 44.....
45. Sodium carbonate is obtained commercially from sodium (1) chloride (2) hydroxide (3) nitrate (4) sulfate 45.....
46. The law of multiple proportions is illustrated by the compounds (1)  $\text{NaI}$  and  $\text{CaI}_2$  (2)  $\text{Ag}_3\text{PO}_4$  and  $\text{Ag}_2\text{SO}_4$  (3)  $\text{FeCl}_2$  and  $\text{FeCl}_3$  (4)  $\text{H}_2\text{O}$  and  $\text{H}_2\text{SO}_4$  46.....
47. The ratio of hydrogen to oxygen in water, by weight, is (1) 1:2 (2) 1:8 (3) 2:1 (4) 8:1 47.....
48. Compounds that are used as fertilizers should be (1) insoluble (2) soluble (3) unstable (4) volatile 48.....
49. Galvanized iron is iron coated with (1) copper (2) lead (3) tin (4) zinc 49.....
50. A base is obtained by the reaction between cold distilled water and (1) copper (2) iron (3) potassium (4) silver 50.....

### Part II

*Answer five questions from part II.*

1. Write balanced equations for *five* of the following: [10]
  - a. The electrolysis of a potassium chloride solution
  - b. The reaction between ferric sulfate and barium hydroxide
  - c. The reaction between silver nitrate and aluminum chloride
  - d. The laboratory preparation of nitric acid
  - e. The reaction between ammonium carbonate and sulfuric acid
  - f. The reaction between acetic acid and sodium hydroxide
2. Describe a test by which you could distinguish between the two compounds in *each* of the following: [10]
  - a. Lithium chloride and potassium chloride
  - b. Aluminum nitrate and cobalt nitrate
  - c. Ferric chloride and ferric sulfate
  - d. Barium nitrate and barium carbonate
  - e. Zinc sulfide and zinc oxide

## RECENT EXAMINATION PAPERS

3. a. Make a diagram of the apparatus used for the laboratory preparation and collection of hydrogen sulfide. [4]  
b. Write a balanced equation for the complete combustion of hydrogen sulfide. [2]  
c. Find the weight of zinc sulfide produced when 8 liters of hydrogen sulfide gas react with a solution of zinc chloride. [4]
4. a. By means of a labeled diagram, show how aluminum is obtained. Name the chemicals used. [3]  
b. Write a balanced equation for the reaction. [2]  
c. Name an aluminum alloy and state its composition. [2]  
d. Find the weight of  $\text{MgCO}_3$  needed to produce 50 tons of  $\text{MgO}$ . [3]  
$$[\text{MgCO}_3 \longrightarrow \text{MgO} + \text{CO}_2]$$
5. a. Why does copper react with hot concentrated sulfuric acid and not with dilute sulfuric acid? [2]  
b. Select *four* reactions that go to completion and give the reason why each goes to completion: [No credit without correct reason] [8]  
(1) Concentrated sulfuric acid heated with solid sodium chloride  
(2) Dilute sulfuric acid and lead nitrate solution  
(3) Dilute sulfuric acid and aluminum chloride solution  
(4) Dilute sulfuric acid and iron  
(5) Dilute sulfuric acid and potassium hydroxide solution
6. a. What is meant by the term *allotropic forms* of an element? [2]  
b. Name *two* elements that have allotropic forms. [2]  
c. Describe briefly the commercial production of *one* of the following fuel gases: (1) producer gas, (2) coal gas. [2]  
d. Find the volume of oxygen needed for the complete combustion of 60 liters of propane,  $\text{C}_3\text{H}_8$ . [4]
7. Explain *each* of the following: [10]  
a. Water is formed when dry crystals of washing soda are heated.  
b. Oxygen forms in a solution of chlorine water.  
c. A solution of acetic acid is a poor conductor of electricity.  
d. The difference between evaporation and distillation.  
e. Caesium (Cs) is more active than sodium.
8. a. Make a diagram to show the probable structure of the chloride *ion*. [2]  
b. Which *two* of these elements—elements of atomic number 9, 40, 54, 79—react with chlorine? [2]  
c. Lithium (Li) has isotopes of atomic weights 6 and 7.  
(1) Make a diagram of the atomic structure of *each* isotope. [4]  
(2) Explain why these isotopes have identical chemical properties. [2]



# CHEMISTRY

Tuesday, January 23, 1945 — 1.15 to 4.15 p. m., only

*Answer all questions in part I and five questions from part II. Answers to the questions in part I should be written on the question paper as directed and handed in with the other answer paper. Answers should be numbered and lettered to correspond with the questions.*

## Part I

*Answer all questions in part I.*

Write on the line at the right of *each* statement the term which, if inserted in the corresponding blank, will make the statement true. [21]

1. When zinc burns in ozone, zinc \_\_\_\_\_ is formed. 1\_\_\_\_\_
2. The products formed by the reaction between sodium hydroxide and phosphoric acid are water and \_\_\_\_\_. 2\_\_\_\_\_
3. The gas produced by the reaction between calcium chloride and concentrated sulfuric acid is \_\_\_\_\_. 3\_\_\_\_\_
4. A nitrogen compound obtained by the destructive distillation of soft coal is \_\_\_\_\_. 4\_\_\_\_\_
5. Producer gas consists almost entirely of nitrogen and \_\_\_\_\_. 5\_\_\_\_\_
6. Boiling a solution of calcium bicarbonate yields a precipitate of \_\_\_\_\_. 6\_\_\_\_\_
7. Roasting zinc sulfide yields the compound zinc \_\_\_\_\_. 7\_\_\_\_\_
8. Sterling silver contains silver and \_\_\_\_\_. 8\_\_\_\_\_
9. When plaster of Paris hardens, it combines with \_\_\_\_\_. 9\_\_\_\_\_
10. The molecular weight of  $\text{Ca}(\text{NO}_3)_2$  is \_\_\_\_\_. 10\_\_\_\_\_
11. The electrolysis of a sodium chloride solution yields chlorine, sodium hydroxide and \_\_\_\_\_. 11\_\_\_\_\_
12. Some metallic oxides react with water to form \_\_\_\_\_. 12\_\_\_\_\_
13. The chlorides of mercury (I), lead and \_\_\_\_\_ are insoluble in water. 13\_\_\_\_\_
14. Acetylene is produced by the reaction between water and \_\_\_\_\_. 14\_\_\_\_\_
15. The cobalt nitrate test is used to identify compounds of magnesium, zinc and \_\_\_\_\_. 15\_\_\_\_\_
16. Hydrogen fluoride may be prepared by the reaction between sulfuric acid and \_\_\_\_\_. 16\_\_\_\_\_
17. A molar solution of lithium chloride is prepared by dissolving \_\_\_\_\_ grams of the salt in water and then adding enough water to form a liter of solution. 17\_\_\_\_\_
18. The activity of the halogen elements \_\_\_\_\_ as their atomic numbers increase. 18\_\_\_\_\_

## RECENT EXAMINATION PAPERS

19. When a metal is refined by electrolysis, the impure metal is made the \_\_\_\_\_ electrode of the cell. 19\_\_\_\_\_
20. A solution of pH 4 is \_\_\_\_\_ acidic than is one of pH 6. 20\_\_\_\_\_
21. The valence of phosphorus, in the compound  $H_3PO_4$ , is \_\_\_\_\_ 21\_\_\_\_\_

Write on the line at the right of *each* statement the *number* preceding the term that best completes the statement. [14]

22.  $MnO_2$  is used in the laboratory preparation of chlorine as (1) a catalytic agent (2) an oxidizing agent (3) a dehydrating agent (4) a reducing agent 22\_\_\_\_\_
23. A mixture of 40 grams of hydrogen and 40 grams of oxygen is ignited. At the end of the reaction, there remains (1) 20 grams of  $H_2$  (2) 20 grams of  $O_2$  (3) 35 grams of  $H_2$  (4) no  $H_2$  or  $O_2$  23\_\_\_\_\_
24. Compared with the number of molecules in 10 liters of  $N_2O$ , the number of molecules in 10 liters of  $NO$ , under the same conditions, is (1) one half as great (2) two thirds as great (3) the same (4) twice as great 24\_\_\_\_\_
25. Tin is replaced from solutions of its salts by (1) copper (2) nickel (3) mercury (4) silver 25\_\_\_\_\_
26. In the commercial production of sodium carbonate use is made of sodium (1) chloride (2) hydroxide (3) nitrate (4) sulfate 26\_\_\_\_\_
27. A strong acid is one that (1) is concentrated (2) is highly ionized (3) is very corrosive (4) tastes very sour 27\_\_\_\_\_
28. Moseley's work resulted in a periodic arrangement of the elements based on their (1) atomic numbers (2) atomic weights (3) number of isotopes (4) relative activity 28\_\_\_\_\_
29. Gasoline is a (1) mixture of ethane and benzene (2) single hydrocarbon (3) mixture of carbohydrates (4) mixture of hydrocarbons 29\_\_\_\_\_
30. The compound represented by the formula  $MgSO_4 \cdot 7H_2O$  is (1) an anhydride (2) a hydrate (3) a hydride (4) a hydroid 30\_\_\_\_\_
31. Water is said to be permanently hard if it contains dissolved (1)  $CaCl_2$  (2)  $Ca(HCO_3)_2$  (3)  $NaCl$  (4)  $Na_2CO_3$  31\_\_\_\_\_
32. An element used in light-weight alloys is (1) beryllium (2) bismuth (3) chromium (4) nickel 32\_\_\_\_\_
33. The water solution of a salt formed by the reaction between a weak base and a strong acid is (1) acidic (2) basic (3) neutral (4) variable 33\_\_\_\_\_
34. Oxygen is usually prepared in the laboratory from (1) air (2) carbonates (3) chlorates (4) nitrates 34\_\_\_\_\_
35. A compound that is deliquescent is (1) barium chloride (2) calcium chloride (3) copper sulfate (4) magnesium sulfate 35\_\_\_\_\_

## RECENT EXAMINATION PAPERS

In *some* of the following statements the term in italics makes the statement incorrect. For each *incorrect* statement, write on the line at the right the term that must be substituted for the italicized term to make the statement correct. For each *correct* statement, write the word *true* on the line at the right. [15]

- |  |         |
|--|---------|
| 36. In the electric furnace, sand and carbon react to form the abrasive with the formula <i>SiC</i> .                | 36..... |
| 37. The products of the combustion of aviation gasoline weigh <i>the same</i> as the original gasoline.              | 37..... |
| 38. A molecule of $C_2H_5 OH$ dissolved in water yields <i>two</i> ions.   | 38..... |
| 39. Salts of hydrobromic acid are called <i>bromates</i> .   | 39..... |
| 40. Metallic sodium is stored in <i>water</i> .  | 40..... |
| 41. The heats of formation of the more stable compounds are relatively <i>low</i> .                                  | 41..... |
| 42. A water solution of hydrogen sulfide is a <i>weak</i> acid.  | 42..... |
| 43. The <i>halogen</i> elements are oxidizing agents.  | 43..... |
| 44. Under standard conditions, the weight of 44.8 liters of hydrogen chloride is <i>thirty-six and a half</i> grams. | 44..... |
| 45. The substance called "dry ice" is solid <i>ammonia</i> .   | 45..... |
| 46. $MnO_2$ , used in the laboratory preparation of oxygen, serves as an <i>oxidizing agent</i> .                    | 46..... |
| 47. Most pottery is composed of mixtures of <i>silicates</i> .   | 47..... |
| 48. The reaction between potassium and water yields oxygen and potassium <i>peroxide</i> .                           | 48..... |
| 49. The solubility of oxygen in water <i>increases</i> as the temperature of the water increases.                    | 49..... |
| 50. Hydrolysis is the reverse of <i>ionization</i> .   | 50..... |

### Part II

*Answer five questions from part II.*

- Make a diagram of the apparatus used for the laboratory preparation and collection of nitric acid. Name the chemicals used. [3]
  - Describe the chemical test for the nitrate radical. [2]
  - Why does the reaction between concentrated nitric acid and iron yield no hydrogen? [2]
  - Find the volume of nitric oxide obtained from the reaction between copper and 100 grams of nitric acid. [3]  
 $3Cu + 8HNO_3$  (dilute)  $\rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$

## RECENT EXAMINATION PAPERS

2. *a.* Each of these processes—complete combustion, esterification, neutralization—is represented by one of the following equations. Write the name of *each* process on your answer paper and after it write the *number* of the proper equation. [3]
- (1)  $\text{Ca}(\text{OH})_2 + 2\text{HCl} \longrightarrow \text{CaCl}_2 + 2\text{H}_2\text{O}$   
 (2)  $2\text{C}_2\text{H}_6 + 3\text{O}_2 \longrightarrow 4\text{C} + 6\text{H}_2\text{O}$   
 (3)  $2\text{C}_4\text{H}_{10} + 13\text{O}_2 \longrightarrow 8\text{CO}_2 + 10\text{H}_2\text{O}$   
 (4)  $\text{C}_2\text{H}_5\text{OH} + \text{HCl} \longrightarrow \text{C}_2\text{H}_5\text{Cl} + \text{H}_2\text{O}$   
 (5)  $\text{C}_3\text{H}_5(\text{C}_{18}\text{H}_{35}\text{O}_2)_3 \text{ (glyceryl stearate)} + 3\text{NaOH} \longrightarrow \text{C}_3\text{H}_5(\text{OH})_3 + 3\text{Na}(\text{C}_{18}\text{H}_{35}\text{O}_2)$
- b.* Find the volume of oxygen that will react with 27 cubic feet of  $\text{C}_2\text{H}_6$ . [This reaction is represented by equation (2), part *a.*] [3]
- c.* How is ethyl alcohol changed to an aldehyde? How is the aldehyde then changed to acetic acid? [2]
- d.* Write the equation that represents the preparation of methyl alcohol from carbon monoxide. [2]
3. *a.* Describe a process by which sulfuric acid is manufactured; include all essential steps. [Equations may be used.] [5]
- b.* Cellulose reacts with a mixture of nitric and sulfuric acids to form nitrocellulose. State the purpose of the sulfuric acid. [2]
- c.* Find the weight of sulfur needed to produce 800 pounds of sulfuric acid. [3]
4. *a.* Atoms consist of electrons, neutrons and protons. Compare the proton, with respect to weight and electrical charge, with (1) the electron, (2) the neutron. [2]
- b.* Write the symbols of *two ions* that have the same number and arrangement of electrons as the argon atom. [2]
- c.* Write an electronic equation that represents the change from a potassium atom to a potassium ion. Does this change represent oxidation? [2]
- d.* Potassium has isotopes of atomic weights 39 and 40.  
 (1) How do the structures of the atoms of these isotopes differ? [2]  
 (2) Calcium has an isotope with an atomic weight of 40. Explain why the chemical properties of this calcium isotope are different from those of the potassium isotope of the same weight. [2]
5. *a.* Describe laboratory procedures that may be used to  
 (1) Separate lead carbonate from a dry mixture of lead carbonate and magnesium nitrate [2]  
 (2) Show that iron is less active than zinc [2]  
 (3) Remove sulfur dioxide from a mixture of oxygen and sulfur dioxide [2]

## RECENT EXAMINATION PAPERS

- b. Explain why
- (1) Black smoke is produced when a hydrocarbon burns in chlorine [2]
  - (2) Water gas is poisonous [2]
6. Write a balanced equation for *each* of the following: [10]
- The reaction between sodium carbonate and nitric acid
  - The reaction between aluminum hydroxide and sulfuric acid
  - The reaction between scandium (Sc) and chlorine
  - The reaction between ammonium chloride and barium hydroxide, when the mixture is heated
  - The reaction between a silver nitrate solution and hydrogen sulfide.
7. a. Pig iron is manufactured in the blast furnace.
- (1) Name the raw materials used and state the purpose of *each* material. [4]
  - (2) Write the chemical name or the formula of a *solid* by-product formed. [1]
- b. Write an equation that represents the production of chromium by the reaction between  $\text{Cr}_2\text{O}_3$  and Al. [2]
- c. Bleaching powder is made from calcium carbonate in three steps. Calcium oxide and calcium hydroxide are the intermediate products. Describe the three processes involved. [3]
8. a. From the reactions indicated below, select *two* that go to completion and give the reason why each goes to completion. [No credit without correct reason] [4]
- (1) Silver nitrate solution and ammonium bromide solution
  - (2) Dilute hydrochloric acid and ferrous sulfate solution
  - (3) Dilute sulfuric acid and ammonium hydroxide solution
- b. Write an ionic equation that represents the reaction between dilute sulfuric acid and ammonium hydroxide solution. [2]
- c. In *each* of the reactions chosen in answer to (a), change *one* of the reacting materials so that the resulting reaction will *not* go to completion. [4]

# CHEMISTRY

Tuesday, June 19, 1945 — 1.15 to 4.15 p. m., only

*Answer all questions in part I and five questions from part II. Answers to the questions in part I should be written on the question paper as directed and handed in with the other answer paper. Answers should be numbered and lettered to correspond with the questions.*

## Part I

*Answer all questions in part I.*

Write on the line at the right of *each* statement the term which, if inserted in the blank, will make the statement true. [25]

1. In the electrolysis of water, ..... is liberated at the anode. 1.....
2. The compound formed when zinc dust combines with sulfur is called ..... 2.....
3. Equal volumes of gases under like conditions have equal numbers of ..... 3.....
4. The electrolytic process for the extraction of aluminum was discovered by ..... 4.....
5. Sulfur dioxide and oxygen, heated with a catalyst, react to form the compound ..... 5.....
6. Magnesium reacts with acetic acid to form hydrogen and ..... 6.....
7. The molecular weight of  $\text{Ca}_3(\text{PO}_4)_2$  is ..... 7.....
8. Metallic sodium reacts with water to form hydrogen and the compound ..... 8.....
9. The physical property of sulfuric acid that makes it valuable in preparing other acids is its ..... 9.....
10. The formula for hydrosulfuric acid is ..... 10.....
11. Diamonds are a crystalline form of ..... 11.....
12. A solution with a pH of 4 will turn litmus ..... 12.....
13. Mercury forms two oxides,  $\text{HgO}$  and ..... 13.....
14. The mineral galena is a source of the metal ..... 14.....
15. The reaction between water and ..... yields acetylene. 15.....
16. Sulfur dioxide is the anhydride of ..... acid. 16.....
17. Equal volumes of molar solutions of hydrochloric acid and sodium hydroxide were mixed and the mixture evaporated to dryness. The residue was ..... 17.....
18. At standard conditions, the volume of one gram-molecular weight of ethane is ..... liters. 18.....
19. Oxygen is obtained commercially by the fractional distillation of ..... 19.....
20. The destructive distillation of soft coal yields the solid product called ..... 20.....

## RECENT EXAMINATION PAPERS

21. The valence of sulfur in the compound  $\text{H}_2\text{SO}_4$  is ..... 21\_\_\_\_\_
22.  $\text{Cl}^-$  — 1e yields ..... 22\_\_\_\_\_
23. The element germanium (Ge) reacts with iodine to form the compound represented by the formula ..... 23\_\_\_\_\_
24. A nitrogen compound used as a refrigerant is ..... 24\_\_\_\_\_
25. The gas produced by the reaction between calcium sulfide and hydrochloric acid is ..... 25\_\_\_\_\_

Write on the line at the right of *each* statement the *number* preceding the term that best completes the statement. [15]

26. Heating limestone produces (1)gypsum (2)plaster of Paris (3)quicklime (4)slaked lime 26\_\_\_\_\_
27. The general formula for the members of the methane series of hydrocarbons is (1) $\text{C}_n\text{H}_n$  (2) $\text{C}_n\text{H}_{2n-2}$  (3) $\text{C}_n\text{H}_{2n}$  (4) $\text{C}_n\text{H}_{2n+2}$  27\_\_\_\_\_
28. A deliquescent compound is (1)NaCl (2) $\text{CaCl}_2$  (3) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (4) $\text{MgSO}_4$  28\_\_\_\_\_
29. A galvanized iron tub would be readily damaged if it were used to store a solution of (1)copper sulfate (2)lithium sulfate (3)potassium nitrate (4)sodium chloride 29\_\_\_\_\_
30. The ratio by weight of sulfur to oxygen in sulfur dioxide is (1)1:2 (2)2:1 (3)1:1 (4)1:3 30\_\_\_\_\_
31. Salts of the metal (1)barium (2)chromium (3)mercury (4)silver decompose when they absorb light. 31\_\_\_\_\_
32. Glass may be etched with (1) $\text{HNO}_3$  (2)HF (3)HI (4) $\text{H}_2\text{SO}_4$  32\_\_\_\_\_
33. Solutions of copper sulfate and sodium chloride were mixed and the mixture evaporated to dryness. The residue contained molecules of (1)NaCl,  $\text{Na}_2\text{SO}_4$ ,  $\text{CuCl}_2$  and  $\text{CuSO}_4$  (2) $\text{CuSO}_4$  and NaCl (3) $\text{CuCl}_2$  and  $\text{Na}_2\text{SO}_4$  (4) $\text{CuSO}_4$ , NaCl and  $\text{Na}_2\text{SO}_4$  33\_\_\_\_\_
34. Silver coins contain the elements silver and (1)copper (2)gold (3)nickel (4)zinc 34\_\_\_\_\_
35. Cast iron is made in the (1)blast furnace (2)electric furnace (3)open hearth furnace (4)reverberatory furnace 35\_\_\_\_\_
36. A type of hard water that can be softened by boiling contains the compound (1) $\text{CaSO}_4$  (2) $\text{Ca}(\text{HCO}_3)_2$  (3) $\text{MgSO}_4$  (4) $\text{MgCl}_2$  36\_\_\_\_\_
37. The reaction between an alcohol and an acid is called (1)hydrolysis (2)hydrogenation (3)esterification (4)saponification 37\_\_\_\_\_
38. The arrangement of the atoms according to their atomic numbers as determined by X-ray spectra was made by (1)Faraday (2)Lavoisier (3)Moseley (4)Mendelejeff 38\_\_\_\_\_
39. A carbon compound that is used to extinguish fires is (1) $\text{CHCl}_3$  (2) $\text{C}_2\text{H}_5\text{OH}$  (3) $\text{CCl}_4$  (4) $\text{CH}_3\text{COOH}$  39\_\_\_\_\_
40. Nitrogen fixation is accomplished in the (1)contact (2)Frasch (3)Haber (4)Solvay process. 40\_\_\_\_\_

## RECENT EXAMINATION PAPERS

In *some* of the following statements the term in italics makes the statement incorrect. For each *incorrect* statement, write on the line at the right the term that must be substituted for the italicized term to make the statement correct. For each *correct* statement, write the word *true* on the line at the right. [10]

41. The formula of sodium hypochlorite is *NaClO*. 41.....
42. Metallic sodium is usually stored in *water*. 42.....
43. All the amalgams contain the element *silver*. 43.....
44. In the preparation of oxygen from  $\text{KClO}_3$ , the  $\text{MnO}_2$  acts as a (an) *oxidizing agent*. 44.....
45. When lead sulfide ore is roasted, *lead oxide* and sulfur dioxide are produced. 45.....
46. Under identical conditions, helium is *four times* as dense as hydrogen. 46.....
47. The products formed when hard coal is burned completely weigh *the same* as the original coal. 47.....
48. A water solution of aluminum nitrate contains *twice* as many nitrate ions as aluminum ions. 48.....
49. The electrolysis of fused sodium chloride yields *sodium* and chlorine. 49.....
50. To burn completely 24 grams of hydrogen, *twenty-four* grams of oxygen are needed. 50.....

### Part II

*Answer five questions from part II. [Show all numerical work when computation is required.]*

1. a. Make a labeled diagram of the apparatus used for the laboratory preparation and collection of
- (1) A gas that is more dense than air and very soluble in water [2]
- (2) A gas that is less dense than air and slightly soluble in water [2]
- b. Write balanced equations that represent the reactions for the preparation of the gases you used in answering a. [4]
- c. Both collection methods provide impure samples. For *each* collection method diagramed in answer to a, name the chief impurity that is found in the collected sample. [2]
2. a. Write an equation that represents the reaction between aluminum powder and ferric oxide. [2]
- b. Basing your answer solely on heats of formation, explain why tin (Sn) can not be used in place of aluminum in part a. [2]
- c. Zinc reacts with hot dilute and with hot concentrated sulfuric acid. Explain why these reactions yield different gases. [2]



## RECENT EXAMINATION PAPERS

- d. Select the reactions that go to completion and explain why each one selected is not reversible: [4]
- (1) The reaction between sulfuric acid and a solution of sodium bicarbonate
  - (2) The reaction between solutions of potassium bromide and silver nitrate
  - (3) The reaction between solutions of copper sulfate and sodium chloride
  - (4) The reaction between solutions of potassium chloride and barium nitrate
3. a. Name the compound indicated by *each* of these procedures:
- (1) The substance forms a blue incrustation when it is heated with cobalt nitrate on a charcoal block. It forms a white precipitate, insoluble in hydrochloric acid, when it reacts with barium chloride. [2]
  - (2) The substance forms a deep blue precipitate when it reacts with potassium ferrocyanide. It reacts with silver nitrate to form a white precipitate, insoluble in nitric acid but soluble in ammonium hydroxide. [2]
  - (3) The substance, fused with borax, yields a green-colored bead. It reacts with hydrochloric acid to form carbon dioxide. [2]
- b. You are given solutions of the sulfates of copper and iron, a strip of copper and an iron nail. Describe procedures by which you may determine their relative positions in the electromotive force series. [2]
- c. Describe simple laboratory procedures by which you may obtain solid zinc chloride from a hydrogen-generating flask that contains hydrochloric acid and an excess of zinc. [2]
4. a. Make separate, clearly labeled diagrams that represent the probable structure of the atom and of the ion of the element of atomic number 11. [2]
- b. Name *two* other elements that have physical and chemical properties very similar to those of the element of atomic number 11. [2]
- c. Which of these elements — elements of atomic numbers 3, 8, 13, 17, 54 — combine readily with the element of atomic number 11? [2]
- d. Boron has isotopes of atomic weights 10 and 11. Explain how the atoms of these isotopes differ in structure. [2]
- e. Explain why a water solution of sodium carbonate turns litmus blue. [2]
5. a. Describe a process by which silicon carbide (carborundum) or calcium carbide is made. [3]
- b. Describe the process by which pure copper is obtained from a bar of impure copper. [4]
- c. Find the theoretical maximum weight of aluminum that might be obtained from 3000 pounds of a bauxite ore that contains 65% aluminum oxide. [3]

## RECENT EXAMINATION PAPERS

6. *a.* Write a balanced equation for the formation of nitroglycerine (glyceryl nitrate) based on a reaction between nitric acid and glycerin. [2]
- b.* What volume of oxygen is required to burn completely 250 liters of ethane?  $[2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O]$  [3]
- c.* Name the materials needed and describe the process that may be used to produce and separate ethyl alcohol in the school laboratory. [3]
- d.* Write the structural (graphic) formula that represents (1) methane, (2) difluorodichloromethane (Freon). [2]
7. *a.* A pupil obtained the following data in an experimental determination of the percentage of water of crystallization in crystalline barium chloride:
- |   |            |
|---|------------|
| weight of empty crucible .....  | 30.2 grams |
| weight of barium chloride crystals and crucible<br>(before heating) ..... | 45.2 grams |
| weight of crucible and anhydrous barium chloride<br>(after heating) ..... | 43.0 grams |
- (1) What percentage of water of crystallization is shown by these data? [2]
- (2) What further laboratory procedure, if any, would you suggest before accepting this figure as the value for the percentage of water of crystallization in crystallized barium chloride? [2]
- b.* Write an ionic equation that represents the neutralization of barium hydroxide by hydrochloric acid. [2]
- c.* At 20°C, 100 grams of water dissolves 35.7 grams of the anhydrous barium chloride; at 40°C, 40.7 grams; at 60°C, 46.4 grams; and at 80°C, 52.4 grams.
- (1) How would you proceed to form one liter of saturated solution of barium chloride at 40°C? [2]
- (2) What change would take place in the solution of barium chloride prepared in answer to (1) if the solution temperature were (a) changed to 60°C, (b) changed to 20°C? [2]
8. *a.* Write the equation that represents the laboratory preparation of bromine. [2]
- b.* Bromine is prepared commercially by reacting chlorine gas with magnesium bromide obtained from sea water. How many liters of chlorine gas will be needed to react completely with 920 grams of magnesium bromide in this process? [3]
- $$[MgBr_2 + Cl_2 \longrightarrow MgCl_2 + Br_2]$$
- c.* Bleaching can be either a reducing action or an oxidizing action. Give an example of (1) a substance that bleaches by reduction, (2) a substance that bleaches by oxidation. [2]
- d.* Describe briefly a laboratory procedure for making each of three allotropic forms of sulfur. [3]

# CHEMISTRY

Tuesday, January 29, 1946 — 1.15 to 4.15 p. m., only

Answer all questions in part I and five questions from part II. Answers to the questions in part I should be written on the question paper as directed and handed in with the other answer paper. Answers should be numbered and lettered to correspond with the questions.

## Part I

Answer all questions in part I.

Write on the line at the right of each statement the word or expression which, when inserted in the blank, will make the statement true. [20]

1. The valence of tin, in the compound  $\text{SnO}_2$ , is ..... 1.....
2. The reaction between an alcohol and an acid yields water and a (an) ..... 2.....
3. The nucleus of an atom consists of 42 protons and 54 neutrons. The chemical symbol of this atom is ..... 3.....
4. The reaction between acetic acid and aluminum yields hydrogen and ..... 4.....
5. Fluorine collects at the ..... electrode during the electrolysis of a water solution of potassium fluoride. 5.....
6. The formula that represents rubidium (Rb) carbonate is ..... 6.....
7. A water solution of pH 5 has a higher concentration of ..... ions than a water solution of pH 10. 7.....
8.  $\text{Ca}^{++} + 2e \rightarrow$  ..... 8.....
9. An ion that has a ..... charge is formed whenever a sodium atom loses an electron. 9.....
10. The general formula that represents the compounds of the methane series is  $\text{C}_n\text{H}_{2n+2}$ . The formula that represents the fifth member of the series (pentane) is ..... 10.....
11. Ozone has a molecular weight of ..... 11.....
12. Thirty-two grams of hydrogen react with ..... grams of oxygen to form water. 12.....
13. A molecule of  $\text{Mg}_3(\text{PO}_4)_2$  contains ..... atoms. 13.....
14. The molecular weight of  $\text{Fe}_2(\text{CO}_3)_3$  is ..... 14.....
15. Chlorine may be prepared by the reaction between manganese dioxide and ..... 15.....
16. The reaction between calcium oxide and carbon yields carbon monoxide and ..... 16.....
17. The impure copper is made the ..... electrode in the electrolytic refining of copper. 17.....
18. Potassium ferrocyanide is used to detect the presence of the ..... ion. 18.....
19. The allotropic form of carbon used commercially as an abrasive is ..... 19.....

## RECENT EXAMINATION PAPERS

20. One liter of nitrogen reacts with three liters of hydrogen to form ..... liters of ammonia. 20.....
- Write on the line at the right of *each* statement the *number* preceding the term that best completes the statement. [19]
21. Soap forms an insoluble compound when it reacts with a solution of (1)  $\text{Na}_2\text{CO}_3$  (2)  $\text{KHCO}_3$  (3)  $\text{MgSO}_4$  (4)  $\text{NH}_4\text{Cl}$  21.....
22. The reactions between water and nonmetallic oxides yield (1) acids (2) acid anhydrides (3) bases (4) salts 22.....
23. Nitric acid, when it reacts with zinc, is (1) catalyzed (2) dehydrated (3) oxidized (4) reduced 23.....
24. A formula that represents an acid is (1)  $\text{CH}_3\text{CHO}$  (2)  $(\text{CH}_3)_2\text{O}$  (3)  $\text{CH}_3\text{CO}$  (4)  $\text{CH}_3\text{COOH}$  24.....
25. An important alloy composed of the metals Al, Cu and Mg is called (1) bronze (2) duralumin (3) magnalium (4) monel metal 25.....
26. Sand will react with a water solution of the compound (1)  $\text{HNO}_3$  (2)  $\text{H}_2\text{SO}_4$  (3)  $\text{HF}$  (4)  $\text{HC}_2\text{H}_3\text{O}_2$  26.....
27. The weight of 22.4 liters of carbon monoxide at standard conditions is (1) 2 grams (2) 4 grams (3) 16 grams (4) 28 grams 27.....
28. The reaction between sea water and (1) bromine (2) chlorine (3) hydrogen bromide (4) hydrogen iodide yields the gas bromine. 28.....
29. The molecular weight of heavy water (deuterium oxide) is (1) 18 (2) 20 (3) 22 (4) 34 29.....
30. A neutral salt is formed by the reaction between sulfuric acid and (1) zinc hydroxide (2) sodium hydroxide (3) ammonium hydroxide (4) copper hydroxide 30.....
31. The energy released when uranium atoms disintegrate results from (1) the change of uranium to radium (2) the emission of high-speed electrons (3) the conversion of matter to energy (4) the formation of the isotope U235 31.....
32. The reaction between copper and sulfur, heated, yields the compound copper (1) sulfate (2) sulfide (3) sulfite (4) hyp-sulfite 32.....
33. The reaction between steam and hot iron yields (1) oxygen (2) hydrogen (3) ozone (4) hydrogen peroxide 33.....
34. Iodine is *least* soluble in (1) alcohol (2) carbon disulfide (3) carbon tetrachloride (4) water 34.....
35. A metal that forms a self-protecting film when it corrodes is (1) aluminum (2) cast iron (3) gold (4) wrought iron 35.....
36. A gas that is collected by the displacement of water in the laboratory is (1) ammonia (2) chlorine (3) nitric oxide (4) sulfur dioxide 36.....
37. By increasing the active surface area of the reactants, the rate at which chemical reactions occur is (1) unaffected (2) increased (3) decreased (4) tripled 37.....

## RECENT EXAMINATION PAPERS

38. Drinking water may be made from sea water by (1) adding alum (2) distilling the water (3) aerating the water (4) filtering the water 38.....
39. A lead-lined container should not be used to store solutions of (1)  $\text{Na}_2\text{SO}_4$  (2) concentrated  $\text{H}_2\text{SO}_4$  (3)  $\text{AgNO}_3$  (4)  $\text{ZnCl}_2$  39.....
- In *some* of the following statements the term in italics makes the statement incorrect. For each *incorrect* statement, write on the line at the right the term that must be substituted for the italicized term to make the statement correct. For each *correct* statement, write the word *true* on the line at the right. [11]
40. The alloys called bronzes all contain the element *aluminum*. 40.....
41. Avogadro's hypothesis states that equal volumes of gases, under like conditions, contain the same number of *atoms*. 41.....
42. A carbon compound frequently used as a fire-extinguishing material is *carbon monoxide*. 42.....
43. The dissociation of one molecule of  $(\text{NH}_4)_3\text{PO}_4$  yields *two* ions. 43.....
44. The number of extranuclear electrons in any atom equals the number of *neutrons* in the atom. 44.....
45. The valence of the inert elements is *four*. 45.....
46. Salts of nitrous acid are called *nitrates*. 46.....
47. Methanol, glycerin and ethylene glycol belong to a class of organic compounds called *alcohols*. 47.....
48. Repeated analyses of samples of red iron oxide, from widely separated locations, show that the compound consists of 7 parts of iron and 3 parts of oxygen, by weight. These data illustrate the law of *conservation of matter*. 48.....
49. The incomplete combustion of hydrogen sulfide yields water and *sulfur dioxide*. 49.....
50. A coating of silver *oxide* forms when silver tarnishes. 50.....

### Part II

*Answer five questions from part II. [Show all numerical work when computation is required.]*

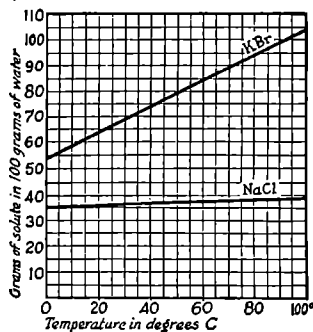
1. *a.* Make a diagram that represents the apparatus for the laboratory preparation and collection of oxygen by a method that employs a solid catalyst. [4]
- b.* Write a balanced equation that represents the reaction that occurs in *a.* [2]
- c.* A pupil prepares oxygen by the method indicated in *a.*, using chemically pure compounds. He then places the remaining contents of the generator in distilled water; after agitation, the mixture is filtered. Tell how the filtrate may be tested to show that a new compound has been formed. [2]
- d.* Write a balanced equation that represents another reaction by which oxygen is prepared in the laboratory. [2]

## RECENT EXAMINATION PAPERS

2. *a.* Explain how slaked lime is prepared from limestone. [2]
- b.* A blast furnace charge consists of iron ore ( $\text{Fe}_2\text{O}_3$ ), coke and limestone. As the furnace operates, the iron ore is reduced; the limestone, converted to the oxide, reacts with the silicon dioxide impurities to form slag. Write a balanced equation that represents
- (1) The reduction of the ore [2]
  - (2) The formation of the slag [2]
- c.* (1) Make a labeled diagram that represents the apparatus used to extract aluminum from its ore. [2]
- (2) Name the ore that is the source of the metal and name the solvent used for the  $\text{Al}_2\text{O}_3$ . [2]
3. *a.* Write a balanced equation for the preparation of each of *three* gases often made in the school laboratory. Each equation must represent a reaction between two of these compounds:  $\text{Na}_2\text{CO}_3$ ,  $\text{NaOH}$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{ZnS}$ ,  $(\text{NH}_4)_2\text{SO}_4$ ,  $\text{KCl}$  [6]
- b.* Find the volume of sulfur dioxide that can be obtained by the reaction between 150 grams of sodium sulfite and sulfuric acid. [4]
4. *a.* Explain the cracking of petroleum, in terms of its effect on the molecules. [2]
- b.* Write the structural (graphic) formula\* that represents (1) ethane, (2) acetylene. [2]
- c.* Find the volume of oxygen needed to burn completely 30 liters of propane. [ $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$ ] [2]
- d.* Describe briefly the process by which each of the following is made commercially: (1) water gas, (2) coke. [4]
5. Describe briefly *laboratory procedures* that could be used for *five* of the following: [10]
- a.* To dry hydrogen gas
  - b.* To demonstrate that zinc is more active than copper
  - c.* To show that a white powder is a sulfate
  - d.* To prepare ferric hydroxide
  - e.* To demonstrate that hydrogen is a reducing agent
  - f.* To change ferrous to ferric chloride
6. *a.* Write the ionic equation that represents the reaction between water solutions of magnesium sulfate and potassium nitrate. [2]
- b.* Describe a simple procedure by which the poor conductivity of dilute acetic acid may be demonstrated. [2]
- c.* Make diagrams of the isotopes of calcium of atomic weights 40 and 42. [2]
- d.* Explain why sodium and caesium (Cs) are members of the same chemical "family." [2]
- e.* Describe briefly the electron transfer that occurs at the cathode of a cell in which a water solution of cadmium chloride is electrolyzed. [2]

## RECENT EXAMINATION PAPERS

7. *a.* (1) Describe briefly the processes used to obtain sulfur trioxide from sulfur. [2]  
 (2) Explain why the commercial process for the manufacture of sulfuric acid does not employ the reaction between water and the trioxide. [1]
- b.* Describe and explain the reaction between cane sugar and concentrated sulfuric acid. [2]
- c.* Write the balanced equation that represents the reaction between copper and hot concentrated sulfuric acid. [2]
- d.* Find the weight of sulfuric acid needed to react completely with 250 pounds of calcium phosphate to yield the superphosphate.  
 $[2\text{H}_2\text{SO}_4 + \text{Ca}_3(\text{PO}_4)_2 \rightarrow \text{Ca}(\text{H}_2\text{PO}_4)_2 + 2\text{CaSO}_4]$  [3]
8. *a.* Write an ionic equation that represents the reaction between sodium hydroxide and nitric acid. [2]
- b.* Explain why the product that collects at the cathode when a water solution of sodium chloride is electrolyzed differs from the product that collects at the cathode when fused sodium chloride is electrolyzed. [2]
- c.* Explain why a water solution of potassium carbonate is basic. [2]
- d.* Describe simple procedures by which a given salt may be identified as sodium nitrate. [2]
- e.* A solution of potassium bromide, saturated at 50° C, is made by dissolving 79 grams of the salt in 100 grams of water. This solution is heated to 70° C. Find the weight of the salt that must be dissolved to yield a saturated solution at the new temperature. [2]



# CHEMISTRY

Tuesday, June 18, 1946—1.15 to 4.15 p. m., only

Answer all questions in part I and five questions from part II. Answers to the questions in part I should be written on the question paper as directed and handed in with the other answer paper. Answers should be numbered and lettered to correspond with the questions.

## Part I

Answer all questions in part I.

Write on the line at the right of *each* statement the word or expression which, when inserted in the blank, will make the statement true. [25]

1. When manganese dioxide is heated with potassium chlorate, oxygen is produced. The source of the oxygen is the ..... 1.....
2. The completed innermost orbit of an atom contains \_\_\_\_\_ electrons. 2.....
3. The chemical reaction in which hydrogen ions and hydroxyl ions unite to form water is known as \_\_\_\_\_. 3.....
4. A solution of pH 3 is \_\_\_\_\_ to indicators. 4.....
5. In the modern periodic table, elements are classified according to their atomic \_\_\_\_\_. 5.....
6. To produce acetylene, add water to \_\_\_\_\_. 6.....
7. The fractional distillation of liquid air yields \_\_\_\_\_, which is used in welding. 7.....
8. Copper reacts with dilute nitric acid to form the colorless gas \_\_\_\_\_ which turns brown on exposure to air. 8.....
9. The halogen that is a solid at room temperature is \_\_\_\_\_. 9.....
10. Methane is 75% carbon and 25% \_\_\_\_\_. 10.....
11. The lower the atomic number of a halogen, the \_\_\_\_\_ its activity. 11.....
12.  $\text{Fe}^{+++} + e \longrightarrow$  \_\_\_\_\_. 12.....
13. When potassium liberates hydrogen from water, a solution of \_\_\_\_\_ remains. 13.....
14. The valence of nitrogen, in the compound  $\text{NaNO}_3$ , is ..... 14.....
15. Calcium oxide is the anhydride of \_\_\_\_\_. 15.....
16. Equal volumes of gases under like conditions of temperature and pressure contain equal numbers of \_\_\_\_\_. 16.....
17. If 50 cc of hydrogen are produced by the electrolysis of water, \_\_\_\_\_ cc of oxygen will also be formed. 17.....
18. When steam reacts with hot carbon, the element released is \_\_\_\_\_. 18.....



## RECENT EXAMINATION PAPERS

19. Solder, type metal and Wood's metal contain the elements lead and \_\_\_\_\_ 19\_\_\_\_\_
20. The number of atoms in a molecule of  $Zn(NO_3)_2$  is \_\_\_\_\_ 20\_\_\_\_\_
21. Carbon is burned. The product is dissolved in water. The compound formed is \_\_\_\_\_ 21\_\_\_\_\_
22. Nitric acid is prepared by the reaction between sulfuric acid and sodium \_\_\_\_\_ 22\_\_\_\_\_
23. The source of the carbon dioxide produced from baking powders is the compound \_\_\_\_\_ 23\_\_\_\_\_
24. The formulas of two different oxides of antimony are  $Sb_2O_3$  and \_\_\_\_\_ 24\_\_\_\_\_
25. Silicon dioxide reacts with carbon to form carbon monoxide and \_\_\_\_\_ 25\_\_\_\_\_

Write on the line at the right of *each* statement the *number* preceding the term that best completes the statement. [25]

26. The number of molecules of sulfuric acid that can be neutralized by one molecule of calcium hydroxide is (1) one (2) two (3) three (4) four 26\_\_\_\_\_
27. Sulfuric acid is used in the preparation of hydrochloric acid because (1) it is an oxidizing agent (2) it is a reducing agent (3) it has a higher boiling temperature (4) it is a dehydrating agent 27\_\_\_\_\_
28. Chemically pure water may be prepared by (1) chlorination (2) coagulation (3) distillation (4) filtration 28\_\_\_\_\_
29. Concentrated sulfuric acid reacts with a white solid to produce a gas. The white solid (1) can not be a sulfite (2) may be a sulfite (3) must be a sulfite (4) is a sulfate 29\_\_\_\_\_
30. An example of a strong electrolyte is (1)  $NaOH$  (2)  $H_2O$  (3)  $NH_4OH$  (4)  $HC_2H_3O_2$  30\_\_\_\_\_
31. Stainless steel usually contains (1) aluminum (2) chromium (3) tin (4) zinc 31\_\_\_\_\_
32. When a mixture of olive oil, soap and water is thoroughly shaken (1) a compound (2) an emulsion (3) a solution (4) a tincture is formed. 32\_\_\_\_\_
33. The name of a famous American chemist is (1) Carver (2) Dalton (3) Moseley (4) Priestley 33\_\_\_\_\_
34. A diamond strongly heated in an excess of oxygen will (1) form only carbon dioxide (2) merely melt (3) not react with the oxygen (4) produce ozone 34\_\_\_\_\_
35. Copper reacts with hot concentrated sulfuric acid to form the gas (1)  $H_2$  (2)  $H_2S$  (3)  $SO_2$  (4)  $SO_3$  35\_\_\_\_\_
36. When zinc sulfide ore is heated in air,  $SO_2$  and (1)  $Zn$  (2)  $ZnO$  (3)  $ZnSO_3$  (4)  $Zn_3N_2$  are formed. 36\_\_\_\_\_
37. Esters are produced by the reaction between an alcohol and (1) an acid (2) a carbohydrate (3) a fat (4) a hydrocarbon 37\_\_\_\_\_

## RECENT EXAMINATION PAPERS

38. A salt whose water solution is acid is (1)  $\text{Cu}(\text{NO}_3)_2$  (2)  $\text{K}_2\text{SO}_4$  (3)  $\text{NaCl}$  (4)  $\text{Ca}(\text{HCO}_3)_2$  38.....
39. Carbon monoxide is poisonous because it (1) causes severe throat irritation (2) destroys lung tissue (3) has no odor (4) prevents the red cells from distributing oxygen 39.....
40. The molecular weight of  $\text{Ca}(\text{OH})_2$  is (1) 57 (2) 58 (3) 74 (4) 114 40.....
41. The element of atomic number 19 is probably (1) amphoteric (2) inert (3) metallic (4) nonmetallic 41.....
42. A gas that may be used as an oxidizing agent is (1) ammonia (2) chlorine (3) helium (4) hydrogen 42.....
43. During the electrolysis of a copper sulfate solution, the copper ions (1) gain electrons (2) lose electrons (3) share electrons (4) neither gain nor lose electrons 43.....
44. Superheated water is used in the Frasch process to (1) clean the pipes (2) dissolve the impurities (3) dissolve the sulfur (4) melt the sulfur 44.....
45. A double replacement reaction will go to completion when one of the products is (1) barium sulfate (2) lead nitrate (3) potassium bromide (4) sodium chloride 45.....
46. Superphosphate is made by treating calcium phosphate with (1) calcium hydroxide (2) hydrochloric acid (3) sodium hydroxide (4) sulfuric acid 46.....
47. A substance that illustrates the law of definite proportions is (1) bronze (2) a saturated solution (3) air (4) sugar 47.....
48. Naphthalene is used as a (1) cleansing agent (2) moth repellent (3) paint remover (4) solvent 48.....
49. A source of lead is (1) bauxite (2) galena (3) malachite (4) pyrites 49.....
50. Water gas is a mixture of hydrogen and (1) ammonia (2) carbon monoxide (3) sulfur dioxide (4) water vapor 50.....

### Part II

*Answer five questions from part II. [Show all numerical work when computation is required.]*

1. a. Make a diagram of the apparatus used for the laboratory preparation and collection of ammonia. Indicate the materials used. [4]
- b. State *two* properties of ammonia that determine the laboratory method of collection. [2]
- c. Write the equation that represents the reaction between ammonia and hydrochloric acid. [2]
- d. Describe a laboratory test for an ammonium salt. [2]

## RECENT EXAMINATION PAPERS

2. Write a balanced equation for each of *five* of the following: [10]
- The electrolysis of a water solution of sodium chloride
  - The reaction between ferric chloride and ammonium hydroxide
  - The reaction between concentrated sulfuric acid and a mixture of potassium iodide and manganese dioxide
  - The reaction between calcium hydroxide and aluminum nitrate
  - The reaction between sodium carbonate and magnesium chloride
  - The reaction between zinc and copper nitrate
3. *a.* Make separate and clearly labeled diagrams that represent the probable structure of the atom and of the ion of magnesium. [2]
- b.* Chlorine has isotopes of atomic weights 35 and 37. Make a diagram of the atomic structure of *each* isotope. [2]
- c.* In the reaction  $2\text{Na} + 2\text{H}_2\text{O} \longrightarrow 2\text{NaOH} + \text{H}_2$ , name (1) the element that is oxidized, (2) the element that is reduced. [2]
- d.* Write an ionic equation that represents the neutralization of a solution of sodium hydroxide by hydrochloric acid. [2]
- e.* Explain why a solution of sodium carbonate turns red litmus blue. [2]
4. Solve *two* of the following problems:
- How many liters of nitrogen react with 75 liters of hydrogen to produce ammonia? [5]
  - How many pounds of chlorine are produced when 100 pounds of hydrochloric acid react with manganese dioxide? [5]
  - Find the volume in liters of the carbon dioxide liberated when 20 grams of calcium carbonate are decomposed by heat. [5]
5. *a.* Hydrogen is passed over hot iron oxide. Name the two products formed. [2]
- b.* Describe the chemical reaction that occurs when a soda-acid fire extinguisher is inverted. [2]
- c.* Describe a procedure that could be used to distinguish between a saturated and an unsaturated solution of a salt. [2]
- d.* A pupil prepares hydrogen in the laboratory. Explain why the gas first collected may explode when ignited. [2]
- e.* When some anhydrous copper sulfate is dropped into a test tube containing ethyl alcohol, the copper sulfate changes color. State the color change and name the compound that causes this change. [2]
6. *a.* Name the compound indicated by *each* of these procedures:
- (1) The substance forms a white precipitate when it reacts with hydrogen sulfide. It reacts with silver nitrate to form a white precipitate insoluble in nitric acid. [2]
  - (2) The substance, fused with borax, forms a blue bead. A brown ring is formed when a solution of this substance, mixed with a solution of ferrous sulfate, is in contact with concentrated sulfuric acid. [2]
  - (3) The substance colors a Bunsen flame violet. It reacts with hydrochloric acid to form carbon dioxide. [2]

## RECENT EXAMINATION PAPERS

- b.* Explain the reaction that occurs when sodium carbonate is added to "permanent" hard water. [2]
- c.* When carbon dioxide reacts with lime water, a white precipitate is formed. Explain why an excess of carbon dioxide will cause the precipitate to disappear. [2]
7. *a.* Make a labeled diagram that represents the laboratory preparation and collection of bromine. [4]
- b.* Write an ionic equation that represents the reaction between chlorine and potassium bromide. [2]
- c.* Explain how sulfurous acid may be prepared in the laboratory from sulfur. [2]
- d.* What change will occur in the volume of concentrated sulfuric acid when it is stored in an open graduated cylinder? State a reason for your answer. [2]
8. *a.* Describe briefly the manufacture of soap. [2]
- b.* Name the important by-product of this process and give a use of it. [2]
- c.* Write a formula that represents *each* of the following: (1) an alcohol, (2) a carbohydrate, (3) an ester, (4) a hydrocarbon. [4]
- d.* Write the structural (graphic) formula that represents (1) propane, (2) benzene. [2]

# INDEX \*

- Absolute humidity, 110  
 Absolute temperature scale, 280  
 Absolute zero, 280  
 Acetic acid, 181  
 Acetylene, 169; series, 177  
 Acheson, 145  
 Acid anhydride, 32, 77  
 Acids, properties of, 69, 78; chemical activity of, 70; definition of, 77; preparation of, 78; method of naming, 81-82; organic, 181-182  
 Acid salts, 81  
 Activity series of metals, *see* "Electromotive series"  
 Adsorption, 270  
 Aération, 16, 36  
 Air, liquid, 13, 111; ocean of, 103; composition of, 103; is a mixture, 104; determination of percentage of oxygen in, 104-105; four essential constituents of, 105-106; water vapor in, 110; inert gases in, 110-111  
 Air-slaked lime, 212  
 Alcohol, methyl, 179; wood, 179; ethyl, 179-180; grain, 180; denatured, 180  
 Aldehydes, 181  
 Alkali metals, 199 *ff.*  
 Allotropic forms, 16; of oxygen, 17; of sulfur, 128-129; of carbon, 145  
 Alloys, 195; properties of, 195-196; of steel, 227; table of, 342  
 Alpha rays, 250  
 Aluminum, 234 *ff.*; preparation of, 234; properties of, 235; uses of, 236; oxide, 236; hydroxide, 237; silicates, 237  
 Alums, 237  
 Ammonia, occurrence of, 114; preparation of, 114-115; properties of, 117; uses of, 117-118  
 Ammonium radical, 115; test for, 118  
 Ammonium salts, 119  
 Amphoteric elements, 58, 131  
 Analysis, 32, 52; of water, 35  
 Anhydrides, 32, 76-77  
 Anhydrous compounds, 40  
 Aniline dyes, 264  
 Anions, 66 *ff.*  
 Anode, 66 *ff.*  
 "Anti-freeze," 181  
 Antimony, 247  
 Aqua regia, 94, 122, 250  
 Argon, 103, 110-111  
 Arrhenius, 65  
 Arsenic, 247  
 Artificial silk, *see* "Silk substitute"  
 Asbestos, 264; platinized, 137  
 Atom, early conception of, 45; definition of, 47; structure according to Electron Theory, 55; compared with ion, 68; energy within, 275  
 Atomic numbers, 56  
 Atomic Theory, 45  
 Atomic weight, 47; according to Electron Theory, 56  
 Avogadro's Law, 285  
 Babbitt metal, 246, 247  
 Bacteria, in nitrogen cycle, 108; in decay, 15-16, 257-258  
 Bakelite, 181, 276  
 Baking powder, 155, 206  
 Baking soda, 155, 206  
 Balanced aquarium, 109  
 Barium, flame test, 201  
 Bases, properties of, 68, 79; chemical activity of, 70; definition of, 79; preparation of, 79; method of naming, 82  
 Basic anhydride, 32, 77  
 Basic salt, 81  
 Bauxite, 234 *ff.*  
 Benzene, 178  
 Benzine series, 164  
 Bergius Process, *see* "Hydrogenation of coal"  
 Beryllium, 274, 276-277  
 Bessemer converter, 223-224; uses of steel, 224  
 Beta rays, 250  
 Binary acid, 81  
 Binary salt, 82  
 Bismuth, 247  
 Blast furnace, 221-222  
 Bleaching, with ozone, 17; with chlorine 8, 90, 91; with bleaching powder, 91, 215-216; with bromine, 96; with sulfur dioxide, 136  
 Bleaching powder, 91, 215-216  
 Blue prints, 231  
 "Boiler scale," 146, 214  
 Boiling points, abnormal, 72  
 Boneblack, 140-147  
 Bordeaux mixture, 242  
 Boyle's law and its applications, 281-283  
 Brass, 242, 343  
 Bromides, test for, 100  
 Bromine, occurrence of, 94; preparation of, 94-95; properties of, 95-96; uses of, 96  
 Bronze, 242, 343  
 Brownian movement, 269  
 Bunsen burner, 170; flame, 171  
 Burning, changes of weight resulting from, 5; need of oxygen for, 16  
 Burton Process, *see* "Cracking of oils"  
 Calcium, carbide, 150; flame test, 201; metallic, 210; carbonate, 210-211; oxide, 212, 213; sulfate, 213; phosphate, 215; chloride, 215  
 Calomel, 245  
 Calorie, 255  
 Calorimeter, 255-256  
 Candle flame, 171  
 Cane sugar, 186-187; test for, 254  
 Caramel, 187  
 Carbohydrates, 184-187; charring by sulfuric acid, 139; in food, 253, 254  
 Carbolic acid, 181  
 Carbon, importance of, 143; occurrence of, 143-145; allotropic forms of, 145; crystalline forms of, 145; amorphous forms of,

\*Consult also the list of *Important Definitions* and other tabular material provided in the Appendix.

- 146; properties of, 148-149; as a reducing agent, 148, 193; oxides of, 153 *f.*  
 Carbonated beverages, 155  
 Carbonates, test for, 157  
 Carbon dioxide, cycle, 109; occurrence of, 153; in the air, 153; preparation of, 154; properties of, 155; "snow," 155; uses of, 155; fire extinguisher, 156; test for, 157, comparison with carbon monoxide, 160-161  
 Carbon disulfide, 151  
 Carbonic acid, 155  
 Carbon monoxide, preparation of, 158-159; properties of, 159-160; uses of 160; test for, 160; comparison with carbon dioxide, 160-161  
 Carbon tetrachloride, 179  
 Carbonyls, 159  
 Carborundum, 150  
 Carboxyl group, 181  
 Cassiterite, 246  
 Castile soap, 184  
 Cast iron, 222  
 Catalytic agent, 12; water as a, 32  
 Cathode, 66 *f.*  
 Cations, 66 *f.*  
 Caustic potash, 207  
 Caustic soda, 204  
 Cell, of living organism, 258  
 Cellophane, 186  
 Celluloid, 186  
 Cellulose, 185-186  
 Cement, 238, 271  
 Ceramics, 237-238  
 Cerium, 277  
 Cesium, 199  
 Chalk, 210, 211; precipitated, 211  
 Chamber process for making sulfuric acid, 136  
 Change, physical, 2; chemical, 2  
 Charcoal, 146  
 Charles' Law and its applications, 279-281  
 Chemical activity, according to electron theory, 58  
 Chemical industry, advances in, 275 *f.*  
 Chemical properties, 2  
 Chemical union, according to electron theory, 59  
 Chemistry, definition of, 1; reasons for studying, 1; progress of, 273-278  
 Chile saltpeter, 206-207  
 Chloride of lime, 215-216  
 Chlorides, 94; test for, 94; of iron, 229  
 Chlorination, of water, 36  
 Chlorine, occurrence of, 87; preparation of, 87-88; properties of 88-90; bleaching action of, 90; uses of 90-91  
 Chloroform, 178-179  
 Chlorophyll, 109, 155  
 Chrome steel, 227  
 Chromium, 247-248  
 Citric acid, 182  
 Clay, 9, 234, 237-238  
 Coagulation, in water purification, 36, 237; in preparing colloids, 268  
 Coagulum, 237  
 Coal, 147; anthracite, 147; bituminous, 147-148; gas, 166; tar, 166  
 Coal stove; reactions in, 158  
 Coal tar dyes, 264  
 Cobalt glass, 201  
 Coke, 146  
 Collodion, 185  
 Colloids, importance of, 267; varieties of, 267-268; preparation of, 268; properties of, 268; protective, 270; applications of, 270-271  
 Combination, direct, 52  
 Combustion, 14; spontaneous, 14  
 Compounds, 4; comparison with mixtures, 5  
 Concrete, 238  
 Congo red, 264  
 Conservation of Energy, Law of, 6  
 Conservation of Matter, Law of, 5, 46  
 Contact process of making sulfuric acid, 136-138  
 Cooling, to preserve foods, 257  
 Copper, 240 *f.*; metallurgy of, 240; electrolytic refining of, 240-241; properties of, 241; uses of, 242; compounds of, 242  
 Copper oxide, reduction of, 25  
 Corrosion, 196  
 Corundum, 236  
 Cotton, 261-262; mercerized, 262  
 Cottrell Process, 269  
 Cracking of oils, 166  
 Cream of tartar, 182  
 Crucible steel, 225  
 Crude oil, 164  
 Cryolite, 234 *f.*  
 Crystallization, water of, 40; fractional, 202  
 Crystalloids, 267, 269  
 Crystals, 39-40  
 Cyanamide process, 115, 124  
 Cycle, nitrogen, 108; carbon dioxide, 109  
 Dalton, 45  
 Decay, 15, 16, 257-258  
 Decomposition, 52  
 Decrepitation, 40, 202  
 Definite Proportions, Law of, 4, 35, 46  
 Dehydration by sulfuric acid, 139  
 Deliquescence, 40-41  
 Denaturing of alcohol, 180  
 Destructive distillation, 115, 146, 166  
 Deuterium, 275; oxide, 275  
 Deuteron, 274, 275  
 Developing, in photography, 249  
 Dewar flasks, 111  
 Dextrose, 186  
 Dialysis, 269  
 Diamond, 145  
 Diastase, 186  
 Diet, 257  
 Digestion, 258  
 Direct combination, 52  
 Direct dyeing, 264  
 Dissociation, illustrations of, 66-68  
 Distillation, of water, 36, 37; destructive, 146, 166-167; fractional, 164  
 Double replacement, 53  
 Double salts, 81  
 Drawing the temper, 227  
 "Dry ice," 155, 156  
 Drying, to preserve foods, 257  
 Dutch process of making white lead, 246  
 Dyeing, 237, 264-265  
 Dynamite, 122, 181, 183  
 Earth's crust, composition of, 9  
 Efflorescence, 40-41  
 Electric arc process of fixing nitrogen, 124  
 "Electric eye," 201  
 Electric furnace, 150; for making steel, 226  
 Electrolysis, of water, 12-13; ionic explanation of, 69; of copper chloride, 69; of sodium chloride, 70  
 Electrolyte, 66

# INDEX

ix

- Electromotive (electrochemical) series, 194, 342
- Electron Theory, 55-63
- Electroplating, 197, 242, 247
- Elements, 3; classification of, 3, 192; transmutation of, 61, 274; complete table of, 347
- Emery, 236
- Emulsion, 37, 270
- Energy, definition of, 5; forms of, 5; within the atom, 275
- "Enriching" (water gas), 166
- Enzymes, 180; in digestion, 258
- Epsom salts, 245
- Equations, 51; how to write, 51; aids in balancing, 51; problems involving, 296 *f.*
- Equilibrium, 73
- Esterification, 182
- Esters, 182-184
- Etching of glass, 99-100
- Ethereal salts, *see* "Esters"
- Ethylene series, 177
- Eudiometer, 33
- Explosion, 165
- Explosive mixture, 171-172
- Exposing, in photography, 249
- Fats, 183; in food, 253; test for, 255
- Fermentation, 180
- Ferric compounds, oxide, 228; hydroxide, 228; chloride, 229; sulfate, 229
- Ferrous compounds, oxide, 228; hydroxide, 228; chloride, 229; sulfate, 229
- Fertilizers, 122, 123-124, 207, 215
- Fibers, 261 *f.*
- Filtration, 35
- Fire, conditions necessary for, 15; control of, 15
- Fire extinguishing, with carbon dioxide, 156; with foamite, 157
- Fixation of nitrogen, 123-124
- Flame, 169; color of, 170; Bunsen burner, 170-171; candle, 171
- Flame tests, 341
- Floating soap, 184
- Fluorine, occurrence of, 99; preparation of, 99; properties of, 99; a typical non-metal, 191
- Flux, 221
- Foamite, 157
- Food, importance of, 253; nutrients in, 253; vitamins in, 254-255; tests, 254-255; energy value of, 255-256; preservation of, 257; digestion of, 258
- Formaldehyde, 181
- Formic acid, 182
- Formulas, 48; how to write, 49-50; structural, 174 *f.*; and percentage composition, 291 *f.*
- Frasch method of extracting sulfur, 127-128
- Freezing points, abnormal, 72
- Fructose, 187
- Fuels, definition, 163; desirable properties of, 163; classes of, 163; liquid, 163-166; gaseous, 166-169
- Galena, 246
- Galvanized iron, 197
- Gamma rays, 250
- Gasoline 164-166; increasing production of, 166
- Gay-Lussac's Law, 285
- Glass, 99-100, 278
- Glazing, 238
- Glucose, 186; test for, 254
- Glycerin, 180-181
- Gold, 249-250
- Gram-molecular volume, 288
- Gram-molecular weight, 48, 288
- Grape sugar, 186; test for, 254
- Graphite, 145-146
- Gravimetric analysis, 35
- Guncotton, 122
- Gun metal, 2-2, 343
- Gypsum, 213-214
- Haber Process, 115, 124
- Halite, 201
- Hall Process, 234-235
- Halogens, importance of, 85; properties of, 85; uses of, 86; replacement power of, 194-195
- Hard soap, 183-184
- Hard water, 214-215
- Heavy hydrogen, 275
- Heavy water, 275
- Helium, electronic structure of, 57; in the air, 103, 110-111
- Hematite, 219, 224, 226
- Hoffman apparatus for electrolysis of water, 12, 13
- Homologous series, 176-177
- Humidity, 110
- Hydrates, 40
- Hydration, water of, 40
- Hydrocarbons, 176 *f.*
- Hydrochloric acid, occurrence of, 91; preparation of, 91-92; properties of, 93; uses of, 94
- Hydrogen, importance of, 20; occurrence of, 20; uses of 21, 26; preparation of, 22-23; properties of, 23-24; occlusion of, 24; as a reducing agent, 25; test for, 26; electronic structure of, 56; heavy, 275; number of atoms in molecule of, 286
- Hydrogenation, of oils, 26-27; of coal, 166
- Hydrogen bromide, 96
- Hydrogen chloride, properties of, 93; fountain, 93
- Hydrogen fluoride, 99; used in etching glass, 99-100
- Hydrogen iodide, 98-99
- Hydrogen peroxide, 41
- Hydrogen sulfide, occurrence of, 132; preparation of, 132; properties of, 132-133; uses of, 133
- Hydrolysis, 71-72; of starch, 186
- Hydroxyl group, 79
- Hygroscopic substance, 263
- Hypo, 38, 249
- Hypochlorous acid, 90
- Ice, artificial 117-118; "dry," 155, 156
- Indicators, 78
- Ingot, 224
- Ink, 230, 276
- Intestinal juice, 258
- Inversion of sugar, 187, 254
- Invertase, 187
- Invert sugar, 187
- Iodides, test for, 106
- Iodine, occurrence of, 96; preparation of, 97-98; properties of, 98; uses of, 98
- Iodoform, 179
- Ion, 66; compared with atom, 68
- Ionization Theory, 65-73
- Iridium, 277
- Iron, importance of, 219; occurrence of, 219; metallurgy of, 219 *f.*; cast, 222; wrought, 222; compounds of, 227 *f.*
- Isotopes, 60-61

- Kaolin, 237  
 Kerosene, 164  
 Kindling temperature, 14  
 Krypton, 103, 110-111
- Lactic acid, 182  
 "Lake," in dyeing, 237, 265  
 Lampblack, 147  
 "Laughing gas," 119  
 Laundry soap, 184  
 Lavoisier, 5  
 Lead, 246; white, 246  
 Leguminous plants, 108, 123  
 Levulose, 187  
 Lignite, 148  
 Lime kiln, 212  
 Limestone, 210, 211; caves, 211  
 Limewater, 213; used in tests for carbon dioxide and carbonates, 157  
 Limonite, 219  
 Linen, 262  
 Liquid air, 13, 111  
 Lithium, electronic structure of, 57; a typical metal, 191; flame test, 201; uses, 277  
 Litmus, 78 *f.*  
 Lodestone, 228  
 Lunar caustic (silver nitrate), 248  
 Lye, 204
- Magnesium, 245  
 Magnetite, 228  
 Manganese, 247  
 Manganese dioxide, 12  
 Manganese steel, 227  
 Marble, 210, 211  
 Matter, 1; states of, 1-2; properties of, 2; kinds of, 3, 4-5  
 Mendelejeff, 62  
 Mercerized cotton, 262  
 Mercuric chloride, 245  
 Mercuric oxide, decomposition of, 13  
 Mercurous chloride, 245  
 Mercury, 245  
 Metallurgy, 192-193; of iron, 219 *f.*; of aluminum, 234-235; of copper, 240-241  
 Metals, 3; as borrowers of electrons, 58; comparison with non-metals, 191-192; occurrence of, 192; extraction of, 192-193; relative activity of, 193; replacement of, 193-194; weathering of, 196; new uses for, 276-277  
 Metathesis, 53  
 Methane, 176, 177; series, 176  
 Mica, 234  
 Milk, a perfect food, 257; as an emulsion, 270  
 Milk of lime, 213  
 Mineral matter in food, 254; test for, 255  
 Miscible liquids, 37  
 Mixtures, 4; comparison with compounds, 5  
 Molecular weight, 48; determination of, 287; of a gas, 287-289  
 Molecule, 47-48  
 Molybdenum steel, 227  
 Mordant, 237, 265  
 Mortar, 213, 271  
 Moseley, 56  
 Multiple Proportions, Law of, 42, 46  
 Muriatic acid, 93  
 Mustard gas, 91
- Naphthalene, 178  
 Nascent state, 41  
 Negative, of a photograph, 249  
 Nelson cell, 203
- Neon, electronic structure of, 57; in the air, 110-111  
 Neoprene, 276  
 Neutralization, 70-71  
 Neutrino, 274  
 Neutron, 55, 273  
 Nichrome, 196  
 Nickel, 247  
 Nickel steel, 227  
 Nitrates, 122; test for, 122  
 Nitric acid, preparation of, 120; properties of, 121-122; uses of, 122  
 Nitric oxide, 119  
 Nitrocellulose, 185  
 Nitrogen, occurrence of, 106; preparation of, 106-107; properties of, 107; uses of, 107; cycle, 108; compounds of, 114 *f.*; fixation, 123-124  
 Nitrogen pentoxide, 120  
 Nitrogen peroxide, 120  
 Nitrogen trioxide, 120  
 Nitroglycerine, 122, 131, 183  
 Nitrous oxide, 119  
 Nobel Prize, 326  
 Non-electrolyte, 66  
 Non-metals, 3; as lenders of electrons, 58; comparison with metals, 191-192  
 Non-miscible liquids, 37; in formation of emulsions, 270  
 Non-polar compound, 59  
 Normal salt, 81  
 Nucleus, atomic, 55; structure of, 273-274  
 Nutrients, 253 *f.*  
 Nutrition, 253 *f.*
- Oils, 183  
 Olein, 183  
 Open-hearth process, 224-225; uses of steel produced by, 225  
 Organic chemistry, 174-187  
 Osmium, 277  
 Oxalic acid, 182  
 Oxidation, slow, 14; rapid, 14; control of, 15; according to Electron Theory, 59; and reduction, 229; of ferrous chloride, 229-230  
 Oxides, formation of, 13; definition, 76; methods of preparing, 77; naming, 77  
 Oxy-acetylene torch, 16, 169  
 Oxygen, importance of, 9; occurrence of, 9, 11; uses of, 10, 15-16; percentage of in the earth's crust, 11; preparation of, 11-12; properties of, 13; percentage of in the air, 104-105  
 Oxy-hydrogen torch, 16, 26  
 Ozone, 16-17; in purification of water, 36
- Paint, composition of, 247  
 Pancreatic juice, 258  
 Paper, 174, 185  
 Paraffin series, *see* "Methane series"  
 Pasteurizing, 257  
 Peat, 148  
 Pepsin, 258  
 Peptization, 268  
 Percentage composition, from formula, 291-292; formula from, 292-294  
 Periodic law and table, 62-63  
 Permanent hard water, 214  
 Permutit, 215  
 Petroleum, refining of, 164; products, 164  
 Phenol, 181  
 Phosgene, 91, 159  
 Photo-electric cell, 201  
 Photography, 248-249  
 Photon, 274



- Photosynthesis, 155, 185  
 Physical properties, 2  
 Pig iron, 222  
 Pigment, in paints, 247  
 Plaster, 213, 271  
 Plaster of Paris, 213-214  
 Plastics, 276  
 Platinum, as a catalyst in contact process, 137; occurrence and uses, 250  
 Poison gases, used in warfare, 91, 159  
 Polar compound, 59  
 Polymerization, 276  
 Portland cement, 238  
 Positive, of a photograph, 249  
 Positron, 274  
 Potash, 207; caustic, 207  
 Potassium, metallic, 200-201; flame test, 201; compounds, 207  
 Potassium chlorate, in preparation of oxygen, 11-12  
 Preservation of foods, 257-258  
 Priestley, 11  
 Printing, in photography, 249  
 Producer gas, 168  
 Properties of matter, physical, 2; chemical, 2  
 Protective colloids, 270  
 Proteins, 253; test for, 255  
 Proton, 55  
 Protoplasm, 258  
 Ptyalin, 258  
 Puddling of iron, 223  
 Pyroxylin, 185  
  
 Quartz, 9, 249  
 Quicklime, 212  
 Quicksilver, *see* "Mercury"  
  
 Radioactivity, 61, 250-251  
 Radium, 61, 250-251  
 Rare gases of the air, 110-111  
 Rayon, 263  
 Reactions, types of, 52; reversible, 72; going to completion, 73  
 Reducing agent, 25  
 Reduction, by hydrogen, 25, 27; according to Electron Theory, 59; and oxidation, 229; of ferric chloride, 230  
 Reinforced concrete, 238  
 Refrigeration, with ammonia, 117-118; with sulfur dioxide, 136; with carbon dioxide, 156; new developments in, 278  
 Regenerative furnace, 225  
 Relative humidity, 110  
 Replacement, simple, 52; double, 53; of metals, 193-194; of non-metals, 194-195  
 Research, chemical, 273  
 Respiration, 16; use of oxygen in artificial, 16; and carbon dioxide cycle, 109  
 Roasting of ores, 193, 240  
 Rubber, vulcanization of, 127, 132; synthetic, 276  
 Rubies, 236  
 Russia iron, 196  
 Rusting, conditions necessary for, 15; control of, 15, 196-197  
 Rutherford, 325  
  
 Saltpeter, 207  
 Salts, definition, 80; properties of, 80; preparation of, 80-81; types of, 81; method of naming, 82  
 Saponification, 183  
 Sapphires, 236  
 Saturated compounds, 177  
  
 Scientific habit of mind, 1  
 Selenium, 269  
 Siderite, 219  
 Silicates, 234 *ff.*  
 Silicon, 9, 234  
 Silicon carbide, 150  
 Silicon steel, 227  
 Silk, 263  
 Silk substitute, 123, 186  
 Silver, 248; compounds of, 248-249  
 Simple replacement, 52  
 Slag, 222  
 Slaked lime, 213  
 Soap, 183-184; kinds of, 183; cleansing action of, 184  
 Soda ash, 205  
 Soda-lime process of preparing sodium hydroxide, 204  
 Soda water, 155  
 Sodium, electronic structure of, 57; metallic, 200-201; compounds, 201 *ff.*; flame test, 201; compounds of 201-207  
 Sodium bicarbonate, 206  
 Sodium carbonate, 204-205  
 Sodium chloride, dissociation of, 66-68; occurrence, 201; extraction, 201-202; purification, 202; properties, 202; uses, 202  
 Sodium hydroxide, 203-204  
 Sodium nitrate, 206-207  
 Solubility, effect of temperature on, 38; effect of pressure on, 38-39  
 Solute, 36  
 Solution, definition, 36; dilute, 37; concentrated, 37; unsaturated, 37; saturated, 38; supersaturated, 38; aids to, 39; compared with suspension, 39; conductivity of, 65-66  
 Solvay Process, 204-205  
 Solvent, 36  
 Specific gravity, of a liquid, 31; of a gas, 289  
 Spiegeleisen, 224  
 Spontaneous combustion, 14  
 Stalactites, 211  
 Stalagmites, 211  
 Standard conditions of temperature and pressure, 279  
 Starch, 185; test for, 254  
 Stearin, 183  
 Steel, 223 *ff.*; Bessemer, 223; open-hearth, 224-225; crucible, 225; electric furnace, 226; tempering of, 226-227; alloys, 227; stainless, 227, 248  
 Sterilizing, 257  
 Sterling silver, 248  
 Strontium, flame test, 201  
 Structural formulas, 174 *ff.*  
 Styptic, 237  
 Sublimation, 40, 97  
 Substitution, 52; products, 178-179  
 Sucrose, 186-187; test for, 254  
 Sugars, 186-187; as a food, 253 *ff.*  
 Sulfates, test for, 140; of iron, 229  
 Sulfides, test for, 133  
 Sulfur, importance of, 127; occurrence of, 127; extraction of, 127-128; allotropic forms of, 128-131; uses of, 130; properties of, 131-132; compared with oxygen, 131  
 Sulfur dioxide, occurrence of, 133; preparation of, 134-135; properties of, 135; uses of, 136  
 Sulfuric acid, preparation of by contact process, 136-138; fuming, 136, 138; properties of, 138-139; uses of, 139  
 Sulfurous acid, 135  
 Sulfurous anhydride, 135  
 Sulfur trioxide, 136

- Superphosphate of lime, 215  
 Supersaturated solution, 38  
 Suspension, 37; compared with a solution, 39  
 Symbols, 3, 47  
 Synthesis, 32, 52; of water, 32-34  
 Synthetic substances, 276 *f.*
- Tantalum, 277  
 Tartaric acid, 182  
 Tear gas, 96  
 Tempering steel, 226-227  
 Temporary hard water, 214  
 Ternary acid, 81-82  
 Ternary salt, 82  
 Thermit, 236  
 Thyroxin, 98  
 Tin, 246  
 Tinning, 197  
 Toluol, 178  
 Toning, in photography, 249  
 Trinitrotoluene (T.N.T.) 122, 183  
 Tungsten steel, 227  
 Turpentine, 178  
 Tyndall effect, 268  
 Type metal, 247
- Ultramicroscope, 268-269  
 Ultra-violet rays, in water purification, 36  
 Unsaturated compounds, 177  
 Unsaturated solutions 37  
 Uranium, 250-251, 275  
 Urey, 275
- Valence, definition of, 48; of radicals, 49;  
 positive and negative, 49; table of 50;  
 according to Electron Theory, 57-58  
 Vanadium steel, 227
- Vat dyeing, 265  
 Vinegar, 182  
 Viscose silk, 263  
 Vitamins, 254, 255  
 Vitriol, oil of, 138  
 Volatility, 73  
 Volume of gases, correction of, 279 *f.*  
 Volumetric analysis, 35  
 Voice cell, 203  
 Vulcanization of rubber, 127, 132
- Water, importance of, 29, 30; occurrence of,  
 29; formation of 29, 31; uses of, 30; prop-  
 erties of, 31-32; solvent action of, 31; con-  
 ductivity of, 31; as a standard for specific  
 gravity, 31; of crystallization, 32; analyses  
 of, by volume, 35; synthesis of, by volume,  
 33; synthesis of, by weight, 34; impurities  
 in, 35; purification of, 35; hard, 214-215; in  
 food, 254; heavy, 275
- Water gas, 167  
 Weathering, of metals, 196  
 Welding, by monatomic hydrogen, 26; by  
 thermit, 236  
 White lead, 246  
 Whitewash, 213  
 Wohler, 174  
 Wood's metal, 247  
 Wool, 262-263  
 Wrought iron, 222
- Xenon, 103, 110-111  
 X-ray spectra, 56
- Yeast, 155, 180, 187
- Zinc, 193 *f.*, 244-245; compounds of, 244  
 Zirconium, 277  
 Zymase, 180, 187

## ADDENDUM

(This addendum to the Index covers Chapter 29, "Chemistry in Warfare," pages 302-321.)

- Adamsite, a sneezing gas, 319  
 Air raid, precautions during an, 315-318  
 Amatol, a high explosive, 305  
 Bomb, aerial, 304; all-thermit, 313; ef-  
 fects of aerial, 312-313; electron, 310;  
 incendiary, 310-314; types of, 309-310  
 Booster, an explosive material, 306  
 Bromaceton, a tear gas, 319  
 Cartridges, construction of, 305-306  
 Charge, bursting, defined, 302  
 Chloracetophenone, a tear gas, 319  
 Chlorine, as a lung irritant, 319  
 Chlorpicrin, a military gas, 319  
 Combustible ingredients, of gunpowder, 302  
 Defense, air, measures to take, 315  
 Detonating cap, purpose of, 306  
 Dunnite, a U. S. A. high explosive, 305  
 Dynamite, how made, 303  
 Explosive, explained, 302; *high*, term de-  
 fined, 304; source of destructive force  
 of, 302; -train, 307  
 Fuse, time, for bombs, 310; construction  
 of, 307-308  
 Gas, in modern warfare, 319; in World  
 War II, 320-321; kinds of, 319-320  
 Grenade, hand, 304, 308-309  
 Gun cotton (See Nitrocellulose.)  
 Gunpowder, invention of, 302; composi-  
 tion of, 302
- Incendiary bombs, how to combat, 317-  
 318; types of, 313  
 Lacrimators (tear gases), 319  
 Lewisite, a poison gas, 320; its action,  
 320  
 Mask, gas, construction of, 321; military,  
 321  
 Mine, explosive power of, 304  
 Mustard gas, 320  
 Nitrocellulose, 303-304  
 Nitrogen compounds, 302  
 Nitroglycerine, formula for, 303  
 Percussion cap (See Detonating cap.)  
 Phosgene, a military gas, 319  
 Phosphorus, use of, in bombs, 313  
 Picric acid, composition of, 305  
 Powder, gun-, 302; smokeless, 303  
 Propellent, defined, 302  
 Shell, 304, 306; anti-aircraft, 308;  
 chemical, 309; kinds of, 308-309;  
 structure of, 306  
 Shrapnel shells, construction of, 308  
 Sternutators (sneezing gases), 319  
 Tear gas (See Lacrimators.)  
 T. N. T., 302, 305  
 Torpedo, an implement of warfare, 304  
 Trinitrotoluene (See T. N. T.)  
 Vesicants (blistering gases), 320; pro-  
 tection against, 321

A TABLE OF THE IMPORTANT ELEMENTS

ELEMENT	SYMBOL	APPROXIMATE ATOMIC WEIGHT	ATOMIC NUMBER	VALENCE
Aluminum.....	Al	27	13	3
Antimony.....	Sb	122	51	3, 5
Argon.....	A	40	18	.....
Arsenic.....	As	75	33	3, 5
Barium.....	Ba	137	56	2
Bismuth.....	Bi	209	83	3, 5
Boron.....	B	11	5	3
Bromine.....	Br	80	35	1
Cadmium.....	Cd	112	48	2
Calcium.....	Ca	40	20	2
Carbon.....	C	12	6	4
Chlorine.....	Cl	35.5	17	1
Chromium.....	Cr	52	24	2, 3, 6
Cobalt.....	Co	59	27	2
Copper.....	Cu	63.6	29	1, 2
Fluorine.....	F	19	9	1
Gold.....	Au	197	79	1, 3
Helium.....	He	4	2	.....
Hydrogen.....	H	1	1	1
Iodine.....	I	127	53	1
Iron.....	Fe	56	26	2, 3
Lead.....	Pb	207	82	2, 4
Lithium.....	Li	7	3	1
Magnesium.....	Mg	24	12	2
Manganese.....	Mn	55	25	2, 4
Mercury.....	Hg	200	80	1, 2
Neon.....	Ne	20	10	.....
Nickel.....	Ni	58.7	28	2
Nitrogen.....	N	14	7	3, 5
Oxygen.....	O	16	8	2
Phosphorus.....	P	31	15	3, 5
Platinum.....	Pt	195	78	4
Potassium.....	K	39	19	1
Radium.....	Ra	226	88	2
Silicon.....	Si	28	14	4
Silver.....	Ag	108	47	1
Sodium.....	Na	23	11	1
Strontium.....	Sr	87.6	38	2
Sulphur.....	S	32	16	2, 4, 6
Tin.....	Sn	119	50	2, 4
Zinc.....	Zn	65	30	2

