# Experiment 2

## Heat of Combustion: Magnesium

**Purpose**

Hess's Law states that when are going from a particular set of reactants to a particular set of products, the heat of reaction is the same whether the reaction takes place in one step or in a series of steps. In this experiment, you will use Hess's Law to determine a heat of reaction that would be difficult to obtain by direct measurement—the heat of combustion of magnesium turnings. The reaction is represented by the equation:

(4) $Mg(s) + ½ O_2(g) \longrightarrow MgO(s)$

This equation can be obtained by combining equations (1), (2), and (3):

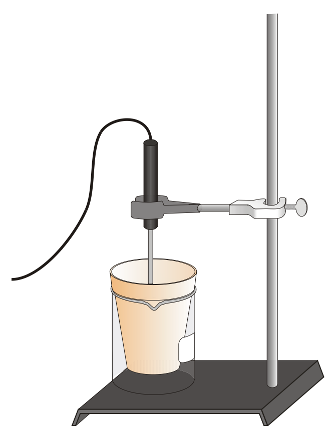
(1) $MgO(s) + 2 HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$

(2) $Mg(s) + 2 HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

(3) $H_2(g) + ½ O_2(g) \longrightarrow H_2O(l)$

You will use a Styrofoam cup in a beaker as a calorimeter, as shown in Figure 1. For purposes of this experiment, you may assume that the heat loss to the calorimeter and the surrounding air is negligible. Heat lost to either of these is a fairly constant factor in each part of the experiment, and has little effect on the final results.

*Figure 1: "Coffee cup" calorimeter*



# Experiment 2 Turn-in Sheets

## Heat of Combustion: Magnesium

**Procedure**

1. Obtain and wear safety glasses.

2. Prepare the computer for data collection by opening the Heat of Combustion experiment. The vertical axis has temperature scaled from 0 to 50°C. The horizontal axis has time scaled from 0 to 500 seconds.

**Reaction 1**

3. Place a Styrofoam cup into a 250-mL beaker as shown in Figure 1. Measure out 100.0 mL of 1.00 M HCl into the Styrofoam cup.

4. Use a utility clamp and a slit stopper to suspend a temperature probe from a ring stand as shown in Figure 1. Lower the temperature probe into the solution in the Styrofoam cup.

5. Weigh out about 1.00 g of magnesium oxide, MgO, on a piece of weighing paper. Record the actual mass used in your data table.

6. Click [Collect] to begin data collection and obtain the initial temperature, t1. After three or four readings at the same temperature (t1) have been obtained, add the white magnesium oxide powder to the solution. Use a stirring rod to stir the cup contents until a maximum temperature has been reached and the temperature starts to drop. Click [Stop] to end data collection.

7. Examine the initial readings in the Table window to determine the initial temperature, t1. To determine the final temperature, t2, click the Statistics button, [STAT]. The maximum temperature is listed in the statistics box on the graph. Record t1 and t2 in your data table.

8. Discard the solution as directed by your instructor.

**CAUTION**: Handle the HCl solution with care. It can cause painful burns if it comes in contact with the skin.

CAUTION: Avoid inhaling magnesium oxide dust.

**Reaction 2**

Repeat Steps 3-8 using about 0.50 g of magnesium turnings rather than magnesium oxide powder. Be sure to record the measured mass of the magnesium.

**CAUTION**: Do not breathe the vapors produced in the reaction!!

# Data

Construct an organized data table containing your measurements for Reactions 1 and 2.

Include a column for each of the following: The reaction (1, 2), the mass of solute (Mg or MgO) used, the moles of solute used, the total mass of the solution, the specific heat capacity of the solution, the change in temperature of the solution, the total change in enthalpy (ΔH, kJ), and the change in enthalpy per mol of solute (ΔH, kJ/mol). Read the next page for instructions on how to fill in this table.

Don't forget to include units in your table headings.

# Calculations

1. Determine the mass of the solution for each reaction (assume the density of each solution is 1.00 g/mL). You may include the mass of the solute. Fill these values into the data table on the previous page.

2. Calculate the change in temperature, $\Delta t$, for Reactions 1 and 2.

3. Calculate the heat released by each reaction, q, using the formula $q = Cp \bullet m \bullet \Delta t$

($Cp = 4.18$ J/g°C, and $m = 100.0$ g of HCl solution).

Convert joules to kJ in your final answer.

Use q to determine $\Delta H$ for both reactions. ($\Delta H = -q$)

Fill in the data table with this value.

4. Determine the moles of MgO and Mg used.

5. Use your Step 3 and Step 4 results to calculate $\Delta H$ per mol of MgO and Mg.

6. On the next page, you will fill in the heat of reaction ($\Delta H$/mol) for reactions 1 and 2.

Reaction 3 is the formation of liquid water. You can find the enthalpy of formation of water in Appendix J of your textbook.

The experimental enthalpy of reaction 4 can then be found using Hess' Law and the enthalpies of reactions 1-3.

You may use the space below and/or the back of this sheet to show your work.

# Results

Follow the directions from Step 6 in the Calculations section to fill in this table. Use Hess' Law (as in the ASA) to find $\Delta H_4$.

| Reactions | ΔH (kJ/mol) |
|---|---|
| 1. $MgO_{(s)} + 2\,HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_2O_{(l)}$ | $\Delta H_1$ = |
| 2. $Mg_{(s)} + 2\,HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$ | $\Delta H_2$ = |
| 3. $H_{2(g)} + 1/2\,O_{2(g)} \longrightarrow H_2O_{(l)}$ | $\Delta H_3$ = |

| | |
|---|---|
| 4. $Mg(s) + ½\,O_2(g) \longrightarrow MgO(s)$ | $\Delta H_4$ =<br><br>(by using Hess' law)<br>[Experimental] |

| | |
|---|---|
| $Mg(s) + ½\,O_2(g) \longrightarrow MgO(s)$ | ΔH =<br><br>(from Appendix J)<br>[Accepted] |

# Calculation and Discussion of % error:

Using the value in Appendix J for the formation of MgO as the accepted value and $\Delta H_4$ determined above as the experimental value, find the percent error for the experiment.

**1) Calculate your percent error.**

$$\%\,\text{error} = \left|\frac{\text{experimental - accepted}}{\text{accepted}}\right| \times 100\%$$

**2) Describe two ways you would you improve the design of the calorimeter.**

**3) Give two possible sources of experimental error.** (Experimental error refers to the uncertainties associated with experimental limitations, and not necessarily mistakes made by the experimentalist.)

# Pre-Lab Assignment

This experiment covers material from Sections 8.7 and 8.8 in your textbook.
Refer to these sections for assistance with the following questions.

1) Construct a data table for Page 4 of this lab. Read the instructions for what to include. The table will not be turned in with the ASA but it may be checked during lab.

2a) Use Hess's Law to show how equations 1-3 can be summed to obtain equation 4.
(Re-write each equation in the form necessary to produce equation 4, and show Hess' law by crossing out terms that cancel.) Refer to your textbook on how to do this.

(1) $MgO_{(s)} + 2\ HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_2O_{(l)}$ $\Delta H_1$

(2) $Mg_{(s)} + 2\ HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$ $\Delta H_2$

(3) $H_{2(g)} + 1/2\ O_{2(g)} \longrightarrow H_2O_{(l)}$ $\Delta H_3$

(4) $Mg_{(s)} + 1/2\ O_{2(g)} \longrightarrow MgO_{(s)}$ $\Delta H_4$

b) Write an equation for how $\Delta H_4$ will be solved using the terms $\Delta H_1$, $\Delta H_2$, and $\Delta H_3$.
(Example: Will $\Delta H_4 = \Delta H_1 - \Delta H_2 + \Delta H_3$? or $\Delta H_1 + \Delta H_2 + \Delta H_3$? etc…)

3) A student measures out 100.0 ml of 1.00 M HCl and places it in a styrofoam cup. The temperature of the HCl is 22.3$^0$C. She then addes 1.002 g of MgO to the cup. The temperature rises to 29.7$^0$C.

a) Determine the enthalpy of the reaction. Assume that the density and the specific heat of the solution are the same as water. Include the correct sign.

b) Determine the change of enthalpy for the reaction in units of kJ per mol of MgO. Include the correct sign!