Nature has given us two ears, two eyes, and but one tongue, to the end that we should hear and see more than we speak. --Socrates

The Determination of the Heat of Formation of Magnesium Oxide

Chemical reactions proceed with the evolution or absorption of heat. This heat flow represents differences in chemical energy associated with the arrangement of atoms into molecules or ions. It includes the electrical potential energy arising from attractions between electrons and nuclei in atoms or ions for each other. Thus when a piece of magnesium ribbon is burned in air to yield magnesium oxide, a large quantity of heat is released to the surroundings:

$$Mg(s) + \frac{1}{2} O_2(g) \rightarrow MgO(s) + \text{heat}$$

Because heat is *evolved* in this reaction, we know that the sum of the chemical energies of the amounts of elemental magnesium and oxygen combining is greater than the chemical energy of the amount of magnesium oxide formed.

When the reaction takes place at constant pressure, the heat flow, q_p, is called the change in ENTHALPY of the system and for MOLAR quantities is given the symbol ΔH. The absolute value of the enthalpy, **H**, for a substance is indeterminate, but ΔH for a particular process is directly measurable, at least in principle, and is a quantity of fundamental interest in thermochemistry and thermodynamics.

When one mole of a substance is formed from its constituent elements in their standard states at standard conditions, ΔH for the reaction is called the *standard heat of formation* and given the symbol ΔH^o_f . In this experiment you will try to find ΔH^o_f for magnesium oxide, i.e., for the reaction:

$$Mg(s) + \frac{1}{2} O_2(g) \rightarrow MgO(s)$$

You are faced with a practical difficulty, however. To cause magnesium and oxygen to react they must be heated. The reaction then proceeds in an uncontrolled and highly exothermic fashion. The only device that can deal with this reaction directly is a bomb calorimeter [we don't own one].

However, so long as the system starts at and returns to standard conditions, the path it follows is not important. Any series of reactions that begins with Mg and O_2 and ends with MgO will involve the same overall change in enthalpy.

Adapted from "An Introduction to Small Scale Thermochemistry", *Heat of Reaction: Mg-HCl-H₂O System*, Walter Rohr, Flinn Scientific

Thus you can use Hess's Law and a number of reactions which are more easily dealt with to find the information about the reaction of magnesium and oxygen.

Consider the following reactions:

$$Mg(s) + 2\, HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad (1)$$

$$MgO(s) + 2\, HCl(aq) \rightarrow MgCl_2(aq) + H_2O(\ell) \qquad (2)$$

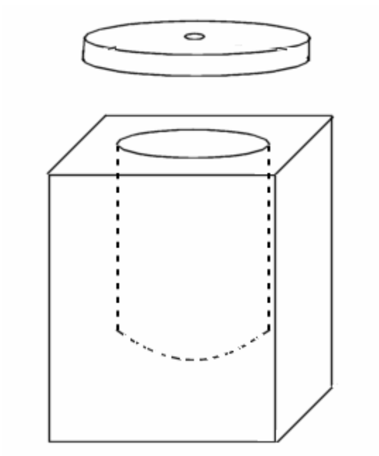
$$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(\ell) \qquad (3)$$

For the first two reactions, the heats of reaction are easily found in a simple calorimeter. For the third reaction, the heat of formation of liquid water, the enthalpy change is well known (-286 kJ/mol). By reversing equation (2) and adding all three we obtain:

$$Mg(s) + \frac{1}{2} O_2(g) \rightarrow MgO(s) \qquad (4)$$

And the *enthalpy change* for this reaction (ΔH^o_f), the heat of formation for magnesium oxide, is found by a similar reversal and addition of the enthalpies of the three reactions.

The constant-pressure calorimeter used in this experiment consists of a specially treated block of expanded polystyrene foam (commonly called "styrofoam", but not the same stuff as the rough and porous material often used in crafts). The block has a hole drilled in it with a practical volume of about 15 mL and is coated to help prevent it from absorbing liquid. Although the calorimeter is fairly well insulated some heat *will* escape or enter through hole in the lid provided for the thermometer, and some heat will be absorbed into or from the calorimeter contents by the calorimeter itself. For the purposes of this experiment we will ignore the former random error. But the systematic error introduced by ignoring the calorimeter itself can be reduced by using the calorimeter constant in heat calculations. The average calorimeter constant for this set-up has been determined to be 5.6 J/°C.



Preparing to experiment

You will be provided with the following materials:

1. 3.0 cm of Mg ribbon
   [the mass in grams per metre will be given in the lab--be sure to record it!]
2. 2.0 M HCl (use about 15 mL each time) [see **Technique** section]
3. a calorimeter, lid and stirring bar
4. a magnetic stirrer
5. MgO (use about 0.5 g) [see **Technique** section]
6. a thermometer probe

Design an experiment to determine the heat of reaction when a measured amount of Mg reacts with HCl (the amounts suggested make Mg the limiting reagent).

Design an experiment to determine the heat of reaction when a measured amount of MgO reacts with HCl (the amounts suggested make MgO the limiting reagent).

BE SURE TO BRING YOUR TI-83/84 CALCULATOR TO CLASS FOR THIS EXPERIMENT. YOU WILL ALSO NEED A COPY OF THE HCHEM.83G FILES IN YOUR CALCULATOR MEMORY.

Pre-lab take-home quiz

Answer these questions on a separate sheet of paper to be turned in on the day you do this experiment.

1. Most Mg ribbon is uniform enough that the mass of a small measured length is proportional to the mass of a larger measured length. Assume that the mass of 1.00 metre of Mg ribbon is 0.4368 g. Show that using 3.0 cm of this ribbon and 15 mL of 2.0 M HCl makes Mg the limiting reagent in the reaction between the two substances.

2. Show how the three reactions given in the introduction actually add up to give the formation reaction for MgO (i.e., set up the three equations and show how they add up, canceling like terms, etc.)

Technique

1. Transfer errors

In calculating the heat released by the reactions it is important to know the mass of the reaction mixture. Because we are using small amounts in this experiment we need to try and minimize the error that would normally not be very significant when, for example, liquid is added from a graduated cylinder.

In the Carbonate Project you used a technique of massing the amount of liquid added to a reaction by the difference of the graduated cylinder full of the liquid and the cylinder after the liquid had been poured out. That would be appropriate in this experiment once again.

The MgO is a very light powder and tends to stick to glass to some extent so it too is a problem if you plan on measuring some into a beaker and dumping it into the calorimeter. You could mass by difference but there is the additional problem of adding the powder to a calorimeter full of HCl. The reaction begins immediately at the surface and heat is lost (spray too) before the lid is replaced. One way to reduce this error is to measure the MgO directly into the calorimeter and then add the HCl quickly.

For this to work it is important that the calorimeter is dry.

2. Drying the calorimeter

These blocks are fairly soft and you can damage the surface coating if you are careless handling them. Drying the inside with our normal paper toweling requires a gentle touch. A final drying with a softer Kimwipe is a good idea.

3. Dealing with the thermometer probe

You need to suspend the probe so that the tip is in the solution but will not hit the stirring bar. One way to do this is to use your ring and ring stand. The probe cable can be draped over the ring and the probe will dangle over the calorimeter. A little adjustment to line things up, and it's done!

The chemicals

Magnesium oxide occurs in nature as the mineral *periclase*. It is a white, very fine, odorless powder. It gradually picks up CO_2 and water from the air, combining with the water to form magnesium hydroxide (milk of magnesia).

Magnesium oxide is used the manufacture of refractory crucibles and bricks as well as in medical applications as a laxative and antacid.

Analysis

These questions should be answered in your laboratory notebook following your data and observations.

1. Use your data to determine the heat of reaction *per mole of Mg* for the reaction between Mg and HCl. You did not use one mole of Mg, but enthalpy changes are proportional to moles. Don't forget to include the calorimeter constant in the calculations [the specific heat of the reaction mixture is 3.75 J/g^{o}C]

2. Use your data to determine the heat of reaction *per mole of MgO* for the reaction between MgO and HCl. You did not use one mole of MgO, but enthalpy changes are proportional to moles. Don't forget to include the calorimeter constant in the calculations [the specific heat of the reaction mixture is 3.75 J/g^{o}C]

3. Use the molar enthalpies from (1) and (2) above, and the enthalpy value given earlier for the formation of water to determine the heat of formation for MgO (ΔH^{o}_{f}) as you demonstrated in your answer to question #2 in the pre-lab quiz.

4. With the final result the question arises (as it does in any experiment) as to its accuracy. Using values obtained from handbooks, the following heats of reaction have been calculated:

Reaction (1)	-462 kJ/mol Mg
Reaction (2)	-146 kJ/mol MgO
Reaction (4)	-602 kJ/mol MgO

For each, determine the % error* in your experimental results and *briefly* discuss any significant deviations with specific reference to things you did and would not do again, etc.

$$* \quad \frac{\text{experimental - actual}}{\text{actual}} \times 100$$