Thermochemistry
Chem 111/112 Laboratory Experiment

Part I.

Introduction

The purpose of Part 1 of this experiment is to determine the enthalpy of formation of a compound, magnesium oxide, MgO. The formation reaction is,

$$Mg(s) + \tfrac{1}{2}\, O_2(g) \quad \rightarrow \quad MgO(s) \qquad\qquad \Delta H_f = ?$$

To determine this, we will perform two reactions and measure their enthalpy change using calorimetry.

$Mg(s) + 2\, H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$ $\quad$ ΔH_1 = we will find this experimentally

$MgO(s) + 2\, H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(\vartheta)$ $\quad$ ΔH_2 = we will find this experimentally

$\tfrac{1}{2}\, O_2(g) + H_2(g) \rightarrow H_2O(\vartheta)$ $\quad$ ΔH_3 = this can be found in a data table

Once the three values, ΔH_1, ΔH_2 and ΔH_3, are obtained, Hess's Law can be used to combine those values such that the desired heat of formation can be calculated.

EXPERIMENTAL PROCEDURE

PART Ia. Determining ΔH_1 and the Relation Between the Quantity of Material Reacting and the Heat Transferred

In this portion of the experiment we want to explore the relationship between the quantity of magnesium metal reacting with hydrochloric acid and the heat evolved by the reaction.

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

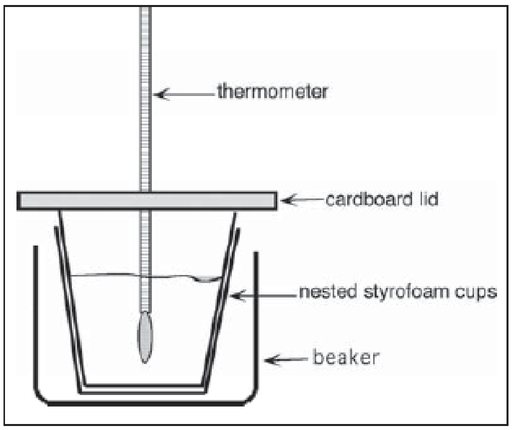
(a) Set up your coffee-cup calorimeter and support it in a ring stand or beaker as illustrated by your instructor and in the Figure.

(b) Weigh out, to the nearest 0.001 g, three different portions of magnesium metal, say about 0.2 g, about 0.4 g, and about 0.5 g. Record the masses in your lab notebook.

(c) Place one of the magnesium samples in a clean, dry coffee-cup calorimeter.

(a) Calorimeter consists of thermometer, two styrofoam cups, paper cover, and beaker.

(b) Two cups together, supported in a beaker. Cover in place with thermometer through the hole in the cover.



(c) Diagram of the experimental setup for the thermochemistry experiment.

Figure Equipment and setup for the thermochemistry experiment.

(d) Using a graduated cylinder, measure out as accurately as possible 100.mL of 1.0 M HCl. Measure the temperature of the HCl solution and record this initial temperature in your lab notebook. Replace the thermometer in the calorimeter setup.

(e) Add the HCl solution to the calorimeter and swirl gently but steadily. At 30 second intervals record the temperature (to the nearest 0.5 ˚C) until the temperature has held constant or decreased for three consecutive readings.

(f) Repeat the steps above with another sample of Mg. Do all three samples, making sure the calorimeter is clean and dry each time.

(g) In each case, determine ΔT, the change in temperature between the temperature of the HCl solution before adding it to the magnesium and the maximum temperature of the reacting system.

PART Ib: Determining ΔH_2: Heat of Reaction of Magnesium Oxide with Hydrochloric Acid

(a) Set up the calorimeter as in PART I. Make sure it is clean and dry.

(b) Weigh out about 0.7 g of MgO to the nearest 0.001 g and place the powder in the calorimeter.

(c) Repeat steps (d) and (e) as in PART I above with another sample of magnesium oxide.

<u>Calculations and Report</u>

In each reaction, we will assume that the calorimeter absorbs no heat. Therefore, the amount of heat energy released is determined by the temperature rise of the solution. Each solution is dilute, so we will assume that they have a specific heat capacity of 4.18 J/g·°C and a density of 1.00 g/mL.

Use the equation,

$$\# \text{ J} = \text{Specific heat capacity x mass (g) x } \Delta T$$

to determine the quantity of heat released. Then, using this and the number of moles of compound reacting, determine the value of ΔH for each reaction.

$$\Delta H = \frac{\# \text{ J}}{\# \text{ moles}}$$

Remember that if heat is released, then ΔH must be a negative value.

Report the value of ΔH for each reaction.

Use the three reactions performed in Part 1 to verify if the amount of heat released is proportional to the number of moles of compound reacting. **Report** on whether this is true.

Use the ΔH values determined experimentally as well as that for formation of H_2O liquid to calculate the value of ΔH of formation for MgO(s). To do so, you must find a way of combining the three reactions for which you know ΔH so that they add up to give the reaction of interest. Remember that you can reverse reactions (and then must switch the sign of ΔH) or multiply reactions by a constant (and then must multiply the ΔH by that same constant). Report how you do this calculation and the final value of the heat of formation of MgO(s).

Part II.

Exploring the heat of neutralization of acids and bases.

Introduction
All acid-base reactions involve the same basic step: the transfer of an H^+ ion (which is just a proton) from the acid to the base. An example is the reaction between NH_4^+ and F^- ions.

$$NH_4^+(aq) + F^-(aq) \rightarrow NH_3(aq) + HF(aq)$$

In this case the ammonium ion, NH_4^+, is the acid and fluoride ion (F^-) is the base.

In this experiment you will perform reactions between the following two acids and bases:

<u>Acids</u>	<u>Bases</u>
hydrochloric (HCl)	ammonia (NH_3)
Acetic (CH_3CO_2H)	sodium hydroxide (NaOH)

You will react each acid with each base, so you will perform four reactions.

Procedure
Similarly to Part I, pour 25 mL of the acid (measured with a graduated cylinder) into the Styrofoam cup and measure its temperature. Add 25 mL of the base solution, place the lid with thermometer on top, stir, and measure the temperature change.

Calculations
Calculate the enthalpy change for each reaction in terms of kJ/mol.

Report
Report the four enthalpy changes you determine and explain the trends you see in terms of the strengths of the acids and bases used. As part of your analysis of the results, write net-ionic equations for each of the reactions.