experiment 11

Enthalpy of Reaction

LECTURE AND LAB SKILLS EMPHASIZED

- Calculating the heat and enthalpy of reactions
- Writing net ionic equations
- Using Hess's law to determine the enthalpy of a reaction
- Defining calorimetry

IN THE LAB

- Students will work in pairs.
- Record your procedure and original data in your lab notebook along with your calculations.
- Report data collected and the subsequent calculations to **www.chem21labs.com**.
- *All equipment should be returned to the correct location after use.*



WASTE

- All solutions should be disposed of in the acid–base waste containment.
- *Styrofoam cups should be rinsed, reused, and returned at the end of the lab session. Do not throw away.*

SAFETY

- Concentrated acids and bases can cause serious burns. Use with caution.
- Gloves and safety goggles are mandatory when anyone is performing an experiment in the lab.
- Long pants, closed–toed shoes, and shirts with sleeves. Clothing is expected to reduce the exposure of bare skin to potential chemical splashes.
- Always wash your hands before leaving the laboratory.

Additional information can be found at: **https://wp.as.uky.edu/genchem/home/che-113/11-enthalpy/**

To continue to show your analytical ability, your company has asked you to determine the enthalpy for a chemical reaction using three different methods. Remember, you will need to identify these three methods and be careful of your laboratory technique to ensure your results come out the same.

All chemical reactions involve an exchange of heat energy; therefore, it is tempting to plan to follow a reaction by measuring the enthalpy change (ΔH). However, it is often not possible to directly measure the heat energy change of the reactants and products (the system). We can measure the heat change that occurs in the surroundings by monitoring temperature changes. A **calorimeter** is a device used to measure the heat changes when two or more substances are combined, whether it be two substances at different temperatures or two substances that will undergo a chemical reaction. The key feature of a calorimeter is that it is insulated enough to minimize the loss of heat to the surroundings. A Styrofoam cup is an effective calorimeter because it acts as a good insulator. Inserting the calorimeter into a beaker provides extra insulation and stability to the cup. Ideally, a calorimeter would not absorb any heat, but in reality it does. By performing a calibration of the calorimeter, we will be able to determine the amount of heat absorbed by the calorimeter and will account for that loss when determining the heat change for our chemical reactions.

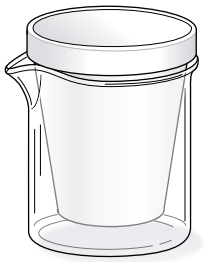

Figure 11.1. Styrofoam cup calorimeter.

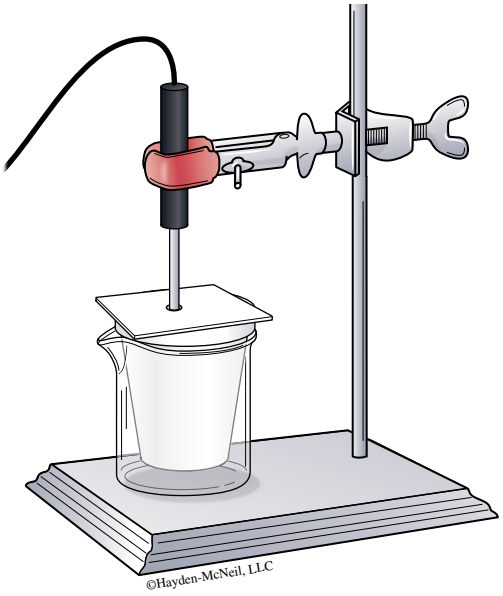

Figure 11.2. Setup for measuring temperature change during a reaction.

In this experiment, you will design a procedure to determine the enthalpy of three acid–base reactions. To do this, you will need to determine the heat capacity for the calorimeter by mixing samples of hot and cold water. Then, based on the differences in the temperature of the two water samples, determine the amount of heat lost to the calorimeter.

If we assume no heat is lost to the surroundings, then the following expression is true:

$$q_{\text{cold water}} + q_{\text{hot water}} + q_{\text{cal}} = 0 \qquad (1)$$

and this can be rearranged to the following,

$$q_{\text{cal}} = -q_{\text{hw}} - q_{\text{cw}} \qquad (2)$$

which will allow us to determine the heat of the calorimeter. There are two formulas which we will use to calculate the individual terms in the equation. The first formula, which is shown in equation 3, is used when you have varying masses of a substance but you know the specific heat of the substance. The **specific heat**, s, has been determined for many substances (i.e., water, iron, etc.) and can be found in a variety of reference sources.

$$q = ms\Delta T \qquad (3)$$

The second formula, shown in equation 4, is used for determining the heat when you are dealing with an object with a fixed mass whose composition will not change, such as a calorimeter. The value of the **heat capacity**, C, is unique to a particular item and must be determined experimentally for that object.

$$q = C\Delta T \qquad (4)$$

These formulas can be substituted into equation 3 to yield the following equation,

$$C_{\text{cal}}\Delta T_{cal} = -(m_{\text{cw}})(s_{\text{water}})(\Delta T_{\text{cw}}) - (m_{\text{hw}})(s_{\text{water}})(\Delta T_{\text{hw}}) \qquad (5)$$

where

ΔT_{cal} = change in temperature of the calorimeter (equals ΔT_{cw} because they both start and end at the same temperature)

m_{cw} = mass of the cold water

s_{water} = specific heat of water

ΔT_{cw} = change in temperature of the cold water

m_{hw} = mass of the hot water

ΔT_{hw} = change in temperature of the hot water

We can expand equation 5 further to the following

$$C_{cal}(T_f - T_{i,cw}) = -(m_{cw})(s_{water})(T_f - T_{i,cw}) - (m_{hw})(s_{water})(T_f - T_{i,hw}) \quad (6)$$

Note that T_f values will all be the same because the cold water, hot water, and calorimeter reach the same temperature. When we look at equation 6, we can find all the values either from experimental data or reference values except C_{cal}. The calculated result will serve as our experimental value for C_{cal}, which will be used when we determine the heat of reaction for three chemical reactions using the same calorimeter.

If we conduct a reaction between two substances in an aqueous solution, any heat gained or lost by the chemical reaction will be absorbed by either the solution or by the calorimeter. By measuring the temperature change as the reaction occurs, we can determine the amount of heat absorbed by the solution and the amount of heat absorbed by the calorimeter. The sum of those values will equal the amount of heat change from the actual chemical reaction. If we assume that no heat is lost from the system, then the following is true (similar to equation 1):

$$q_{rxn} + q_{soln} + q_{cal} = 0 \quad (7)$$

and this can be rearranged into the following form,

$$q_{rxn} = -q_{soln} - q_{cal} \quad (8)$$

which will allow us to determine the heat of the reaction by calculating the heat of the solution and calorimeter. As before, we can expand this equation to obtain

$$q_{rxn} = -m_{soln} s \Delta T - C_{cal} \Delta T \quad (9)$$

where m_{soln} is the total mass of the combined solution, s is the specific heat of water since we are using aqueous solutions, and ΔT is the change in temperature of the solution and the calorimeter during the reaction. The initial temperatures of the solutions, as well as the calorimeter, should be the same since they are all starting at room temperature. The final temperature will be the maximum temperature reached by the solution which will also be the final temperature of the calorimeter. Using the heat capacity determined in the calibration process along with the other experimental data will allow us to calculate the heat of the reaction.

While the heat can be readily calculated based on experimental data, the values are only relevant when using identical amounts of reagents with the same calorimeter. A more universal value is that of enthalpy which is reported in kJ/mole of a particular reactant, which in this case will be the limiting reagent. To determine the enthalpy, one can use the following:

$$\Delta H = \frac{q}{n} \quad (10)$$

where q is the heat change of the reaction and n is the number of moles in the limiting reagent. The enthalpy values can then be compared to the same reaction performed with different amounts of a substance or to accepted enthalpy values.

In this experiment, we will look at three reactions as outlined in the pre-lab exercise. For each of these reactions, we will determine the heat of the reaction by looking at the change in temperature of the solution as well as using the heat capacity determined for our calorimeter. The values determined in the pre-lab exercise will serve as your accepted values for determining your percent error. We will also get the opportunity to use Hess's law (see your textbook for review of this concept) which will let us determine two different enthalpy values for the same reaction which can be compared to the accepted value.

Tips for Procedure

- Instructions on using the temperature probe are available in Chapter 4.
- Make sure the MeasureNet station will graph and display limits of y = 0 to 100, x = 0 to 600.

- Record the exact masses of your hot and cold water as well as their containers. You will need to do this by measuring the mass of the Styrofoam cup, then the mass of cup + cold water. Repeat the process with the hot water and a 100 mL beaker. The hot water should be between 35 and 40°C.
- Use approximately 25 mL of each solution. Record the exact volume of all solutions used in your lab notebook.
- Always add the base to the acid solution.
- Record all initial and final temperatures for every solution.
- Read the labels carefully, as some solutions may appear to be similar.
- Solutions can be carefully mixed using the temperature probe, as long as you do not remove the probe from the solution.
- Save the data file from your calibration as "010," then your reactions in order as "001," "002," and "003."

Pre-Lab Exercise

You will conduct the following three reactions in this experiment. On Achieve, write the balanced net ionic reaction equations from the descriptions. Use the table of thermodynamic data in your general chemistry textbook to calculate the molar enthalpy of the reactions.

Reaction 1: An aqueous solution of sodium hydroxide reacts with an aqueous solution of hydrochloric acid, yielding water and aqueous sodium chloride.

Reaction 2: An aqueous solution of sodium hydroxide reacts with an aqueous solution of ammonium chloride, yielding aqueous ammonia (NH_3), water, and aqueous sodium chloride.

Reaction 3: An aqueous solution of hydrochloric acid reacts with aqueous ammonia, NH_3, yielding aqueous ammonium chloride.

Procedure

You must write the procedure and prepare your lab notebook before coming to lab to do the experiment. Make sure that you will collect the data necessary to complete the data analysis questions. It is better to have too much information and not need it, than to need something and not have it.

Details about collecting data with MeasureNet are in Chapter 4.

Suggested Materials

temperature probe

Styrofoam cup calorimeter

250 mL beaker

50 mL graduated cylinder

glass stirring rod

2.0 M hydrochloric acid solution

2.0 M sodium hydroxide solution

2.0 M ammonium chloride solution

2.0 M ammonium hydroxide solution (NH_3/NH_4OH)

ring stand

iron clamp

cardboard square

beakers, assorted sizes

Data Analysis

1. What was the mass and value of ΔT for the calorimeter?
2. What was the mass and value of ΔT for the cold water?
3. What was the mass and value of ΔT for the hot water?
4. Determine the heat change (q) for the cold and hot water.
5. Calculate the heat change for the calorimeter (pay attention to the sign on the value of q).
6. Using ΔT for the calorimeter, determine the heat capacity of the calorimeter.
7. Calculate the heat change for each of the three reactions. Use 1.03 g/mL for the density of all solutions.
8. For each reactant, determine the limiting reagent (all reactions have a 1:1 molar ratio) and use the amount of that reactant, in moles, to determine the enthalpy change for each of the three reactions.
9. Use the experimental data for reactions 1 and 2 along with Hess's law to determine the enthalpy of Reaction III.
10. Using the two enthalpy values calculated previously (questions 8 and 9), determine the percent error in each of these values compared to the accepted value calculated in the pre-lab exercise.