

# Thermochemistry

## Introduction

Welcome to the Chemistry 2B Laboratory. During this first laboratory period you will go over the laboratory safety rules, become acquainted with the layout and equipment in the laboratory, and check-out the equipment in your locker. Then you will begin the first experiment of the quarter, which involves one of the most important areas of science, thermodynamics.

## Safety Rules and the Laboratory Room

After reviewing the safety rules with your TA, sign the back of the safety sheet and return it to your TA. Remember to always follow the safety instructions when performing all experiments! **WEAR YOUR GOGGLES!**

## Checking Out Your Locker

Checking out your locker is a process in which you make sure your locker contains all of the proper equipment in the correct quantities. This is sometimes difficult, as some of the equipment have rather complicated names. Please turn to pages 4-5 of this manual for a locker list and drawings of common laboratory equipment. If you are missing any items, first check the box of extra glassware that is located at the back of the laboratory. If you still cannot find the missing equipment, visit the stockroom (Room 71). They will give you the glassware that you are missing. Please replace all missing equipment the first day of laboratory since the stockroom is only prepared to replace glassware during the first week and the process of replacing glassware later in the quarter is more complicated. If you have extra glassware, please place it in the box as someone else in the laboratory is probably missing the item.

## Thermochemistry Experiment

This experiment is an introduction to the basic principles of thermochemistry and involves the exchange of energy as heat. The ideas and concepts involved in thermodynamics are illustrated in your everyday experiences. For example, on a hot summer day the hood of a car can get hotter than the sidewalk cement and when cooking, you have probably noticed that a wooden spoon does not heat as fast as metal one. After completing this experiment, you will better understand the reasons behind these and other thermal phenomena.

In the first part of this experiment you will construct a simple "coffee-cup" calorimeter. Please note that while this calorimeter is not expensive, when used properly it can give very good results. In the next part of the experiment you will measure the specific heat of an unknown solid. You must carefully follow the procedure outlined to avoid the experimental problems described. In the third and fourth parts of the experiment you will determine the enthalpies,  $\Delta H_{\text{rxn}}$ , of endothermic and exothermic reactions. You will be exploring the factors that cause a reaction to occur. In the fifth

part of the experiment, you will design your own procedure to determine the heat of fusion,  $\Delta H_{\text{fus}}$ , of water.

In order to make sense of your observations for the third and fourth parts of the experiment you will need to think about an additional concept. It is fairly easy to rationalize why an exothermic reaction occurs; the reaction releases heat, implying that the products are of lower energy than the reactants ( $\Delta H_{\text{rxn}}$  is negative). However, in an endothermic reaction heat is absorbed, indicating that the products are higher in energy ( $\Delta H_{\text{rxn}}$  is positive). What provides the driving force for an endothermic reaction?

The answer to this question is entropy, symbolized "S". Entropy will be fully discussed later in the course, but a brief introduction is provided here. Entropy can be thought of as a measure of the disorder or randomness in a system; the greater the disorder the higher the entropy. For instance, compare your lecture at the beginning of the hour and in the middle of the hour. At the beginning of the lecture everyone is coming into the room and milling around, finding seats and getting settled. Entropy is high. In the middle of the lecture everyone is seated in rows of chairs, all quietly facing the same direction with their attention focused in pretty much the same place. Entropy is low.

When entropy is discussed in chemistry, attention is focused on the number and motion of particles in a system. A reaction that results in an increase in the total moles of particles ( $n_f - n_i > 0$ ) is said to have an increase in entropy ( $\Delta S > 0$ ). Entropy also depends in part upon particles' distribution in space and in part on the distribution of energy (and motion) among the particles. The more freedom particles have to move around the more entropy they will have. Changing a specific sample of a solid to a liquid does not increase the number of moles of the sample, but the energy and motion of the molecules definitely increases. Therefore, the entropy of the sample has increased. Changing the liquid to a gas dramatically increases the entropy of the system. Similarly dissolving a salt in water will increase the entropy because the particles go from a very organized crystal to a less organized solution of free-moving ions.

Nature tends to **minimize** enthalpy ( $\Delta H$ ) and **maximize** entropy ( $\Delta S$ ). Entropy can therefore be a driving force for a reaction since greater entropy is a preferred condition. Endothermic reactions occur because entropy is increasing. The gain from increasing entropy ( $+\Delta S$ ) in these reactions is enough to counterbalance the unfavorable enthalpic conditions ( $+\Delta H$ ). Increasing the entropy of a system has the same effect as minimizing the enthalpy of the system --- it drives the reaction forward. In the third and fourth parts of this experiment you will look at reactions that vary in their enthalpic and entropic properties. As you will see, one of the reactions is enthalpy-favored ( $-\Delta H$ ) but not entropy-favored ( $-\Delta S$ ), one is entropy-favored ( $+\Delta S$ ) but not enthalpy-favored ( $+\Delta H$ ), and one is favored by both enthalpy ( $-\Delta H$ ) and entropy ( $+\Delta S$ ).

Finally, in the last part of this experiment you will design your own procedure to determine the heat of fusion for ice,  $\Delta H_{\text{fus}}$ . Please note that you must come to the laboratory with an outline of the procedure you plan to use. **As preparation for this experiment you should read Chapter 7 in Petrucci et.al., 8th ed.**

### **Background: Heat, Specific Heat, Heat Capacity, and Molar Heat Capacity**

All parts of this experiment require the use of a calorimeter. In the first part of this experiment, you will construct an inexpensive but effective coffee-cup calorimeter. Before you can use this calorimeter to determine thermodynamic quantities you must determine the **heat capacity** of the calorimeter itself. You will do this by adding a weighed sample of hot water to a known amount of cold water in the calorimeter and measuring the temperature change.

When measuring energy transfers, it is necessary to know how much energy is required to cause the temperature of a sample to change. Think about a summer day. The same sun is shining down on your car, the patio and the pool, but by afternoon, you cannot touch the hood of the car, you have to tread carefully on the patio if you are in bare feet, and the pool is a nice place to cool down. All three of these things have received the same amount of energy, but they are at different temperatures.

The amount of energy required to change the temperature of an object or a sample of a substance by one degree Celsius or Kelvin is called that object's

**heat capacity**, symbolized  $c_p \left( \frac{J}{^\circ C} \right)$ .

There are two variations on heat capacity that you also need to be familiar with:

**specific heat**,  $C_{sp} \left( \frac{J}{g^\circ C} \right)$       and      **molar heat capacity**,  $C_p \left( \frac{J}{mol^\circ C} \right)$ .

The specific heat of a substance is the heat required to raise the temperature of one gram of the substance one degree, and the molar heat capacity is the amount of energy required to raise one mole of the substance by one degree. All substances have characteristic specific heats and molar heat capacities.

When two substances having different temperatures come into contact, energy in the form of heat is exchanged between them until they reach a common temperature. If they are insulated from their surroundings, the amount of heat lost from the hotter substance equals the heat gained by the colder one. The heat lost or gained is related to the mass, the specific heat of the substance, and the temperature change. This relationship is expressed as

$$q = m * C_{sp} * \Delta T$$

where  $q$  is the heat,  $m$  is the mass,  $C_{sp}$  is the heat capacity of that substance, and  $\Delta T$  is the change in temperature. This equation can also be used if moles are substituted for mass, and molar heat capacity is substituted for specific heat. In this experiment, you will determine the heat capacity of your calorimeter.

When dealing with the calorimeter itself, you will combine the mass and specific heat of the calorimeter into a single term, the calorimeter's heat capacity. This can be done since the mass of the calorimeter does not change.

$$q = c_{p(\text{calorimeter})} * \Delta T$$

**Safety: To avoid burns use crucible tongs to pick up hot metal. Never pick up a heated metal with your bare hands. Wear gloves and use caution whenever handling acids and bases. All waste from this experiment can be poured down the drain. WEAR YOUR GOGGLES.**

**YOU WILL WORK INDIVIDUALLY ON THIS EXPERIMENT.**

### **Experimental Procedure**

#### **Part I. Determining the Heat Capacity of the Calorimeter**

1. Set up a Bunsen burner (or a hot plate) and heat 500 mL of deionized water to boiling in an 800 mL beaker. (You may share a waterbath only.) You may need to refresh your water supply periodically to prevent the water from boiling away completely.
2. Put two Styrofoam coffee cups in a 250 mL or a 400 mL beaker. Take the top Styrofoam cup from the calorimeter, place it on the balance, and calibrate the balance to zero mass. This is called "taring" the container. Weigh out about 70 grams of room temperature deionized water into the calorimeter and record the mass of the water to the nearest thousandth of a gram. We will refer to this mass as the "mass of the cool water",  $mass^{cool\ water}$ .
3. Place the cup back into the calorimeter set-up. Take a 4"x 4" piece of cardboard, with the hole in the center, and place it on top of the coffee cups. If the hole is too big, cover it with two pieces of masking tape and punch a hole with your pen or pencil big enough so the thermometer can fit. Insert a thermometer through the hole.
4. Using the buret holder, gently clamp the thermometer and lower it into the cup so that the whole bulb is covered with water but is not touching the bottom of the cup. Avoid positioning the calorimeter too close to a Bunsen burner or hot plate so that the water inside the calorimeter remains cool.
5. Place your **large** labeled test tube in a 150 mL beaker and tare this assembly. Put 30 ml of room temperature deionized water in the test tube and record the mass of the water to the nearest thousandth of a gram in your notebook along with the corresponding test tube label. We will refer to this mass as the "mass of hot water." Repeat this step with two more large test tubes.
6. Place one of the labeled test tubes in the beaker of boiling water using a utility clamp to hold it in place. Make sure that the water level in the test tube is below the water level in the 800 mL beaker. Allow the test tube of water to sit in the boiling water for 15 minutes in order for the temperature of water in the test tube to equilibrate to the temperature of the boiling water in the beaker.

7. After the 15 minutes measure the temperature of the boiling water in the beaker with your second thermometer and record to the nearest 0.2 °C. We will refer to this temperature as the initial temperature of the hot water,  $T_i^{\text{hot water}}$ .
8. Just before transferring the hot water in the test tube to the calorimeter, measure the temperature of the water in the calorimeter and record to the nearest 0.2 °C. We will refer to this temperature as the initial temperature of the cool water,  $T_i^{\text{cool water}}$ . Next, remove the thermometer and cardboard top from the calorimeter. Using your test tube clamp, grasp the test tube containing the 30 ml of hot water near the top of the test tube and quickly but carefully pour the hot water into the calorimeter. Be careful that no hot water on the outside of the test tube drips into the calorimeter.
9. Replace the thermometer and cardboard top on the calorimeter. Adjust the thermometer's height so that it is not touching the bottom or sides of the calorimeter yet the water is covering the thermometer bulb.
10. Gently stir the water in the calorimeter until the highest temperature is reached. This is the equilibrium temperature. Watch the thermometer closely as it rises. Sometimes it will rise, fall, and rise again due to the initial uneven distribution of heat within the calorimeter. Stirring the water in the calorimeter distributes the heat throughout the calorimeter. Monitor the temperature and record the highest temperature attained to the nearest 0.2 °C. We will refer to this temperature as the final temperature,  $T_f$ .
11. Place the next test tube in the boiling water and allow it to equilibrate to 100°C for 15 minutes. While you are waiting, repeat steps 2, 3 and 4. Once the water in the test tube has equilibrated complete steps 7, 8, 9, and 10. Dry the calorimeter and thermometer between trials.
12. Repeat the same procedure using your final test tube of water.

## **Part II. Determining the Specific Heat of a Metal**

Using the same calorimeter for which you determined the heat capacity, you will analyze an unknown metal sample to find its characteristic specific heat capacity and identify the sample as **lead**, **aluminum** or **copper**.

1. Continue to heat 500 mL of deionized water to boiling in an 800 mL beaker either using a Bunsen burner or a hotplate. Again, you may need to refresh your water supply to prevent the water from boiling away completely.
2. Set up your calorimeter by placing the two Styrofoam cups in a 250 mL or a 400 mL beaker as before. Take the top Styrofoam cup from the calorimeter, place it on the balance and tare it. Weigh out about 70 grams of room temperature deionized

water into the calorimeter and record the mass of the water to the nearest thousandth of a gram.

- Place the cup back into the calorimeter set-up. Take your 4"x 4" piece of cardboard, with the hole in the center, and place it on top of the coffee cups. Insert a thermometer through the hole.
- Using the buret holder, gently clamp the thermometer and lower it into the cup so that the whole bulb is covered with water but is not touching the bottom of the cup.
- Obtain a sample of **the unknown metal** from the box at the front of the room. Identify the metal based on density and color and **write down the type of metal** you obtained in your laboratory manual. The sample should be a piece of metal strung with nylon string. Do not remove the nylon string. Weigh the unknown metal sample using a weigh boat to protect it from contamination. Tare the weigh boat. Record the mass of the metal to the nearest thousandth of a gram.
- Suspend the string of metal disks from a utility clamp that is attached to the superstructure at the laboratory bench above the 800 mL beaker of boiling water. Adjust the height of the utility clamp so that the metal is completely submerged in the boiling water. Suspension of the sample insures that the metal will have the same equilibrium temperature as the boiling water by preventing direct heating by the flame or hot plate (which would result if the metal were allowed to rest on the bottom of the beaker). Add more water to the beaker if necessary and return to a boil. Make sure the beaker does not tip over.
- Allow the metal to sit in the boiling water for 3-4 minutes. This will insure the temperature of the metal to be approximately 100 °C.
- After the 3-4 minutes measure the temperature of the boiling water in the beaker with your second thermometer and record to the nearest 0.2 °C. This will be the initial temperature of the metal,  $T_i^{\text{metal}}$ .
- Just before transferring the metal to the calorimeter, measure the temperature of the water in the calorimeter and record to the nearest 0.2 °C. This is the initial temperature of the water,  $T_i^{\text{water}}$ , and calorimeter,  $T_i^{\text{calorimeter}}$ . Next, lift and shake the suspended metal vertically so that a maximum amount of hot water will drip off the metal surface and back into the beaker. Quickly but carefully drop the metal into the calorimeter and cover with the cardboard. Make sure that the metal sample is completely covered with water.
- Replace the thermometer through the hole in the cardboard top on the calorimeter. Adjust the thermometer's height so that it is not touching the metal or the Styrofoam cup yet the water is covering the thermometer bulb.

11. Gently swirl the calorimeter cup until an equilibrium temperature (highest temperature) is reached. Watch the thermometer closely as it rises. Sometimes it will rise, fall, and rise again due to the uneven distribution of heat within the calorimeter. Swirling the water in the calorimeter distributes the heat uniformly. Monitor the temperature and record the highest temperature attained to the nearest 0.2 °C. This is the final temperature,  $T_f$ .
12. Repeat this procedure two more times using the same metal sample. Dry the calorimeter and thermometer between trials.

### **Part III. Calculating the Enthalpy of an Endothermic Reaction**

The cold packs in some first-aid kits are made of ammonium nitrate pellets encased in a plastic bag surrounded by water. When the cold pack is bent, the inner bag is broken and an endothermic reaction occurs as the ammonium nitrate dissolves in the water. As a result the pack gets colder. You will be simulating this reaction in your calorimeter in order to calculate the enthalpy of reaction,  $\Delta H_{\text{rxn}}$ , in J/mol.

1. Set up your calorimeter by placing the two Styrofoam cups in a 250 mL or a 400 mL beaker as before. Take the top Styrofoam cup from the calorimeter, place it on the balance and tare it. Weigh out about 25 grams of room temperature deionized water into the calorimeter and record the mass of the water to the nearest thousandth of a gram.
2. Place the cup back into the calorimeter set-up. Take your 4"x 4" piece of cardboard, with the hole in the center, and place it on top of the coffee cups. Insert a thermometer through the hole.
3. Using the buret holder, gently clamp the thermometer and lower it into the cup so that the whole bulb is covered with water but is not touching the bottom of the cup.
4. Tare a clean weigh boat and weigh out about 5 g of ammonium nitrate. Record the mass of the ammonium nitrate to the nearest thousandth of a gram.
5. Just before transferring the ammonium nitrate to the calorimeter, measure the temperature of the water in the calorimeter and record to the nearest 0.2 °C. This is the initial temperature of the water,  $T_i^{\text{water}}$ , and calorimeter,  $T_i^{\text{calorimeter}}$ . Next, remove the thermometer and cardboard top from the calorimeter. Carefully, add the ammonium nitrate to the calorimeter and cover with the cardboard. Make sure that none of the ammonium nitrate or water spills out of the calorimeter.
6. Replace the thermometer through the hole on cardboard top on the calorimeter. Adjust the thermometer's height so that it is not touching the bottom or sides of the calorimeter yet the water is covering the thermometer bulb.

7. Gently stir the water in the calorimeter until the ammonium nitrate is dissolved and the lowest temperature is reached. Watch the thermometer closely. Stirring the solution in the calorimeter achieves a uniform temperature throughout the calorimeter. Monitor the temperature and record the highest temperature attained to the nearest 0.2 °C. This is the final temperature,  $T_f$ .
8. Repeat this procedure two more times. Clean and dry the calorimeter and thermometer between trials.

#### **Part IV. Calculating the Enthalpy of Exothermic Reactions**

Neutralization reactions are exothermic reactions. You will be measuring quantities to estimate the enthalpy change for the neutralization of hydrochloric acid with sodium hydroxide. The number you will calculate is not, strictly speaking, the enthalpy of reaction of hydrochloric acid and sodium hydroxide. The heat released by diluting the acid and the base is also included in that number.

1. Clean and DRY your Styrofoam cups and set up your calorimeter by placing the two Styrofoam cups in a 250 mL or a 400 mL beaker as usual. Take your 4"x 4" piece of cardboard, with the hole in the center, and place it on top of the coffee cups. Insert a thermometer through the hole. DO NOT ADD WATER TO YOUR CALORIMETER.
2. Using the buret holder, gently clamp the thermometer and lower it into the cup so that the bulb is near but not touching the bottom of the cup.
3. Carefully, measure out about 15 mL of **6.0 M** hydrochloric acid in a clean and dry graduated cylinder. Record the volume to the nearest 0.2 mL.
4. Lift cardboard top from the calorimeter keeping the thermometer in place. Carefully, transfer the hydrochloric acid from graduated cylinder to the calorimeter. Make sure that none of the hydrochloric acid splashes out of the calorimeter. DO NOT ADD ANY WATER TO THE CALORIMETER.
5. In a CLEAN and DRY graduated cylinder, carefully, measure out about 15 mL of **6.0 M** sodium hydroxide. Record the volume to the nearest 0.2 mL.
6. Just before transferring the sodium hydroxide to the calorimeter, measure the temperature of the hydrochloric acid in the calorimeter and record to the nearest 0.2°C. This is the initial temperature of the water,  $T_i^{\text{water}}$ , and calorimeter,  $T_i^{\text{calorimeter}}$ . The solutions 6.0 M HCl(aq) and 6.0 M NaOH(aq) already contain water and we are using this water as the calorimeter water.
7. Carefully, add the sodium hydroxide to the calorimeter and cover with the cardboard. Make sure that none of the sodium hydroxide or hydrochloric acid spills out of the calorimeter. Adjust the thermometer's height, if needed, so that it is not

touching the bottom or sides of the calorimeter yet the solution is covering the thermometer bulb.

8. Gently stir the water in the calorimeter until the solutions are well mixed and the highest temperature is reached. Watch the thermometer closely as it rises. Stirring the solution in the calorimeter distributes the heat throughout the calorimeter. Monitor the temperature and record the highest temperature attained to the nearest 0.2 °C. This is the final temperature,  $T_f$ .
9. Repeat this procedure two more times. Clean and dry the calorimeter, thermometer, and graduate cylinder between trials.

### **Part V. Calculating the Heat of Fusion of Water**

Now is your chance to design an experiment. This is an especially important experience since many of you will be pursuing careers in fields that require you to do research. In this part of the experiment you will design an experiment to determine the heat of fusion of ice. You will probably need your calorimeter, ice, water, and a balance. You may use any of the equipment in your locker. Be sure your method is repeatable. See how close you can come to the known result. Have fun!

1. Design an experiment to determine the heat of fusion of ice. You should come to the laboratory with an outline of the procedure you plan to use.
2. Do the experiment performing 3 separate trials. Write-up the detailed procedure you used.

Clean Up. All solutions may be disposed of by washing down the sink with copious amounts of water. Be sure to rinse out the calorimeter before returning it to the box at the front of the room.

### **Data Analysis**

#### **Part I**

1. In step 8, why do you not want any of the water on the outside of the test tube to drip off into the calorimeter?
2. For each trial, calculate the heat lost by the hot water,  $q_{\text{hot water}}$ . Is this quantity positive or negative? The specific heat of water is 4.184 J/g·C.
3. For each trial, calculate the heat gained by the cool water,  $q_{\text{cool water}}$ . Is this quantity positive or negative?

- For each trial, calculate the heat gained by the calorimeter,  $q_{\text{calorimeter}}$ . This can be done by using the equation:  $-q_{\text{hot water}} = (q_{\text{cool water}} + q_{\text{calorimeter}})$ . Is  $q_{\text{calorimeter}}$  a positive or negative quantity? Hint: Be careful with your negative values here. Remember that  $(-q_{\text{hot water}})$  has the opposite algebraic sign value of  $(q_{\text{hot water}})$ .
- For each trial, calculate the heat capacity of your calorimeter. Hint: Write an expression for the heat capacity of the calorimeter in terms of  $q_{\text{calorimeter}}$  and the temperature change of the calorimeter. Note that the temperature change of the calorimeter is assumed to be the same as the temperature change of the "cool water" in the calorimeter.
- Calculate the average heat capacity for the calorimeter.
- Calculate a standard deviation for the average heat capacity.
- Calculate a 90% confidence limit for this data.

## Part II

- Why do we want the water to drip off the metal before it is placed in the calorimeter?
- Calculate the specific heat for your metal for each trial. Remember that the heat lost by the metal is equal to the heat gained by the water in the calorimeter and by the calorimeter itself. This can be expressed as  $-q_{\text{metal}} = (q_{\text{water}} + q_{\text{calorimeter}})$ . The specific heat capacities are positive numbers.
- Calculate the average specific heat capacity for your metal sample.
- Calculate the standard deviation of your average specific heat capacity.
- Using the physical properties of your metal, i.e. density & color, identify your metal.
- Calculate the percent error of your average specific heat as compared to the accepted value.  $C_{\text{sp}}(\text{Pb}) = 0.128 \text{ J/g}\cdot^{\circ}\text{C}$ ;  $C_{\text{sp}}(\text{Al}) = 0.900 \text{ J/g}\cdot^{\circ}\text{C}$ ;  $C_{\text{sp}}(\text{Cu}) = 0.387 \text{ J/g}\cdot^{\circ}\text{C}$

## Part III

- Write a chemical equation that describes the dissolution of ammonium nitrate in water.
- For each trial, calculate the number moles of ammonium nitrate dissolved.
- For each trial, calculate the heat gained by the chemical system of ammonium nitrate,  $q_{\text{rxn}}$ . This can be done by using the equation:  $q_{\text{rxn}} = -(q_{\text{water}} + q_{\text{calorimeter}})$ .

The calorimeter and the water are losing heat. Therefore,  $q_{\text{water}}$  and  $q_{\text{calorimeter}}$  are negative values.

18. The heat transfer in the calorimeter is taking place at constant pressure. Therefore, we can equate the heat gained by the chemical system of ammonium nitrate,  $q_{\text{rxn}}$ , to its enthalpy of reaction,  $\Delta H_{\text{rxn}}$ . For each trial, calculate the enthalpy of the reaction per mole of ammonium nitrate in units of Joules per mole.
19. Calculate the average enthalpy of reaction for the dissolution of ammonium nitrate in J/mol.
20. Calculate the standard deviation of the enthalpy of reaction.
21. Is the dissolution reaction of ammonium nitrate enthalpy-favored? Explain your answer.

#### Part IV

22. Write the chemical equation for the neutralization reaction of hydrochloric acid and sodium hydroxide.
23. Calculate the number moles of hydrochloric acid used in the reaction for each trial.
24. In order to calculate the gained by water,  $q_{\text{water}}$ , the mass of calorimeter must be determined. Calculate the mass of water using the combined volume of 6.0 M hydrochloric acid with 6.0 M sodium hydroxide and the density of water, 1.00g/mL.
25. For each trial, calculate the heat lost by the chemical system,  $q_{\text{rxn}}$ . This can be done by using the equation:  $q_{\text{rxn}} = -(q_{\text{water}} + q_{\text{calorimeter}})$ .
26. The heat transfer in the calorimeter is taking place at constant pressure. Therefore, we can equate the heat lost by the chemical system,  $q_{\text{rxn}}$ , to its enthalpy of reaction,  $\Delta H_{\text{rxn}}$ . For each trial, calculate the enthalpy of the neutralization reaction per mole of hydrochloric acid in units of Joules per mole.
27. Calculate the average enthalpy of reaction for the neutralization in J/mol.
28. Calculate the standard deviation of the enthalpy of reaction.
29. Is the neutralization reaction enthalpy favored? Explain your answer.

#### Part V

30. For each trial, calculate the heat lost by the water in the calorimeter,  $q_{\text{water}}$ .

31. For each trial, calculate the heat lost by the calorimeter,  $q_{\text{calorimeter}}$ .
32. For each trial, calculate the heat gained by the ice,  $q_{\text{ice}}$ .
33. For each trial, calculate the heat of fusion per gram of ice.
34. Calculate the average heat of fusion.
35. Calculate the standard deviation for the average.
36. Calculate your percent error. The accepted value for the Heat of Fusion of ice, according to the textbook, Petrucci et. al., 8<sup>th</sup> edition, is 330 J/g. Note that this value is reported here only to 2 significant figures.

**Conclusion.** Compose a few paragraph summary of this experiment. Include some comments about the sources of error in the experiment that may be responsible for the difference between the values you have obtained and the accepted literature values for the properties you studied in this experiment. Discuss the reasons for your measured value of the specific heat of the metal being too high or too low.