Hess’s Law

# Introduction

Enthalpy is a *state function*. It is important to understand what we mean by this term, and why it is so important in practical applications of chemistry.

Suppose a person wanted to climb to the top of a mountain and there were two paths. One route was an almost straight path and difficult to climb; the other path weaved around and back again so that the trail was never too steep, but it took three times the distance to get to the top. The number of steps taken depends on the trail, but the vertical height reached is the same.

Heat and work are like the choice of a trail: there are many different combinations of the two can be used to complete a thermal energy transfer. But enthalpy as a state function is like the height of the mountain: it does not depend on how you get there; the numerical value can be calculated from the initial and final states only.

Why is this important? Suppose the enthalpy of a chemical reaction has never been calculated. In fact, suppose making the reaction happen would be extremely hard to do. It will still be possible to determine the enthalpy if the enthalpies of a series of reactions were known that added up to the reaction of interest.

Hess’s Law states that enthalpies are additive. For example, suppose the enthalpy for this reaction was inknown:

2C → D ΔH =?

If the enthalpies were known:

C + B → A ΔH = +30 kJ

and

2B + D → 2A ΔH = -20 kJ

then doubling the coefficients of the first equation, including the enthalpy:

2C + 2B → 2A ΔH = +60 kJ

and reversing the second equation, which changes the ign of the enthalpy:

2A → 2B + D ΔH = +20 kJ

Results in the original equation. Therefore, the enthalpy is the sum of the enthalpies: 60 + 20 = 80 kJ.

In this experiment, the enthalpy for the hydration of calcium chloride:

CaCl2(s) + 2 H2O(l) → CaCl2\*2 H2O(s)

Is calculated by dissolving anhydrous calcium chloride and calcium chloride dihydrate separately in water.

CaCl2(s) → CaCl2(l)

CaCl2\*2 H2O(s) → CaCl2(l) + 2 H2O(l)

The enthalpies of solution for these two reactions will allow us to determine the heat gained or lost in the first reaction.

# Materials and Methods

Styrofoam™ cup calorimeter

LabQuest 2 with a temperature probe

Anhydrous calcium chloride

Calcium chloride dihydrate

# Procedure

1. Find the mass of the calorimeter and record it on the Report Sheet.
2. Fill the cup about half-full of distilled water (around 100 g), weigh it, and record the mass on the Report Sheet.
3. Record the initial temperature of the distilled water.
4. Measure around 6.5 g of anhydrous calcium chloride in a weighing boat and record the mass to the closest 0.01 g on the Report Sheet.
5. Add the salt to the water. Stir gently with the temperature probe to dissolve the solid. Record the lowest (or highest) temperature reached, to the closest 0.1 C, on the Report Sheet.
6. Discard the calcium chloride solution. Repeat the procedure with around 8.5 g of calcium chloride dihydrate, also measured to the closest 0.01 g.

# Data and Results

1. Determine the change in temperature for the dissolution of anhydrous calcium chloride.
2. Calculate the heat gained or lost by the anhydrous calcium chloride using the equation q = mc ΔT. Let m be the mass of distilled water and c = 4.18 J/g-C. [It is true that energy will be gained or lost when the CaCl2 ionic salt dissolves into ions, but we will assume this energy change is small compared to the thermal energy change of the water.]
3. Calculate the moles of calcium chloride in sample.
4. Calculate the enthalpy of solution in units of kJ/mol. Include the sign: negative if the temperature fell, and positive if the temperature rose.
5. Repeat these steps for calcium chloride dihydrate.
6. Given:

CaCl2(s) → CaCl2(l) ΔH1

CaCl2\*2 H2O(s) → CaCl2(l) + 2 H2O(l) ΔH2

determine on how the two enthalpies should be combined to calculate the enthalpy of interest:

CaCl2(s) + 2 H2O(l) → CaCl2\*2 H2O(s) ΔH =?

Is this enthalpy positive or negative?