LEARNING OBJECTIVES

• To use a simple coffee-cup calorimeter to determine the enthalpy of formation of magnesium oxide.

• To employ Hess’s Law in the determination.

BACKGROUND

The heat (energy) absorbed in the formation of one mole of a compound in its standard state from its elements in their standard states is referred to as the standard enthalpy of formation ($\Delta H$) of the compound.

For example, the standard enthalpy of formation of magnesium oxide is the energy transferred in the formation of one mole of magnesium oxide from its elements in their standard states.

$$\text{Mg}(s) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{MgO}(s)$$

The enthalpy of the above reaction can be determined by measuring the energy transferred when a known quantity of magnesium metal is burned in the presence of excess oxygen. Direct measurement requires the use of complex equipment and is therefore not practical in the first year laboratory.

This difficulty will be circumvented by using Hess’s Law, which states that if a reaction can be regarded as the sum of two or more reactions, the change in enthalpy for the overall reaction will be equal to the sum of the enthalpy changes for the component reactions.

Consider the following reactions:

$$\text{Mg}(s) + 2 \text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{H}_2(g), \quad \Delta H_1 \quad (1)$$

$$\text{Mg}^{2+}(s) + \text{H}_2\text{O}(l) \rightarrow \text{MgO}(s) + 2\text{H}^+(aq), \quad \Delta H_2 \quad (2)$$

$$\frac{1}{2} \text{O}_2(g) + \text{H}_2(g) \rightarrow \text{H}_2\text{O}(l), \quad \Delta H_3 \quad (3)$$
The sum of these three reactions produces:

\[ \text{Mg}(s) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{MgO}(s) \]

Therefore, the desired enthalpy of formation of magnesium oxide is equal to:

\[ \Delta H_{f(Mg)} = \Delta H_1 + \Delta H_2 + \Delta H_3 \]

In this experiment, \( \Delta H_1 \) will be determined by measuring the energy released when a known quantity of magnesium metal is reacted with a slight excess of hydrochloric acid.

\[ \text{Mg}(s) + 2 \text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{H}_2(g) \quad \Delta H_1 \]

The reverse of \( \Delta H_2 \) will be determined by measuring the energy released when magnesium oxide is dissolved in hydrochloric acid.

\[ \text{MgO}(s) + 2 \text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{H}_2\text{O}(l) \quad -\Delta H_2 \]

The heat of formation of water, \( \Delta H_3 \), is readily available in scientific literature (and text books).

Reactions (1) and (2) will be completed in aqueous solution in the lab. An energy transfer will occur between the reacting species and the solution in which the reaction takes place. This energy transfer will cause a resulting increase (or decrease) in the temperature of the solution. The assumption is made that all energy transfer is in the form of heat. Also, assume that the density of the HCl solution is 1.00 g/mL.

The specific heat capacity of water is 4.184 J C\(^{-1}\)g\(^{-1}\). It takes 4.184 J (or 1 calorie) of energy to raise the temperature of one gram of water one degree Celsius (or one degree Kelvin).

By measuring the change in temperature of a given mass of water as a known quantity of material is reacted, it is possible to calculate the energy transferred during the reaction of that material. The energy which would be released by the reaction of one mole of the material (and thus the change in enthalpy for the reaction) can then be calculated.

One of the problems which is encountered in measuring temperature changes is that while the temperature is still rising, energy is being lost to the surroundings (by radiation and convection), and the observed temperature change will be smaller than the “true” temperature change. Completing the reaction under study in a calorimeter system such as the one illustrated below minimizes the energy loss. The loss of energy to the surrounding cannot be completely eliminated, however, and for this reason a graphical temperature correction is employed.

To accommodate for this heat loss, the temperature of the reacting solution is plotted as a function of time. By extrapolating the slope of the temperature curve as the solution cools to time zero, \( T_i \), (the instant of mixing of the reactants) the true maximum temperature, \( T_f \), and thus the true temperature change, \( T_f - T_i \), of the reacting solution can be determined.
Once the true $\Delta T$ is known, the energy transferred during the reaction can be calculated using the specific heat capacity of water, the mass of solution (solid + HCl) in the calorimeter, and the temperature change of the solution:

$$\text{Energy transferred (Joules)} = (4.184 \text{ J} \, ^\circ \text{C}^{-1} \, \text{g}^{-1}) \times \text{(g of solution)} \times (\Delta T)$$

**SAFETY PRECAUTIONS**

- Wear safety goggles at all times in the laboratory.

- Although the hydrochloric acid used in this experiment is quite dilute, burns can occur if spilled and allowed to concentrate upon evaporation. Wash immediately if spilled.
EXPERIMENTAL PROCEDURE

A device that is used to measure the energy evolved or absorbed during a process is called a calorimeter. In this experiment a calorimeter will be constructed from two Styrofoam cups in a fashion similar to that shown above with another serving as a cover. Stirring can be done occasionally by gently swirling the calorimeter.

The initial temperature should be determined for all trials by measuring the temperature of the HCl.

Part A. Determination of $\Delta H_1$

1. Assemble the calorimeter as shown by the instructor. Leave the top and thermometer off the calorimeter at this point.
2. Measure 0.550-0.600 g of magnesium turnings directly into the calorimeter.
3. Measure 60 mL of 1 M HCl into a graduated cylinder and record the volume to the nearest 0.1 mL.
4. Dry the thermometer and place it in the acid solution and record the initial temperature. You should set up an EXCEL spreadsheet with an active graph.
5. Place the thermometer in the hole in the calorimeter cover.
6. Quickly transfer the HCl into the calorimeter and replace the cover. It will fizz vigorously, so be careful.
7. Collect temperature data every 5 seconds for at least 5 minutes.
8. When the procedure above is complete, clean and dry the Styrofoam cup and repeat the procedure for a second trial.

Part B. Determination of $\Delta H_2$

1. Using the same procedure as outlined above, determine $\Delta H_2$.
2. You should use 0.900-1.000 g of MgO and 60 mL of 1 M HCl.
3. Two trials should be completed. Again, collect temperature data for at least 5 minutes.
4. The calorimeter should be cleaned and dried between each trial.

Data Analysis

1. A temperature vs. time curve must be plotted for each trial. Extrapolation to time zero must be done in order to determine the final temperature and, thus, the $\Delta T$.
2. Calculate $\Delta H_1$ (kJ/mol) for each of the two trials in Part A. Calculate the mean $\Delta H_1$.
3. Similarly, plot the data from Part B and calculate the mean $\Delta H_2$. Note that the reverse of the reaction 2 was actually performed. Adjust the sign appropriately.
4. Calculate the enthalpy of formation of magnesium oxide, $\Delta H_f$, (kJ/mol). In your calculation, show the algebraic summation of component reactions and enthalpy values.
THE ENTHALPY OF FORMATION OF MAGNESIUM OXIDE

DATA/RESULTS

A. Determination of $\Delta H_1$

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<thead>
<tr>
<th>Quantity</th>
<th>Trial 1</th>
<th>Trial 2</th>
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<tbody>
<tr>
<td>Mass of Mg (g)</td>
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<td>Moles of Mg (mol)</td>
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<tr>
<td>Mass of solution + Mg (g)</td>
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<tr>
<td>$\Delta T$ (°C)</td>
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<tr>
<td>Energy transferred (kJ)</td>
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<tr>
<td>$\Delta H_1$ (kJ/mol)</td>
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B. Determination of $-\Delta H_2$

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<thead>
<tr>
<th>Quantity</th>
<th>Trial 1</th>
<th>Trial 2</th>
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<tbody>
<tr>
<td>Mass of MgO (g)</td>
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<td>Moles of MgO (mol)</td>
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<td>Mass of solution + MgO (g)</td>
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<td>Energy transferred (kJ)</td>
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<td>$\Delta H_2$ (kJ/mol)</td>
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<tr>
<td>$-\Delta H_2$ (kJ/mol)</td>
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Determination of $\Delta H'_r$ of MgO

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<tbody>
<tr>
<td>Mean $\Delta H_1$ (Part A)</td>
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<tr>
<td>Mean $\Delta H_2$ (Part B)</td>
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<tr>
<td>$\Delta H'_r$ (from literature)</td>
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<tr>
<td>$\Delta H'_r$, MgO (s) (experimental)</td>
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<tr>
<td>$\Delta H'_r$, MgO (s) (from literature)</td>
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<tr>
<td>% error (of $\Delta H'_r$, MgO (s))</td>
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THE ENTHALPY OF FORMATION OF MAGNESIUM OXIDE

PRE-LABORATORY QUESTIONS

1. In a Styrofoam calorimeter, 50.0 mL of 0.100 M AgNO₃ and 50.0 mL of 0.100 M HCl are mixed. The following reaction occurs:

   \[ \text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl (s)} \]

   If the two solutions were initially at 22.60°C and the final temperature is 23.40°C, calculate \( \Delta H \) per mole of \( \text{Ag}^+ (aq) \) for the above reaction. Assume there is 100.0 g of combined solution with a specific heat capacity of 4.184 J°C⁻¹ g⁻¹.

2. In an experiment similar to the procedure described in Part B of the lab procedure, 2.500 g of MgO were combined with 125 mL of 1.0 M HCl. The temperature increased by 9.6°C. Calculate the enthalpy of reaction per mole of MgO for:

   \[ \text{Mg}^{2+} (aq) + \text{H}_2\text{O (l)} \rightarrow \text{MgO (s)} + 2 \text{H}^+ (aq) \]