Experiment 10
Thermochemistry

OUTCOMES

After completing this experiment, the student should be able to:

- measure the enthalpy of a reaction in the laboratory using temperature data.
- use Hess's Law to estimate the enthalpy change for a reaction.
- write an introduction for a lab report.

DISCUSSION

There are times in the lab when we want to know how much heat is given off or absorbed during a reaction. Sometimes this can be determined quite easily experimentally by measuring temperature changes. However, there will be instances when this is not a straightforward task. For example, in the reactions lab, you observed the burning of magnesium ribbon in oxygen to give magnesium oxide while emitting a bright light.

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

The way in which the experiment was carried out makes it difficult to determine the enthalpy change. How do you measure the temperature change when the heat dissipates in the air quite rapidly? How do you accurately separate the magnesium oxide from the unburned magnesium to determine how much reacted?

In cases like this where we do not have an easy reaction to measure directly, we often use Hess's Law of Heat Summation in order to predict the enthalpy change. When burning magnesium, the reaction can be expressed as the sum of three other reactions:

a) $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ ΔH_a = class average

b) $MgO(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$ ΔH_b = class average

c) $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$ $\Delta H_c = -571.6$ kJ

The first two reactions, (a) and (b), both take place in solution. Thus, it should not be too difficult to measure temperature changes for the reactions and relate them to the enthalpy change, knowing that the enthalpy change for a reaction at constant pressure is equal to the heat of the reaction ($\Delta H_{rxn} = q_{rxn}$). Because the heat gained by each reaction is lost from the surroundings and vice versa, we know that $q_{rxn} = -q_{surroundings}$. Here, the surroundings will be the calorimeter *and* the solution.

There are two different equations that we can use for q: $q = C*\Delta T$ for large objects (like the calorimeter) where the mass does not change and $q = m*s*\Delta T$ for components of a reaction that are variable. In this experiment, the heat capacity (C) of the coffee cup calorimeter is given as 10. J/°C. For the solution, you will use the combined mass of *all* reagents (solids and solutions) for the mass, and the specific heat (s) of the solution will be estimated to be the same as that of water, or 4.18 J/(g°C). In both cases, the change in temperature (ΔT) is determined by $T_{max} - T_{min}$.

In order to make accurate comparisons, the q_{rxn} (and thus ΔH) determined for reactions (a) and (b) will need to be converted to a change in enthalpy per mole of limiting reactant (kJ/mol). Once you have a value of ΔH for one mole of limiting reactant, you will need to convert it to represent the number of moles of limiting reactant in the balanced equation.

MATERIALS

calorimeter (two Styrofoam cups, nested)
cardboard cover with a small notch cutout
3 strips of magnesium metal (~0.15 g each)
150 mL 1.0 *M* HCl
Logger Pro with LabPro® Interface
temperature probe
magnesium oxide (~0.75 g)

PROCEDURE

⚠ ***Wear safety glasses or goggles at all times for this experiment.***
⚠ ***Avoid skin contact with the chemicals in this experiment.***
⚠ ***Never pipet by mouth.***

You will work in the same groups of 3 to 4 students as were assigned in the magnesium/hydrochloric acid lab. Students who served as a manager or computer operator will work as lab techs for this activity. In the event you had only 3 students the first time, or if a group member has dropped since then, you may make your own assignments, but you should take on a different role than last time. Thus, each individual should have had a chance to do some of the dirty work, as well as some of the greater responsibility of getting the assignment submitted to the professor in a timely manner. To refresh your memory on the role played by each person, you may check out the group work page.

1. Before starting the experiment, develop a hypothesis for what you think will happen in the lab. You will be writing an introduction for this experiment and should have a hypothesis to include in it. It could be related to whether the reactions are endothermic or exothermic, which reaction will have a greater change in temperature, or any other question that you think will be answered during the course of the experiment.

2. Plug a stainless steel temperature probe into the LabPro® interface and launch the Logger Pro application. Adjust the experiment length to 200 seconds and set the sampling rate to 5 seconds per sample. Adjust the number of decimal places for the temperature in your data table to the nearest ±0.1 °C.

3. Place your calorimeter onto an electronic balance (use a cheap, less precise balance for this measurement, as you do not want to risk spilling acid on the analytical balance) and tare the balance. Remove the calorimeter from the balance and carefully add 25.0 mL of 1.0 *M* HCl. Place the cup back onto the balance and record the mass of HCl added. Also, record the volume of HCl used. These may be recorded into your lab notebook and into the sample lab report page provided for the lab.

4. Measure and record the mass of a magnesium strip (~0.15 g), using the more precise analytical balance. Roll the magnesium ribbon into a loose ball.

5. Slide the temperature probe into the small notch cutout on the cardboard cover and place the probe into the HCl. Stir the HCl with the probe to maintain a uniform temperature throughout the solution. Wait until the temperature stabilizes.

6. Click the green "Collect" button to begin data collection. After a couple of data points have been collected to establish the starting temperature, slide the cover aside and drop the ball of magnesium into the calorimeter. Slide the cover back into place. Continue stirring until the data collection ends.

7. Select *Store Latest Run* from the *Experiment* menu and save your data to your M:drive.

8. Repeat steps 2 – 6 two more times to obtain a total of three trials of the magnesium reaction with hydrochloric acid.

9. Perform steps 2 – 6 three times using ~0.25 g of magnesium oxide in place of the magnesium strips.

10. If your data looks good, copy and paste the data from each of your trials into an Excel spreadsheet. Use the appropriate function or formula to determine the minimum and maximum temperatures reached in each trial.

⚠ ***Dispose of all chemicals in the proper waste container.***

DATA ANALYSIS

1. Determine which substance is the limiting reactant in the reaction of magnesium with hydrochloric acid and in the reaction of magnesium oxide with hydrochloric acid. How

many moles of limiting reactant were used in each of the reactions? Provide at least one sample calculation for each of the two types of reactions. Remember, you may place formulas in spreadsheet cells for the sample calculations if you submit your postlab report electronically.

2. Calculate the enthalpy change (ΔH) per mole of the limiting reactant, in kJ/mol, for each of the two reactions. Provide at least one sample calculation for each of the two types of reactions. Refer to the lab discussion for helpful information and equations for these calculations.

3. Using the coefficients of the limiting reactant in each of the equations, determine ΔH, in kJ, for the balanced equations in (a) and (b). Calculate the average ΔH value for each of the two reactions and submit your average ΔH values here before leaving the lab (or within 24 hours with instructor approval).

4. Use Hess's Law of Heat Summation, the class average enthalpy changes for reactions (a) and (b), and the given enthalpy change for reaction (c) to determine the enthalpy change for the reaction: $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$.

5. Calculate the percent error in the ΔH that you calculated in step 4, assuming the reaction was carried out under standard thermodynamic conditions. In this case, your answer for step 4 is the experimental value of ΔH. You will need to calculate a theoretical value of ΔH using $\Delta H_f°$ values from your textbook.

$$\text{percent error} = \frac{|\text{experimental} - \text{theoretical}|}{\text{theoretical}} \times 100$$

6. What would happen to the value of ΔH that you calculated if all of your temperature readings were too high by 1 °C?

7. What were some possible sources of error in this experiment? Explain.

POSTLAB ACTIVITY

You will be turning in a group lab report. The report should include the title, an introduction, results, discussion, and references. The information that you obtained from the data analysis should be included at some point in the report. It is up to you whether it is in the results or discussion or both. However, remember that the report is more than just answering some questions and that it should flow smoothly and logically as you discuss the data obtained and what it signifies. Lab report guidelines for how to write the introduction, results, and discussion are found at http://webs.anokaramsey.edu/chemistry/Chem1061.

Follow your instructor's directions for submitting the report. If you are submitting electronically, please submit a single file with all of the required information. Use the following convention for naming your files: *Lastname1 Lastname2 Etc Thermo*. If you are emailing the report, use a subject line of *Chem 1061: Thermochemistry Lab*.

As indicated previously, you will need to show sample calculations in the report. For electronic submissions, you may embed data tables which contain the formulas in calculated cells. For paper submissions, you will need to show these calculations for one trial of each reaction. You will also need to show these calculations if you submit the report electronically but do NOT include formulas in embedded data tables.