

## Hess' Law: Determining $\Delta H_f$ for Magnesium Oxide

### 12.1 Background

In a previous experiment, you used calorimetry to experimentally determine the enthalpy of reaction,  $\Delta H_{\text{rxn}}$ , for neutralization. However, for many reactions, such as those involving gases, or that occur particularly rapidly or slowly (or explosively!), a direct measurement is not feasible. For these reactions, it may be easier to utilize Hess' Law.

In this laboratory, you will use calorimetry and Hess' Law to determine the heat of formation,  $\Delta H_f$ , for  $\text{MgO(s)}$ .

Recall that  $\Delta H_f$  for a substance is defined as the heat released or absorbed when one mole of a substance is produced from its elements in their standard states at 1 atm pressure and 25 °C. For  $\text{MgO(s)}$  the formation reaction is:



As this involves a solid and gas, it cannot be determined directly. Therefore, we will make use of Hess' law in order to determine the value of  $\Delta H_f$ . by measuring  $\Delta H_{\text{rxn}}$  for two related reactions and use these values, along with the  $\Delta H_f$  for  $\text{H}_2\text{O(l)}$  to calculate  $\Delta H_f$  for  $\text{MgO(s)}$ .

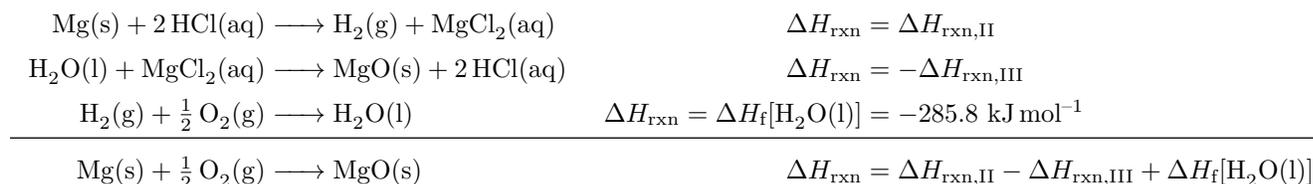
The first reaction that you will study is the reaction between  $\text{Mg(s)}$  and  $\text{HCl(aq)}$ :



The second reaction involves the reaction between  $\text{MgO(s)}$  and  $\text{HCl(aq)}$ :



According to Hess' law, heats of reactions are additive just as are the balanced chemical equations representing those reactions. If we combine the above two reactions appropriately, along with the formation reaction for  $\text{H}_2\text{O(l)}$ , we will obtain the formation reaction for  $\text{MgO(s)}$ , and the  $\Delta H_f$  for  $\text{MgO(s)}$ .



To determine  $\Delta H_{\text{rxn}}$  for the reactions between  $\text{Mg(s)}$  and  $\text{HCl}$ , and  $\text{MgO(s)}$  and  $\text{HCl}$ , we will use the experimental and mathematical techniques we developed last week in lab. As a reminder,

$$\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{\text{moles of limiting reactant}} \quad (12.1)$$

$$q_{\text{rxn}} = -q_{\text{surr}} \quad (12.2)$$

and

$$q = ms\Delta T \quad (12.3)$$

The calculation of the moles of limiting reactant in each case should be straightforward—it will simply be the number of moles of Mg or MgO added to the solution. The calculation of  $q$ , however, is slightly more complicated than in last week's lab. Specifically, recall that you need to know the mass of your solution. Last week, we were able to calculate the mass through the density of the solution that was provided to you. This week, however, not all the reactants are aqueous. Therefore, we must recognize that

$$m_{\text{total}} = m_{\text{HCl solution}} + m_{\text{solid reactant}} \quad (12.4)$$

(We will assume that the mass lost due to H<sub>2</sub> gas escaping in Reaction (II) is negligible.)

At this point, you should be able to calculate the mass of the HCl solution used, based on the density and specific heat provided in the Calculations section below. The mass of the solid reactant (Mg or MgO, depending on the reaction) will simply be the amount you measured on the balance. Once you account for this, the calculations are just like those you did last week.

After you determine the enthalpy of formation for MgO, you can look up the literature ("known") value in a textbook and then calculate the percent error to see how far off you are.

$$\% \text{ error} = \left| \frac{\text{theoretical value} - \text{experimental value}}{\text{theoretical value}} \right| \times 100\% \quad (12.5)$$

There are a few things to notice about percent error. First, the absolute value sign means that % error is always positive. In other words, while it tells you how far off your experimental value is from the theoretical value, it doesn't reflect whether your value is too high or too low. Second, notice that if your experimental value exactly matches the theoretical value, the % error will come out to zero. In other words, a lower value is better. What constitutes a "good" percent error, however, will depend on the type of experiment being done. In some labs, such as titrations, a 5% error is unacceptable. In others, such as determining the freezing point depression due to a solute, it is quite good.

## 12.2 Procedure

Two styrofoam cups, one nested inside the other, and a plastic cover will serve as your calorimeter. If there is not already a small hole in the center of the plastic cover, make one with a pen cap. Assemble the calorimeter by slipping the plastic cup cover over the temperature probe and slide it up the temperature probe. Stirring will be accomplished by swirling the cup gently without bumping the temperature probe.

1. Using the analytical balance, accurately measure out about 0.6 g of magnesium turnings and record the mass. Set the Mg(s) aside.
2. Add 60 mL of 2 M HCl (**CAUTION! STRONG ACID!**) to the calorimeter. Record the volume to the nearest 0.1 mL.
3. Start LoggerPro on the laptop and connect the GoTemp interface to one of the USB ports on the laptop. The temperature should be displayed on the screen.

4. Set up data collection in LoggerPro as follows:
  - a) Under the **Experiment** menu, select **Data Collection**.
  - b) Choose **Time Based** for mode.
  - c) Type in **480 seconds** for Duration.
  - d) Type in 1 second per sample.
  - e) Press **Done**.
5. Hit **Collect** on the computer to start data collection and *swirl the cup vigorously and constantly*.
6. Once you have established a stable start temperature, lift the cover on the calorimeter and quickly add the magnesium. Continue to swirl vigorously and constantly. This will provide the data for the determination of  $\Delta H_{\text{rxn,II}}$ , Reaction (II).
7. When you have collected all the data, clean up the graph and analyze the data just like you did in the previous lab.
8. Copy/paste your graph into Word.
9. Dispose of your solution in the **Recovered Acids and Bases** container.

You will do a similar experiment to determine a value for  $\Delta H_{\text{rxn,III}}$ , Reaction (III).

10. To clear the data from reaction 1 in LoggerPro, go to the **Experiment** menu and click **Clear Latest Run**. Set up data collection in LoggerPro the same way you did previously.
11. Measure out (to the same precision as before) about 1 g of MgO. Repeat the procedure for Reaction (II), (**CAUTION! STRONG ACID!**) *Swirl the mixture vigorously and constantly*.
12. When you have collected all of the data, dispose of your solution in the **Recovered Acids and Bases** container. Perform the same scaling, labeling, and extrapolation on the graph as you did for the graph of the Reaction (II) data. Resize the graphs in Word so that both graphs will print on a single page, and print a copy for each individual.

## 12.3 Calculations and Analysis

1. Determine  $\Delta H_{\text{rxn}}$  for Reactions (II) and (III), remembering that Mg and MgO are your limiting reagents.
2. Use Hess' Law to determine  $\Delta H_f$  for MgO.
3. Look up the literature value for  $\Delta H_f^\circ[\text{MgO(s)}]$  and calculate your % error.

You may find the following values useful:

2 M HCl: specific heat: 4.046 J/(g °C); density: 1.043 g/mL

The heat of formation for  $\text{H}_2\text{O(l)}$  is  $-285.8 \text{ kJ mol}^{-1}$ .

### References and Resources

Refer to Section 9.8 of your textbook for more background on the topics covered in this experiment.