Enthalpy of Formation of Magnesium Oxide

Objectives
- To determine the enthalpy of formation of magnesium oxide by measuring the enthalpy changes for two reactions by constant–pressure calorimetry and applying Hess's law to the collected data and calculations.

Prior Reading

Consult your textbook for a discussion of enthalpy, constant–pressure (or *coffee–cup*) calorimetry, Hess's law and formation reactions.

Equipment & Reagents

coffee–cup calorimeter (two styrofoam cups nestled inside each other) and lid, thermometer, glass stirring rod, 100–mL graduated cylinder, solution of 1 *M* hydrochloric acid, magnesium ribbon, magnesium oxide

Discussion

Enthalpy is one of several quantities, called state functions, that describe a system. Enthalpy is often referred to as the *heat content* of a system and is given the symbol *H*. Although absolute values of enthalpy cannot be measured, a change in enthalpy between reactants and products or between two states can be determined. This change in enthalpy, ΔH, equals the heat flow in the process at constant pressure as depicted by the equation below in which q_P represents the heat flow under constant–pressure conditions.

$$\Delta H = q_P$$

Calorimetry measures heat flow, and measurements at constant pressure lead to the determination of enthalpy changes according to the above equation. Constant–pressure conditions are easily obtained in a calorimeter that has a small opening to the atmosphere. Atmospheric pressure remains relatively constant over the duration of a short experiment and buffers small changes in pressure that may occur in the calorimeter. Styrofoam coffee cups with lids that do not fit too tightly are amazingly effective constant–pressure calorimeters and are used in this experiment.

Calorimeters are isolated systems in which neither energy nor mass is exchanged with the surroundings. A complete system consists of three parts—a reaction, the surrounding solution and the walls of the calorimeter. The fundamental relationship of calorimetry states that energy as heat is simply transferred among these three components of the system (*i.e.*, the reaction itself, the solution in which the reactions occur and the walls of the calorimeter) but not out of the calorimeter. In equations, these concepts are represented as

$$q_{sys} = 0 \quad \text{(no heat leaves the calorimeter)}$$

$$q_{sys} = q_{rxn} + q_{soln} + q_{cal}$$

The subscripts refer to the heat flow within the system (*sys*), reaction (*rxn*), surrounding solution (*soln*) and walls of the calorimeter (*cal*).

The part of the system of interest is the heat of reaction, q_{rxn}, and solving the above equations for q_{rxn} gives the following relationship

$$q_{rxn} = -(q_{soln} + q_{cal})$$

Often the heat that flows into or out of the walls of the calorimeter, q_{cal}, is small in comparison to the total heat flow and can be neglected. Such a situation applies in this experiment, and the last equation then reduces to

$$q_{rxn} = -q_{soln}$$

The heat flow of the reaction (q_{rxn}) is opposite in sign to the heat flow into or out of the solution that surrounds the reaction (q_{soln}). In an *exothermic* reaction, heat flows out of the reaction into the surrounding solution, and the heat q is given a negative sign. For an *endothermic* reaction, heat flows from the surrounding solution into the reaction, and the sign for q is positive.

The figure below depicts a calorimeter with the three components of the system discussed above. The reaction is shown with a discreet boundary, but in reality the reaction is uniformly distributed throughout the solution. A hypothetical boundary is drawn in this model for clarity in thinking about heat flow in the system. Heat (q) is flowing from the reaction into the solution, and thus the reaction depicted is exothermic.
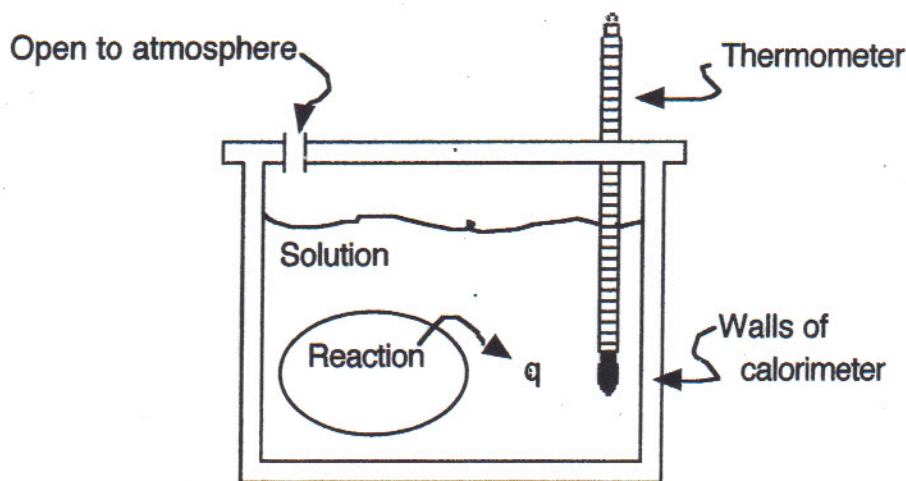
Constant–Pressure Calorimeter

In a calorimeter, temperature measurements are made in the solution surrounding the reaction, and the heat flow that is measured is q_{soln}. Knowledge of the specific heat capacity (*s*) of the solution, its mass (*m*) in grams and the difference in final and initial temperatures, (ΔT), leads
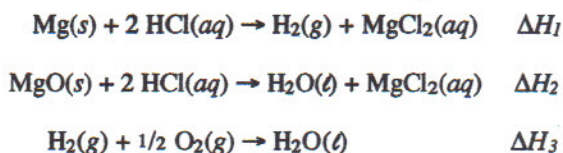
to the equation for calculating the heat flow into or out of the solution.

$$q_{soln} = (s_{soln})(m_{soln})(\Delta T_{soln})$$

The units for specific heat capacity are $J \cdot g^{-1} \cdot {}^{\circ}C^{-1}$. When s is multiplied by the mass in g and ΔT in $^{\circ}C$, the product q is in J.

The first step in a calorimetry experiment is to determine q_{soln}. Then changing the sign of q_{soln} gives q_{rxn}. These values represent heat flow in J for the *given quantity of the reacting substance*. Determining ΔH requires calculating q_{rxn} for a *mole* of the reacting species. Conversion from grams to moles requires the molar mass of the reactant. Your textbook discusses specific examples of these calculations.

In this experiment, you will determine ΔH for two reactions by calorimetry and will use the *accepted value of ΔH for a third reaction*. These reactions are

$$Mg(s) + 2\ HCl(aq) \rightarrow H_2(g) + MgCl_2(aq) \qquad \Delta H_1$$

$$MgO(s) + 2\ HCl(aq) \rightarrow H_2O(\ell) + MgCl_2(aq) \qquad \Delta H_2$$

$$H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(\ell) \qquad \Delta H_3$$

These data (ΔH_1, ΔH_2, ΔH_3) are then combined to calculate the enthalpy change for the formation of magnesium oxide. A *formation reaction* is one in which one mole of a pure substance is formed from its pure elements in their most stable form. Thus for magnesium oxide, the formation reaction is

$$Mg(s) + 1/2\ O_2(g) \rightarrow MgO(s) \qquad \Delta H_f$$
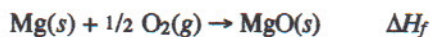
The symbol for the enthalpy of formation includes the subscript f as shown above ΔH_f.

The thermochemical relationship that allows you to combine reactions and ΔH values is called Hess's law. This law states that if you can add a set of reactions to give an overall reaction, then ΔH for the overall reaction is simply the sum of the ΔH values for each reaction in the set. Again, your textbook discusses Hess's law in some detail.

You might note that enthalpy data for formation reactions are usually given under standard state conditions represented by the superscript "$^{\circ}$" in the term ΔH_f°. Standard state conditions for a substance are the most stable form of the substance at 1 atm pressure and a specified temperature, which is usually 25°C. In this experiment, you will be working close enough to standard state conditions at 25°C and 1 atm so that you can use the tabular values in your textbook for the standard enthalpy of formation of magnesium oxide to compare your results.

Experimental Procedure

A.   Reaction with magnesium

Weigh a dry calorimeter (without its cover) to the nearest 0.001 g. Add 100 mL of 1 M HCl(aq) from a graduated cylinder, and reweigh the calorimeter with the solution. This second mass can only be recorded to the nearest 0.01 g.

Obtain a small piece of magnesium ribbon, and weigh approximately 0.24 g. Be sure to record this measurement to the nearest 0.001 g.

Assemble the calorimeter with its lid, thermometer and stirring rod as shown in the figure below.

Coffee–Cup Calorimeter
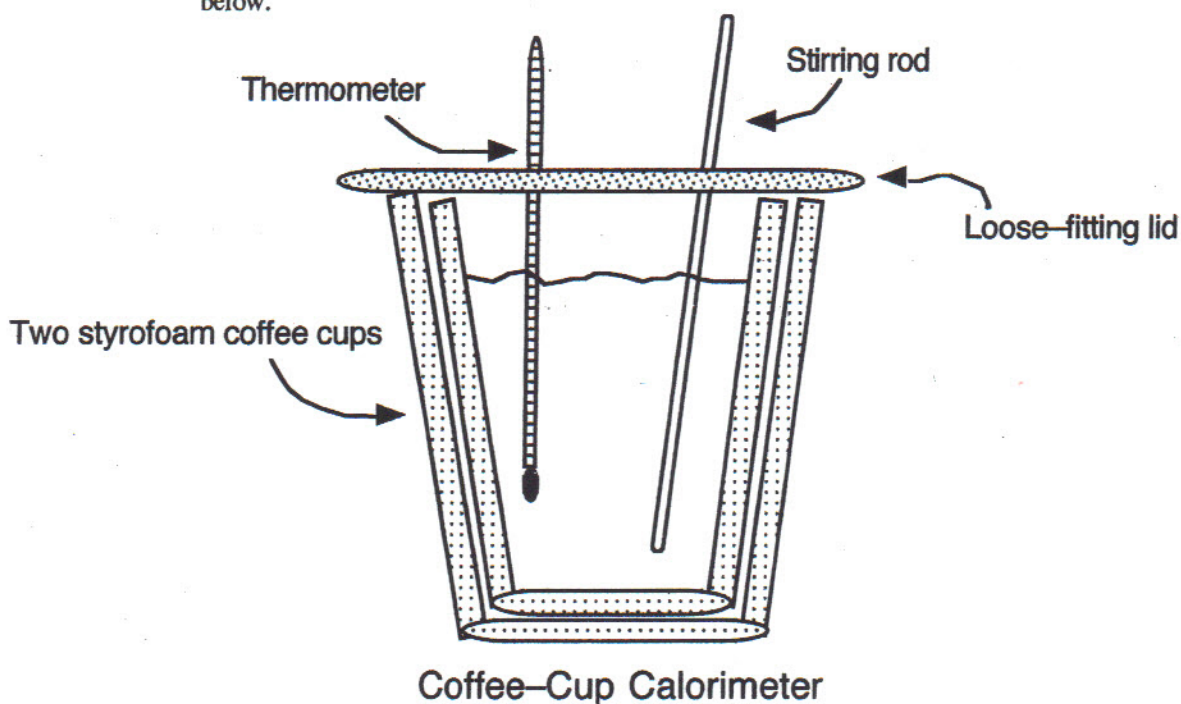
Record the temperature of the hydrochloric acid to the nearest 0.1°C. Remove the cover of the calorimeter and add the magnesium. Then quickly replace the cover. Observe the temperature and gently stir the solution with the glass rod until the temperature stops rising. Record the highest temperature to the nearest 0.1°C.

Empty the calorimeter contents into the waste container provided. Dry the calorimeter with a paper towel.

B.   Reaction with magnesium oxide

Repeat the above procedure except use approximately 0.8 g of magnesium oxide (measured to the nearest 0.001 g) instead of magnesium ribbon.

Empty the calorimeter contents into the waste container and dry the calorimeter.

C.   Calculations

Calculate the mass of hydrochloric acid and ΔT for both reactions by subtraction of the appropriate quantities.

From these respective values, calculate q_{soln} for each reaction using a *specific heat capacity for 1 M HCl as 3.64 J·g-1·°C-1*. From q_{soln}, determine q_{rxn} for each reaction. You may assume that q_{cal} in this experiment is negligible.

Finally, calculate q_{rxn} values on a mole basis using the molar masses for magnesium and magnesium oxide, respectively, and convert from J to kJ. *When q_{rxn} is expressed in terms of kJ (or J) per mole at constant pressure, it becomes ΔH.*

Refer to the *Discussion* of this experiment, and write the balanced equations that you will be using and their respective ΔH values. The ΔH value for the third reaction can be found in the tables for standard enthalpy of formation in your textbook.

Rearrange these equations so that when added they will give the formation reaction for magnesium oxide. Follow the laws of thermochemistry and change ΔH values accordingly. Lastly, add the rearranged equations and ΔH values to give the net reaction and ΔH_f for the formation of magnesium oxide.

Look up the accepted value for ΔH_f under standard state conditions in your textbook, and calculate the percent error of your experiment. Be careful in reporting your percent error to the correct number of significant figures.

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Pre–Laboratory Questions

Name: _____

Date: _____

1. Write and balance the **two** chemical reactions that you will use in the calorimeter analyses in this laboratory.

2. In this experiment you will use *Hess's law* to obtain the enthalpy change for the formation of magnesium oxide. Write a sentence (or two) in your own words that explains *Hess's law*.

3. A student observed that when 0.243 g of magnesium was added to a 100.0 g of hydrochloric acid, the temperature of the solution increased 11.2°C. Without doing any calculations, how can you claim that the reaction that occurred between magnesium and hydrochloric acid was *exothermic*?

Enthalpy of Formation of Magnesium Oxide

Data & Results

Name: _____

Lab Partner: _____

Date: _____

Note: Read thermometers to ±0.1°C, measure small masses to ± 0.001 g and large masses to ±0.01 g, and include units on all measurements and calculations.

I. Measurements

Reaction with:	Magnesium	Magnesium Oxide
Mass of dry calorimeter		
Mass of calorimeter containing HCl		
Mass of reactant		
Initial temperature of HCl		
Final temperature of HCl		

II. Calculations

Note: Specific heat 1 M HCl = 3.64 $J \cdot g^{-1} \cdot °C^{-1}$

Reaction with:	Magnesium	Magnesium Oxide
m_{soln}		
ΔT		
q_{soln} (in J)		
q_{rxn} (in J)		
q_{rxn} (in kJ)		
mass of reactant used (from above)		
molar mass of reactant used		
mol reactant used		
ΔH (in kJ/mol reactant)		

III. Reactions & Enthalpy Data

A. Reactions studied and measured enthalpy data

(1)	$\Delta H_1 =$
(2)	$\Delta H_2 =$
(3)	$\Delta H_3 =$

B. Rearranged reactions and enthalpy data to be added

(1)	$\Delta H_1 =$
(2)	$\Delta H_2 =$
(3)	$\Delta H_3 =$

C. Net reaction and enthalpy of formation of MgO

	$\Delta H_f =$

D. Calculation of percent error

Accepted value for ΔH_f for MgO	$\Delta H_f =$
Percent error	

Name: _____

Date: _____

1. Calculations based on measurements need to be reported to the correct number of significant figures, and in this experiment you calculated ΔH for two reactions. List the measurements that you used for each reaction, indicate how many significant figures were in **your** recorded data, and lastly indicate how many significant figures should be in **your** reported value of ΔH.

Reaction	Quantities Measured / Values	No. Sig. Figures	ΔH *(with correct sig. figures)*
Mg + HCl	1. mass Mg		
	2. mass HCl		$\Delta H_1 =$
	3. T_i and T_f		
	4. ΔT		
MgO + HCl	1. mass MgO		
	2. mass HCl		$\Delta H_2 =$
	3. T_i and T_f		
	4. ΔT		

2. Based on **your calculation** of ΔH (in kJ·mol⁻¹) for the reaction of magnesium oxide with a 1 *M* solution of hydrochloric acid, determine the temperature change for the following conditions. Show your work to receive full credit.

A 1.210–g sample of MgO was dissolved in 85.00 g of 1 *M* HCl. Assume that no heat was absorbed by the walls of the calorimeter and that the specific heat of the hydrochloric acid solution was 3.64 $J \cdot g^{-1} \cdot {}^{\circ}C^{-1}$.

Your Value of ΔH used: _____

$\Delta T =$ _____