# Calorimetry

# Introduction

The word *calorimetry* is derived from Latin, meaning ‘heat measure’. Calorimetry is a special case of conservation of energy, where:

Qgain = -Qlost

The letters ‘Q’ or ‘q’ are standard symbols for heat. The equation assumes that there are no other energies being transferred between the two substances except heat, and that none of the heat is dissipated into the surroundings.

A calorimeter is used to measure the amount of heat gained or lost by a substance in a process, such as a chemical or physical change. A simple but effective calorimeter for limiting heat lost to the surroundings is a Styrofoam™ cup. The cup can include a Styrofoam™ lid to further limit heat loss, but for the sake of this experiment, that should not be necessary.

There are two separate equations employed for measuring heat gain or loss. If the temperature of the substance changes, then

Q = m c Δ T = m c (Tf – Ti).

In this equation: Q is the heat gained or lost; m is the mass of the substance; c is the *specific heat* of the substance; and ΔT is the change in temperature, Tfinal – Tinitial. Specific heat is a property of a pure substance, typically measured in J/g-C. Note that if a substance loses heat, the final temperature is less than the initial temperature and the value of q is negative. The opposite is true if the substance is gaining heat.

In the subject of thermodynamics, heat is the *transfer* of energy between two systems because of a temperature difference. A system does not *store* heat; instead, it possesses thermal energy. The thermal energy of a system may change because of heat flow, but it may also change because of work, which is a second method for transferring energy between systems.

Sometimes heat flows into or out of a system but the temperature does not change. An example of this is a melting ice cube: heat is flowing into the ice cube and the solid is becoming a liquid, but the temperature remains constant during the change. Therefore, we have a second equation for calculating heat flow:

Q = m ΔH

where Q is the heat flow, m is the mass of the substance, and ΔH is called the *enthalpy* of the process. A physical change such as melting or boiling has an enthalpy; chemical reactions also have enthalpies since a chemical change can occur at a constant temperature.

In Part I of this activity, the heat lost by a warm metal cylinder will be equal to the heat gained by a cool cup of water. Energy will flow until both substances are at the same temperature.

In Part II, an ice cube will melt by gaining heat from the surrounding water. Heat will also be needed to raise the newly formed ice water from 0 Celsius to the final temperature of the water.

In Part III, combining acid and base solutions will result in the formation of water molecules:

1 HCl(aq) + 1 NaOH(aq) 1 H2O(l) + 1 NaCl(aq)

or as a net ionic equation:

1 H+(aq) + 1 OH-(aq) 1 H2O(l)

Water molecules have less stored potential energy than the ions, so this reaction releases heat and the solution gets warmer. The salt in the solution, NaCl, also absorbs thermal energy as the temperature of the solution rises, but we will ignore this small component of the total change in energy.

# Materials and Hazards

1.0 M Hydrochloric Acid (HCl) – Corrosive, Poisonous, Toxic

1.0 M Sodium Hydroxide (NaOH) – Corrosive, Mutagenic, Toxic

Baking Soda

Ice

Styrofoam Cup

LabQuest with a temperature Probe

Metal Cylinder

# Procedure

Part I: Calculating Specific Heat

1. Obtain a metal cylinder and a ‘coffee cup’ calorimeter. Weigh the calorimeter, fill it about half-full of water, then weigh it again. (The amount of water in the cup should be enough to submerge the cylinder when it is placed in the cup.) The difference in weight is the mass of the water in the cup.
2. Use a thermometer to find the initial temperature to the closest 0.1 Celsius.
3. Place the cylinder in a 250 mL beaker and add just enough water to submerge it. Heat the beaker to boiling on a hot plate.
4. Use crucible tongs to *quickly* transfer the cylinder into the calorimeter. Stir carefully with the thermometer and record the highest temperature reached by the system to the closest 0.1 Celsius.
5. Complete the calculations for this Part.

Part II: An Endothermic Physical Change

1. Fill the calorimeter about half-full of water and find its mass as in Part I. Record the initial temperature of the water to the closest 0.1 Celsius.
2. Add an ice cube to the calorimeter. Stir the ice cube and water carefully with the thermometer until the ice cube melts and record the lowest temperature reached to the closest 0.1 Celsius.
3. Weigh the calorimeter and water a second time.
4. Complete the calculations for this Part.

Part III: An Exothermic Chemical Change

1. Measure 50.0 ml of a 1.0 M hydrochloric acid solution with a graduated cylinder and pour it into the calorimeter. Record its initial temperature to the closest 0.1 Celsius.
2. Measure 50.0 mL of a 1.0 M sodium hydroxide solution in a graduated cylinder and measure its temperature. It should be close to the first temperature. (Why?) *If not, use the average of the two readings to represent the system's initial temperature.*
3. Pour the second solution into the calorimeter. Stir the mixture gently with the thermometer and record the highest temperature reached to the closest 0.1 Celsius.
4. Measure the mass of the solution after the final temperature is reached. The density of water and these solutions are close to 1.00 g/mL, so the mass of the solution should be close to 100.0 grams.
5. Complete the calculations for this Part.

# Data and Results

Part I

1. Subtract the mass of the cup from the mass of the cup with water to find the mass of water in the cup.
2. Subtract the final temperature of the water from the initial temperature to find the change in temperature of the water. This will be a positive number.
3. Calculate the heat gained by the water from the equation, Q = mcΔT, using the specific heat of water, c = 4.18 J/g-C.
4. Subtract the final temperature of the metal cylinder (which will be the same as the final temperature of water) from the initial temperature (that of boiling water – 100 C) to find the change in temperature of the metal. This change in temperature will be negative. Since the first equation includes a negative sign for the heat lost, we will use the absolute value of the temperature change in the calculations.
5. Record the heat lost by the metal. Based on the assumptions of calorimetry, it will equal the amount of heat gained by the water.
6. Divide the heat lost by the water by the mass of metal and its temperature change to determine the specific heat of the metal.
7. The specific heat of the metal will be given to you. Calculate the percent error of the measurement and record the result on the Report Sheet.

Part II

1. Subtract the mass of the cup (from Part I) from the first reading of the mass of the cup with water to find the mass of water in the cup.
2. Subtract the second mass reading from the first mass reading to find the mass of melted ice.
3. Subtract the final temperature of the water from the initial temperature to find the change in temperature of the water. This will be a negative number because the water was losing heat.
4. Calculate the heat lost by the water from the equation, Q = mcΔT, using the specific heat of water, c = 4.18 J/g-C and the absolute value of the change in temperature.
5. The heat was used for two purposes: to convert the solid ice into liquid water at 0 Celsius; and to raise the temperature of the newly formed ice water from 0 C to the final temperature of the water. Subtract the final temperature of the newly formed ice water (which is the same as the final temperature of the water) from its initial temperature (0 C) to find the change in temperature of the ice water. This will be a positive number.
6. Calculate the heat gained by the ice water from the equation, Q = mcΔT, where m is the mass of the ice water c is the specific heat of water, 4.18 J/g-C.
7. Subtract the heat used to raise the temperature of the ice water from the total heat lost by the water in the cup. This is the amount of heat used to melt the ice cube.
8. Divide the heat used to melt the ice cube by the mass of the ice cube. This is the enthalpy of the phase change.
9. The enthalpy for the melting of ice will be given to you. Calculate the percent error of your measurement and record the result on the Report Sheet.

Part III

1. Subtract the mass of the cup (from Part I) from the mass of the cup and saltwater solution to find the mass of solution in the cup.
2. Subtract the final temperature of the water from the initial temperature to find the change in temperature of the water. This will be a positive number because heat was released I the chemical reaction.
3. Calculate the heat gained by the water from the equation, Q = mcΔT. Assume the specific heat of the dilute salt solution is the same as water, c = 4.18 J/g-C.
4. Use the definition of molarity and the molarity and volume of acid solution to calculate the moles of acid: mol = M \* vol.
5. Use the definition of molarity and the molarity and volume of base solution to calculate the moles of base: mol = M \* vol.
6. Consider the net ionic reaction:

1 H+(aq) + 1 OH-(aq) 1 H2O(l)

How many moles of water were formed in this reaction?

1. Divide the heat gained by the solution by the moles of water formed. This is the enthalpy of the reaction, in J/mol. Convert this number to kJ/mol.
2. The enthalpy of an acid-base neutralization is between 40 – 60 kJ/mol. How did your calculation compare to this range?

# Clean Up

Pour the solution form Part III into a waste beaker with excess baking soda in it. Rinse the Styrofoam (TM) cup and return it and the metal cylinder to their proper locations.