Honors Chemistry Lab #22: Magnesium Enthalpy of Reaction (Inquiry Version)

Introduction:

When we study energy changes in chemical reactions, the most important quantity is usually the enthalpy of reaction (ΔH_{rxn}), the change in enthalpy that occurs during a reaction, such as the dissolution of a piece of copper in nitric acid. If heat flows from a system to its surroundings, the enthalpy of the system decreases, so ΔH_{rxn} is negative. Conversely, if heat flows from the surroundings to a system, the enthalpy of the system increases, so ΔH_{rxn} is positive. In chemical reactions, bond breaking requires an input of energy and is therefore an endothermic process, whereas bond making releases energy, which is an exothermic process. If ΔH_{rxn} is negative, then the enthalpy of the products is less than the enthalpy of the reactants; that is, an exothermic reaction is thermodynamically favorable. Conversely, if ΔH_{rxn} is positive, then the enthalpy of the products is greater than the enthalpy of the reactants; thus, an endothermic reaction is energetically uphill. Two important characteristics of enthalpy and changes in enthalpy are summarized in the following discussion of ice. Reversing a reaction or a process changes the sign of ΔH. Ice absorbs heat when it melts (electrostatic interactions/IMFs are broken), so liquid water must release heat when it freezes (electrostatic interactions/IMFs are formed). Enthalpy is an extensive property, which means it is dependent on the amount of the substance. Therefore, the magnitude of ΔH for a reaction is proportional to the amounts of the substances that react. For example, a large fire produces more heat than a single match, even though the chemical reaction (i.e., the combustion of wood) is the same in both cases. For this reason, the enthalpy change for a reaction is usually given in kilojoules per mole of a particular reactant or product. Consider the thermochemical equation, which describes the reaction of aluminum with iron(III) oxide (Fe2O3) at constant pressure. According to the reaction stoichiometry, 2 mol of Fe, 1 mol of Al2O3, and 851.5 kJ of heat are produced for every 2 mol of Al and 1 mol of Fe2O3 consumed: $2Al_{(s)}+Fe_2O_{3(s)} \rightarrow 2Fe_{(s)}+Al_2O_{3(s)}+815.5$ kJ. One way to report the heat absorbed or released would be to compile a massive set of reference tables that list the enthalpy changes for all possible chemical reactions, which would require an incredible amount of effort. Fortunately, since enthalpy is a state function, all we have to know is the initial and final states of the reaction. This allows us to calculate the enthalpy change for virtually any conceivable chemical reaction using a relatively small set of tabulated data, such as the following:

Enthalpy of combustion (ΔH_{comb}) The change in enthalpy that occurs during a combustion reaction. Enthalpy changes have been measured for the combustion of virtually any substance that will burn in oxygen; these values are usually reported as the enthalpy of combustion per mole of substance.

Enthalpy of fusion (ΔH_{fus}) The enthalpy change that acompanies the melting (fusion) of 1 mol of a substance. The enthalpy change that accompanies the melting, or fusion, of 1 mol of a substance; these values have been measured for almost all the elements and for most simple compounds.

Enthalpy of vaporization (ΔH_{vap}) The enthalpy change that accompanies the vaporization of 1 mol of a substance. The enthalpy change that accompanies the vaporization of 1 mol of a substance; these values have also been measured for nearly all the elements and for most volatile compounds.

Enthalpy of solution (ΔH_{soln}) The change in enthalpy that occurs when a specified amount of solute dissolves in a given quantity of solvent. The enthalpy change when a specified amount of solute dissolves in a given quantity of solvent.

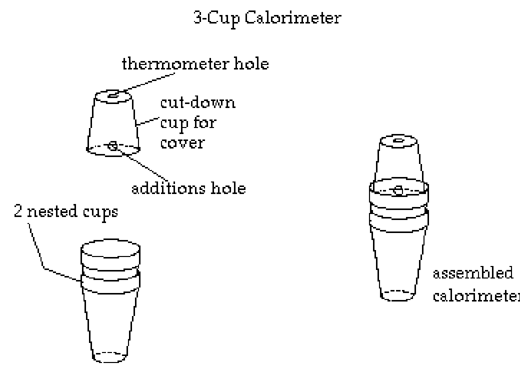
In this lab, you will design an experiment to calculate the enthalpy of reaction (ΔH_{rxn}) for magnesium in hydrochloric acid using calorimetry. After conducting multiple replicates of your experiment, you will compare your experimental values to the accepted value of -462.0 kJ/mol at 25 °C for 1.0 M hydrochloric acid.

Materials: Styrofoam calorimeter, thermometer, digital thermometer, LabQuest (optional), magnesium, 1.0 M HCl

Suggested Procedures:

Suggestion 1: Obtain a ~2-3 cm piece of polished magnesium. Measure the mass of the magnesium ribbon on an analytical balance. Assemble a Styrofoam calorimeter as shown in the diagram.

Suggestion 2: Use 50-100 mL of 1.0 M hydrochloric acid into a Styrofoam calorimeter. Measure and record the exact volume. The specific heat capacity of the hydrochloric acid is the same as water. Caution should always be exercised when using dilute acids.



Suggestion 3: Create a data table for your data set and show every single calculation used in the investigation to determine the enthalpy of reaction.

Pre-Lab Questions:

1. Using the accepted value, write a thermochemical equation for the reaction of magnesium metal with hydrochloric acid.

2. Assign oxidation numbers to each species and identify the oxidizing and reducing agents.

3. Write the net ionic equation for this reaction.

Lab Questions (answer these as you conduct the investigation):

4. Determine which reactant is limiting and excess, showing all calculations. Explain why this information is useful when determining the enthalpy of reaction.

5. Should you include the mass of both the acid and the magnesium when using $q = mc\Delta t$? Justify your response using data from your experimental design.

6. Calculate the enthalpy of reaction (ΔH_{rxn}) in kJ/mol for each trial showing all work.

7. Calculate the percentage error for your values.

Post Lab Questions:

8. The calorimeter itself will absorb some heat and become part of the system. How could you correct for this source variable? Propose an experiment that could be used to calculate the amount of energy absorbed by the calorimeter.

9. List three potential systematic errors that could be made in this experiment.

Please self-assess your lab report using the checklist/rubric.