Experiment 27

FV 8/7/02

HEAT TRANSFER FROM CHEMICAL REACTIONS

MATERIALS: Magnesium metal ribbons, magnesium oxide powder, copper wire (for weighting of Mg ribbon), 100 mL graduated cylinder, magnetic stirrer, 1.00 M hydrochloric acid, coffee cup calorimeter with lid, thermometer.

PURPOSE: The purpose of this experiment is to determine the enthalpy of formation of magnesium oxide.

LEARNING OBJECTIVES: By the end of this experiment, the student should be able to demonstrate the following proficiencies:

1. Quantitatively observe temperature changes during chemical reactions.
2. Develop an understanding of enthalpy.
3. Calculate the enthalpy change of a reaction from calorimetry data.
4. Apply Hess's Law to calculate the enthalpy change associated with a reaction.
5. Explore the additive nature of thermodynamic parameters.
6. Learn how to evaluate "pooled" experimental data.
7. Use a spreadsheet program for data manipulation, graphing, and regression analysis.

DISCUSSION:

Elemental magnesium is one of the principal components of flares used to illuminate nighttime activities, or to aid in signaling one's location to aircraft and ships. Your instructor will ignite a strip of magnesium ribbon to demonstrate the combustion of magnesium in air. It will be evident that a great deal of light energy is released from this reaction. Some chemical reactions (including this one) are also associated with the evolution of thermal energy and are called *exothermic* reactions. When there is an absorption of energy in a chemical reaction, the process is called *endothermic*. The magnitude of the energy change is determined by the particular reaction as well as the amount of product(s) formed. The thermal energy transferred in a balanced chemical reaction carried out at constant pressure is called the *enthalpy of reaction* (or heat of reaction) and is given the symbol ΔH_{rxn}. ΔH_{rxn} is often expressed in units of kJ/mole where mole refers to the amount of a reactant or a product involved in the reaction. In general, the *reactant or product must be specified.*

In this experiment, you will measure the enthalpy changes of exothermic reactions utilizing a simple calorimeter, which consists of a Styrofoam cup, lid, and thermometer. The energy given off by the reaction is absorbed by the calorimeter and the solvent (water in this experiment) resulting in a temperature increase for the calorimeter and the solvent. Assuming the calorimeter absorbs a negligible amount of heat compared to that absorbed by the solvent, the heat evolved by the reaction (q_{rxn}) is calculated using the equation:

$$q_{rxn} = - m \cdot s \cdot \Delta T = - m \cdot s \cdot (T_f - T_i) \qquad (1)$$

where s = specific heat of the solvent = 4.18×10^{-3} kJ·g^{-1}·°C^{-1} (Note: the specific heat of the solvent is assumed to be that of pure water.)

m = mass of solvent (water) in grams (Note: the density of the solvent can be assumed to be 1.00 g/mL even though you will use 1.00 M HCl rather than pure water.)

T_i = the initial temperature of the solvent (in °C)

T_f = the final temperature the solvent (in °C) would reach if there were no heat losses from the calorimeter during the reaction

NOTE: The (-) sign in front of the m in equation (1) is used to indicate that the heat being given off by the reaction is the same magnitude, but opposite sign, as the heat absorbed by the solvent:

$q_{rxn} = - q_{solvent}$ (i.e., when $T_f > T_i$, $q_{solvent} = q_{surroundings}$ will be (+) and $q_{rxn} = q_{system}$ will be (-))

The enthalpy change (ΔH_{rxn}) for the reaction per mole of specified reactant is found by using the equation:

$$\Delta H_{rxn} = q_{rxn} / \text{moles of specified reactant} \qquad (2)$$

The main experimental problem in calorimetry is obtaining an accurate value of ΔT. The initial temperature, T_i, of the reactants can be determined directly using a thermometer. However, it is difficult to obtain a precise value for the final temperature, T_f, which is the instantaneous temperature when the reactants are mixed together and react. This is an open system so heat exchange will occur between the room and the contents of the calorimeter, unlike an isolated system such as a bomb calorimeter. An open system allows the reaction to proceed at constant pressure at the sacrifice of the loss of generated heat to the surroundings. This occurs both during the reaction and after its completion. The rate of heat exchange depends upon the insulating properties of the calorimeter, the rate of reaction, and the effectiveness of mixing the reactants. A correction for this heat exchange is made by an extrapolation process using a temperature versus time curve. First, a plot of the temperature readings as a function of time for the reaction is generated. By extrapolating the linear portion of the curve back to zero time (the time when the reactants were mixed in the calorimeter), T_f is obtained. The T_f determined in this manner will be the temperature that the calorimeter and the solvent would have reached, had the reaction occurred instantaneously and with no heat exchange to the room. This value should be used for the calculation of change in temperature, ΔT. Consult Appendix O for specific instructions for extrapolation using Microsoft Excel and to see what a typical temperature versus time plot looks like.

Enthalpy Changes for Chemical Reactions:

The thermodynamic parameter enthalpy is a state function, which means that the value for a given chemical change is independent of how the change is done. Consequently, the enthalpy change associated with the conversion of a set of reactants to a set of products is independent of how many steps (the path) are involved in the conversion. Consider the formation of dinitrogen tetroxide gas.

Path 1:		
	$2\ NO_2\ (g) \rightarrow N_2O_4\ (g)$	ΔH_{rxn1}

Path 2:		
Step 1	$2\ NO_2\ (g) \rightarrow N_2\ (g) + 2\ O_2\ (g)$	ΔH_{rxn2}
Step 2	$N_2\ (g) + 2\ O_2\ (g) \rightarrow N_2O_4\ (g)$	ΔH_{rxn3}
Net	$2\ NO_2\ (g) \rightarrow N_2O_4\ (g)$	ΔH_{rxn1}

The sum of the reactions in *Path 2* is the reaction in *Path 1*. Since Paths 1 and 2 begin and end with the same compounds, the total change in enthalpy along both paths is identical: $\Delta H_{rxn2} + \Delta H_{rxn3} = \Delta H_{rxn1}$.

Name _____ Section _____

Partner _____ Date _____

PROCEDURE AND DATA SECTION
Experiment 27

Before beginning this experiment, complete all In-Lab Questions found on page E27-7.

What solid are you studying? _____

Write the chemical reaction you are studying. _____

What is your assigned target temperature change? _____

Considering the materials available for this lab, generate a step by step procedure to determine the enthalpy change of the reaction that is assigned to you by your instructor. The procedure should include the compound(s) needed, a brief description of the mechanical execution of the experiment, and a description of any special procedures to ensure all the heat from the chemical reaction is delivered to the solvent in the calorimeter.

These guidelines must be part of your plan:
1. You will need to plan to collect data for **at least 8 minutes per run** to construct a plot of temperature vs. time.
2. You **must use 100 mL of solvent** in each run (so that data from all groups can be compared).
3. You should generate (within 1 or 2 °C) the target temperature change for your reaction as designated by your instructor.
4. **No more than** 0.7 g of Mg metal or 2.5 g of MgO should be needed for any of the assigned temperature changes.

Summarize the steps of your procedure here:

Be sure to verify your procedure with the instructor before performing any chemical reactions.

Perform a test run using your procedure. Be sure to record the amount of solid used for the run and numerous temperature/time measurements. Use the blank page (E27-4) to record all experimental details. Consider the three questions below for each trial run you complete. Sketch a plot of temperature vs. time on graph paper or on your calculator, and use your plot to estimate the final temperature, T_f. Show your graph to your instructor.

How much solute did you use in your test run? _____

What was the temperature change observed in your test run? _____

How should you modify your procedure to achieve the desired temperature change?

Carry out "test runs" until you obtain your assigned temperature change (within 1 or 2 °C). Be sure to carefully record all experimental details as you proceed.

DATA SECTION
Experiment 27

DATA TREATMENT
Experiment 27

1. Using only the data from the trial that resulted in your assigned temperature change (thus ignoring data from any test runs), determine the temperature change (T_f - T_i) for the reaction you studied by constructing a temperature versus time plot with Excel. Report this temperature change and the mass of reactant that led to this change to the class. (Refer to Appendix O for help in determining T_f using Excel.)

2. Create an Excel spreadsheet that lists the results from the entire class (ΔT and mass of reactant for each of the two reactions studied).

3. Using an Excel spreadsheet, calculate the heat of reaction, q_{rxn}, and the number of moles of reactant used for each mass and temperature change for each of the two reactions studied and reported by your classmates.

4. Also in Excel, construct a plot of the heat of reaction vs. the number of moles of reactant for each of the two reactions studied (thus you will have to make two plots for this part).

5. The slope of each of these plots is ΔH_{rxn} (per mole) for that reactant (see equation 2 on page E27-2). Record the ΔH_{rxn} (per mole) for each reactant here.

ΔH_{rxn} (Mg) = _____ ΔH_{rxn} (MgO) = _____

6. Using the scheme you developed in In-Lab question 4b and the values for ΔH_{rxn} that you found graphically, determine the enthalpy change of the combustion reaction in which Mg combines with O_2 (also called the enthalpy of formation, ΔH_f, for the product formed). Show your work.

7. Look up the ΔH of formation, ΔH_f, of MgO (s) in your textbook (see Appendix C). Calculate the percent deviation between your value and this literature value.

8. Speculate as to the accuracy and precision of collecting data as a class as compared to collecting data individually or in small groups. Explain your answer.

9. Military Applications of the reaction studied: Go to http://www.fas.org/man/dod-101/sys/dumb/luu2.htm and in one short paragraph, using complete sentences describe how the reactions studied in this experiment are related to your military career.

Name _____ Section _____

Date _____

IN-LAB QUESTIONS
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TURN IN THIS PAGE BEFORE YOU LEAVE LAB.

1. Write a balanced chemical reaction for the combustion of Mg (s) with O_2 (g).

2. Write the net ionic equation for the reaction of 1.00 M HCl with magnesium metal ribbon.

3. Write the net ionic equation for the reaction of 1.00 M HCl with the product from the combustion reaction that your instructor performed (see question 1 above).

4. The enthalpy of formation is defined as the enthalpy change for the formation of 1 mole of a substance from the most stable forms of its component elements. An example of this would be in the formation of water:

$H_{2\,(g)} + ½\,O_{2\,(g)} \rightarrow H_2O_{(l)}$ $\Delta H_{rxn} = -285.84$ kJ/mole ← this can also be abbreviated ΔH_f since this is a formation reaction.

a. Write the reaction whose energy change would yield the enthalpy of formation (ΔH_f) of the product in question 1 above.

b. Propose a **set of reactions** (similar to the example shown on page E27-2) that would eventually yield the reaction for the enthalpy of formation for the reaction you wrote down in question 4a. You will need to include the reaction for the formation of water found above in your scheme. **Verify your overall scheme with your instructor prior to going any further.**