Experiment 5

Measuring Heat Transfer with a Calorimeter

Introduction

People have been controlling the temperature of their environment for thousands of years. Some of these temperature changes are caused by physically adding or removing heat from a system; for example, water boils when you heat it on a stove and ice freezes in a freezer. Other changes are due to chemical reactions, such as wood burning in a fireplace or a chemical cold pack cooling down when its internal chemicals are mixed. A chemical reaction that generates heat, such as burning wood, is said to be **exothermic**, meaning heat is going out of the system, or "**ex**iting". The heat is being transferred from the system, the burning wood, to the surroundings, the air around it. Thus, you can feel the heat being generated. A reaction that uses oxygen and a fuel to produce heat is called a "**combustion**" reaction. Most furnaces, including oil, gas, and wood, use combustion reactions to provide heat for residences and commercial properties.

Other reactions, like those used in commercial cold packs, absorb heat from their surroundings, which causes them to feel cooler than the air. These reactions are **endothermic**, meaning heat goes in or is "**en**tering". Endothermic reactions require heat to be added to the system before the reaction can go to completion; they gain this heat by pulling it from whatever material the system touches, including the air or skin. The surroundings (skin in the case of a cold pack), become cooler as the heat is being pulled out.

The energy released or absorbed in these reactions is not random or unpredictable. Heat transfer between the system (the reactants), and the surroundings (everything outside the system), can be measured down to the last **joule** though a process called **calorimetry** and is dependent upon the amount of material, the temperature change seen during the heat transfer, and the type of material itself.

<u>Calorimetry and Heat Measurements</u>

It is often difficult to directly measure the amount of heat, **q**, absorbed or generated by a chemical reaction. However, it is easy to measure the change in the temperature of the surroundings that occurred because of that reaction. Provided all the energy generated from the reaction was released as heat, the amount of heat absorbed or generated by a chemical reaction can be determined by measuring the change in the temperature of the surroundings.

Reaction generates heat → surroundings absorb heat

$$q_{rxn} = -q_{surroundings}$$

Temperature decreases for system → Temperature increases for surroundings

$\Delta T = T_f - T_i < 0$ $\quad\quad$ $\Delta T = T_f - T_i > 0$

If the reaction takes place in a well-insulated container, nearly all the heat exchanged in the reaction will be confined to the contents of the container, causing the temperature of the contents to change. In calorimetry, an insulated reaction container, called a **calorimeter**, is filled with a measured amount of a liquid with a known **heat capacity**. Heat capacity is the amount of heat needed to raise the temperature of a material by one degree centigrade. Measuring the temperature change of the liquid in the calorimeter makes it possible to calculate the amount of heat transferred. Ideally all the heat will be transferred to the contents, but inevitably some escapes both outside the calorimeter and to the calorimeter itself. The amount of heat that is absorbed by the calorimeter can be determined and is referred to as the **calorimeter constant**.

This heat must be accounted for when calculating the total heat generated or absorbed by a chemical reaction. The calorimeter constant is defined as the amount of heat per °C. However, if the calorimeter is sufficiently insulated, both the heat lost to the calorimeter and outside the calorimeter will be negligible and can be ignored. In this experiment, we will assume that we have a perfect calorimeter and 100% of the heat generated or absorbed in a reaction is transferred between the system and the liquid in the calorimeter (the surroundings).

Specific Heat and Heat Capacity

We know that if two objects, initially at different temperatures, are placed together, they will eventually come to equilibrium at some intermediate temperature. Provided no heat is lost to or gained from the surroundings, the quantity of heat lost by the warmer object is equal to that gained by the cooler one. Several measurements have been defined to compare the relative amounts of heat that can be absorbed by different materials. First, the **Specific Heat** (s) of a substance is the amount of heat required to raise the temperature of one gram of the substance one degree centigrade. The specific heat has been calculated for many substances and can be used as a means of identification, as it is "specific" for a material. A second quantity, the **Heat Capacity** (C), is the quantity of heat required to raise the temperature of the entire mass of the material by one degree centigrade. Thus:

$$\text{Heat capacity, } C = \text{mass} \times s$$

The heat generated in a reaction can be calculated from the heat capacity of the substance and the temperature change.

$$\text{Heat, } \mathbf{q} = C \times \Delta T = \mathbf{mass \times s \times \Delta T}$$

Enthalpy, H

The quantity of heat, q, released or absorbed in a chemical reaction is the difference in the heat content between products and reactants. If the reaction takes place at constant pressure, we refer to the heat content of a substance as its **Enthalpy**, H. The amount of heat involved in a chemical reaction, q, equals the change in enthalpy, ΔH, and is defined as:

$$q_{rxn} = \Delta H = H_{products} - H_{reactants}$$

If the products contain more heat than the reactants, they must have absorbed heat from the surroundings and $\Delta H > 0$. A reaction that absorbs energy from the environment is called an endothermic reaction and the temperature of the surroundings, such as the water in a calorimeter, decreases. **Endothermic reactions are symbolized by a positive ΔH.**

If the products contain less heat than the reactants, some heat must have been released into the surroundings; so $\Delta H < 0$. A reaction that releases energy to the environment is called an exothermic reaction and the temperature of the surroundings, such as the water in a calorimeter, increases. **Exothermic reactions are symbolized by a negative ΔH.** The enthalpy of a chemical reaction is associated with how the reaction is written so the chemical reaction must either be given or the enthalpy defined on a molar basis of one of the reactants or products.

Standard Enthalpy Tables

Standard enthalpy tables are available that list the enthalpy values for many common types of reactions, such as enthalpy of formation, enthalpy of solution, etc. These tables can be found in your text book and are based on the production of 1 mole of product.

For example: The standard enthalpy of formation, **ΔH°**, of sodium chloride would be based on the elements making up sodium chloride in their common form producing 1 mole of product at 25°C and 1atm.

$$Na(s) + \tfrac{1}{2} Cl_2(g) \rightarrow NaCl(s) \qquad \Delta H° = -411 kJ$$

Tabulated standard enthalpy of formation tables are often written with units of kJ/mol since the definition is based on the production of a single mole of a compound. Tables for enthalpy of solution, combustion, and other common chemical and physical reactions are also available. These tables make it possible to predict how much heat is associated with a chemical reaction without having to do the calorimetry experiment yourself.

Heats of Solution and Reaction

Heat can be transferred in a variety of processes. When heat is absorbed or released during a reaction, it is referred to as the heat of reaction, or ΔH_{rxn}. However, heat is also transferred between the system and the surroundings when ionic salts are dissolved in water to form an aqueous solution. The energy transferred when dissolving a solid is referred to as the heat of solution, ΔH_{soln}, or enthalpy of solvation, ΔH_{solv}, and may be positive or negative depending on whether the action of solvating the ions is an exothermic or endothermic process. The enthalpy of solvation is a function of the particular salt and remains constant per mole of the salt. Thus, measuring the enthalpy of solvation can be used as a means of identification for ionic salts by matching the enthalpy of solvation found experimentally to the accepted value found in thermodynamic tables.

In Your Experiment

You will perform several different calorimetry experiments and use the temperature changes to calculate the specific heat of a metal, the heat of reaction of an acid with a metal, and the heat of solution for several different salts dissolving in water. Some of these temperature changes will be caused by the release of energy, indicating an exothermic reaction and resulting in an increase in the temperature of the surroundings. In at least one of the salt solutions, the temperature of the surroundings (the solution) will decrease upon dissolving, indicating that an endothermic reaction has taken place. Be very careful when measuring the initial and final temperatures. When measuring some of the heats of solution, the temperature will only increase or decrease by a degree or two, so it is necessary to record the temperature to as many significant digits as possible.

Lab Precautions

Chemical Hazards

Magnesium metal
NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 1
REACTIVITY: Will react vigorously with acidic solutions.

3M Hydrochloric Acid Solution
NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 0
DERMAL EXPOSURE:
Wash off with soap and plenty of water.
EYE EXPOSURE:
Flush with copious amounts of water for at least 15 minutes. Contact your TA immediately.

Ammonium chloride (solid)
NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 0
DERMAL EXPOSURE:
Wash off with soap and plenty of water.
EYE EXPOSURE:
Flush with copious amounts of water for at least 15 minutes. Contact your TA immediately.

Sodium chloride (solid)
NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 0
DERMAL EXPOSURE:
Wash off with soap and plenty of water.
EYE EXPOSURE:
Flush with copious amounts of water for at least 15 minutes. Contact your TA immediately.

Calcium Chloride (solid)
NFPA RATING: HEALTH: 2 FLAMMABILITY: 0 REACTIVITY: 0
DERMAL EXPOSURE:
Wash off with soap and plenty of water.
EYE EXPOSURE:
Flush with copious amounts of water for at least 15 minutes. Contact your TA immediately.

Equipment Hazards

Remember that hot plates do not look any different when they are hot. Be sure to turn off hotplates before leaving lab.

Chemical Disposal

Dispose of all waste solutions in waste container.

Experiment 5: Procedures and Data Sheet

(**Submit as part of your informal report**)

Name: ____ Date: _______ Section: _______

TA Signature: __________________________________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all measurements with the correct number of significant figures and units.
TA signature & TA initials on any changes made to the data are required or the data is invalid.

Note: You should start boiling your water before your TA's pre-lab talk.

Part 1: Determining the Specific Heat of a Solid

1. Fill a 250 ml beaker about ½ full of water and place on a hot plate to boil.
2. Tare the balance and record the mass of the metal cylinder at your station.

Mass of metal cylinder: ________________

3. Use the test tube holder to lower the cylinder into the water.
4. Insert the temperature probe into the boiling water and record the temperature as the initial temperature of the metal. It should be approximately 100.0°C. Do not rest the temperature probe on the bottom of the beaker. You want the water temperature, not the temperature of the glass.

Temperature of boiling water bath, T_I (metal): __________________

5. Use your 25mL graduated cylinder (not a beaker!) to add exactly 150.0 mL of water to the calorimeter. (Use the large thermos calorimeter, not the Styrofoam cup calorimeter for this part of the experiment.)
6. Use your temperature probe to record the temperature of the water in the calorimeter as your initial water temperature.

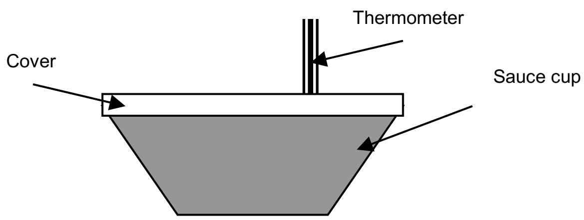
Temperature of water in thermos, T_I (water): __________________

7. Quickly transfer the cylinder from the water bath to the calorimeter. DO NOT SPLASH ANY OF THE WATER OUT OF THE CALORIMETER!
8. Insert the temperature probe into the calorimeter, stir the water gently, and monitor the temperature until it reaches a maximum.
9. Record the maximum temperature reached as the final temperature for both the metal and the water.

Temperature of water in thermos, T_f (water and metal): _________________

Part 2. Determining the Enthalpy of Solvation
Do all 3 salts: 1 exothermic, 1 endothermic and 1 no real change (thermally neutral)
Note the variation in temperature that you see naturally in just the water before the salt is added; if adding salt does not cause any more variation than you see naturally, it is considered thermally neutral.

1. Assemble a Styrofoam calorimeter as shown in the figure below.



2. Add exactly 15.0 mL of water to the calorimeter.
3. Insert the thermometer through the cover of the calorimeter and monitor the water temperature.
4. When the temperature remains constant, record the water temperature as your initial temperature.
5. Weigh out approximately 1g of the calcium chloride, $CaCl_2$.
6. Record the exact mass of the salt.
7. Carefully, but quickly add the salt to the calorimeter. DO NOT SPILL ANY OF THE SALT
8. Monitor the temperature until the temperature reaches a maximum (or a minimum for an endothermic reaction) and begins to change direction. Record any observations.
9. Record this maximum or minimum temperature as your final temperature.
10. Discard the contents of the calorimeter in the waste container.
11. Rinse the calorimeter with water.
12. Repeat the procedure with the remaining 2 salts.

Part 2 Data

	$CaCl_2$	NaCl	NH_4Cl
Temperature of water in calorimeter (Ti)			
Mass of salt			
Temperature of solution in calorimeter (Tf)			

Part 3. Determining the Heat of Reaction

1. Add 10.0 mL of 3.0 M hydrochloric acid to the Styrofoam calorimeter.
2. Insert the thermometer through the cover of the calorimeter and monitor the acid temperature.
3. When the temperature remains constant, record the acid temperature as your initial temperature.

Temperature of acid in calorimeter (T_i): ___________

4. Weigh out approximately 0.2g of magnesium metal.
5. Record the exact mass.

Mass of magnesium metal: ___________

6. Carefully, but quickly, add the magnesium to the calorimeter.
7. Monitor the temperature until the temperature reaches a maximum.
8. Record this temperature as your final equilibrium temperature.

Temperature of solution in calorimeter (T_f): ___________

9. Record any observations you saw during the reaction.

____

____

____

Experiment 5: Data Rubric (20pts)

Points

Data are neat and legible	5pts	__________pts
Significant figures (>80% correct)	3pts	__________pts
Units (>80% correct)	2pts	__________pts
All data are present and make sense	10pts	__________pts

Deductions (sliding scale based on TA discretion)

Lab area left unclean	–20pts	__________pts
Improper waste disposal	–20pts	__________pts
Disruptive behavior	–20pts	__________pts
Lab coat or safety glasses removed while in lab	–20pts	__________pts
Data sheet is missing TA signature	–20pts	__________pts
Other: ____		__________pts

Comments: ____

Grade for Data Sheet __________**pts**

Experiment 5: Results Table
(Submit as part of your informal report)

Name: ____ Date: ________ Section: ________

All results must be written in pen. **Pencil is not allowed!!**
Record all results with the correct number of significant figures and units
No stray marks or notes should be present on this page. Only the tabulated results are allowed

Part 1: Determining the Specific Heat of a Solid

Mass of water (m_w)	
Temperature change of the water (ΔT_w)	
Heat gained by water (q_w)	
Temperature change of the metal (T_m)	
Heat lost by metal (q_m)	
Specific heat of the metal	
Metal identity	
Percent error	

Part 2: Determining the Enthalpy of Solvation

Salt used	$CaCl_2$	NaCl	NH_4Cl
Mass of water (m_w)			
Combined mass of reactants			
Temperature change of the water (ΔT_w)			
Heat lost or gained by the solution (q_{soln})			
Heat lost or gained by the salt (q_{salt})			
Moles of salt			
ΔH_{solv}			

Part 3: Determining the Heat of Reaction

Mass of acid solution (m_a)	
Total mass of aqueous reaction solution (m_{soln})	
Temperature change of the aqueous solution (ΔT_{soln})	
Heat absorbed by the solution (q_{soln})	
Heat generated by the reaction (q_{rxn})	
Moles of magnesium	
ΔH_{rxn}	

Experiment 5: Results Table Rubric (20pts)

Points

Tables are neat and legible	5pts	__________pts
Significant figures (>80% correct)	3pts	__________pts
Units (>80% correct)	2pts	__________pts
All results are present and make sense	10pts	__________pts

Deductions (sliding based on TA discretion)

Results to not match data	-20pts	__________pts
Plagiarism!!! Results are identical to another student	**-100pts**	**__________pts**

Other: ____ __________pts

Comments: ____

Grade for Results Table __________pts

Experiment 5: Calculations

(Submit as part of your informal report)

Part 1: Determining the Specific Heat of a Solid

Mass of Water, m_w
Use the value 1.000g/mL as the density of water to calculate the mass of the 150.0mL of water that you added to your calorimeter.

Mass of Water, m_w: ____________

Temperature change of the water, ΔT_w
Subtract the initial temperature of the water in the calorimeter from the final equilibrium temperature of the water in the calorimeter. Since the water absorbed heat from the metal, you should get a positive number.

$$\Delta T_w = T_f - T_i$$

Temperature change of the water, ΔT_w: ____________

Heat Gained by Water, q_w
Use the equation below to calculate the heat absorbed by the water in your calorimeter. Since you are calculating the heat gained by the water, you will want to use the mass of water in your calorimeter, the change in temperature of the water, and the specific heat of water. (The specific heat of water is 4.184J/g°C).

$$q_w = m_w \times S_w \times \Delta T_w$$

Heat Gained by Water, q_w: ____________

Temperature change of the metal, ΔT_m
The initial temperature of the metal is the same as the temperature of the boiling water bath in which it was heated. During the calorimetry experiment, the temperature of the metal and the water in the calorimeter were allowed to equilibrate – i.e. they exchanged heat until their temperatures were the same – so the final temperature of the metal is equal to the final temperature of the water in the calorimeter. Subtract these temperatures to get the change in temperature of the metal. Since the metal lost heat to the water, you should get a negative number.

$$\Delta T_m = T_f - T_i$$

Temperature change of the metal, ΔT_m: ____________

Heat Lost by Metal, q_m
If we assume that no heat was lost in the transfer of the metal to the water, then:

$$q_m = -q_w$$

Heat Lost by Metal, q_m: ____________

Specific Heat of the Metal, S_m
Rearrange the calorimetry equation to solve for the specific heat (s) of the metal. Since you are now looking at the metal, you will want to use the mass of the metal, the heat lost by the metal, and the temperature change of the metal.

$$s_m = \frac{q_m}{(m_m x \Delta T_m)}$$

Specific Heat of the Metal, s_m: ____________

Metal Identity
Compare your value for the specific heat to the list of values provided below for several different metals. Identify your metal based both on observation (e.g. copper and zinc have similar specific heats but are different colors) and the calculated specific heat.

Copper	0.386 J/g°C	Aluminum	0.900 J/g°C
Zinc	0.387 J/g°C	Lead	0.128 J/g°C
Iron	0.449 J/g°C		

Metal Identity: ____________

Percent Error in Specific Heat
Calculate the percent error between your value and the accepted value for specific heat given in the table above.

Percent Error: ____________

Part 2: Determining the Enthalpy of Solvation (3 salts)

Mass of Water
Use the density of water given in Part 1 to convert the 15.0mL of water to grams.

Mass of Water, m_w: ____________ ____________ ____________

Mass of Solution, m_{sol}
The mass of the solution is the combined mass of the water and the salt.

Mass of solution, m_{sol}: ____________ ____________ ____________

Temperature change of the Solution, ΔT_{sol}: $\Delta T_{sol} = T_f - T_i$
Values should be positive for an exothermic reaction and negative for an endothermic reaction.

Temperature change of the Solution, ΔT_{sol}: ____________ ____________ ____________

Heat Lost or Gained by the Solution, q_{soln}

Use the calorimetry equation to calculate q_{soln}. Your values here are all measured for the solution, which is the "surroundings" for the reaction. Assume that the specific heat of the solution is the same as the specific heat of water. (4.184 J/(g °C)).

$$q_{solution} = (m_{solution} \times S_{solution} \times \Delta T_{solution})$$

Heat Lost or Gained by the Solution, q_{soln}: ____________ ____________ ____________

Heat Lost or Gained by the dissolving salt, q_{salt}

In this experiment, we are considering the system to be the reaction itself and the calorimeter and resulting solution to be the surroundings. Assume that all the heat generated by the reaction, (the system) is transferred to the solution (the surroundings). Therefore, the heat generated or absorbed by dissolving the salt is simply the negative of the heat absorbed by or removed from the solution.

$$q_{salt} = -q_{soln}$$

Heat Lost or Gained by the dissolving salt, q_{salt}: ____________ ____________ ____________

Moles of salt

Use the molar mass of the salt to convert the mass of the salt to moles.

Moles of salt: ____________ ____________ ____________

ΔH_{solv}

The molar enthalpy of solvation is based on the heat lost or gained per mole of material dissolved, and is generally reported in kJ/mol. Use the equation below to calculate the ΔH_{solv} for each salt. Since your measured q values are in joules, you will need to convert to kilojoules. The enthalpy, ΔH_{solv}, will be negative for an exothermic reaction and positive for an endothermic reaction.

$$\Delta H_{solv} = q_{salt} / (\text{moles salt})$$

ΔH_{solv}: ____________ ____________ ____________

Part 3. Determining the Heat of Reaction

Mass of Acid solution, m_a

The acid solution is dilute, so the density is approximately that of water. You can use 1.00g/mL to convert the volume of acid to mass. (Note that this is the mass of HCl and water together.)

Mass of acid, m_a: ____________

Total Mass of Reaction Solution, m_{soln}

The total mass of the solution is the combined mass of the acid, the water, and the magnesium.

Mass of Solution, m_{soln}: ____________

Temperature Change of the Aqueous Solution, ΔT_s

You should get a positive number for an exothermic reaction and a negative number for an endothermic reaction.

$$\Delta T_{soln} = T_f - T_i$$

Temperature Change of the Aqueous Solution, ΔT_{soln}: ___________

Heat Absorbed by the Solution, q_{sol}

The solution is again the "surroundings" in this experiment. You are assuming the specific heat of the solution is the same as the specific heat of water (4.184 J/g°C).

$$q_{soln} = (m_{soln} \times S_{soln} \times \Delta T_{soln})$$

Heat Absorbed by the Solution, q_{soln}: ___________

Heat Generated by the Reaction, q_{rxn}

In this experiment, we are considering the system to be the reaction itself and the calorimeter and resulting solution to be the surroundings. Assume that all the heat generated by the reaction, (the system) is transferred to the solution (the surroundings). Therefore, the heat generated by the reaction is simply the negative of the heat absorbed by the solution.

$$q_{rxn} = -q_{soln}$$

Heat Generated by the Reaction, q_{rxn}: ___________

Moles of Metal

Use the molar mass to convert the mass of the magnesium to moles.

Moles of magnesium: ___________

ΔH_{rxn}

Use the equation below to calculate the enthalpy of reaction in kilojoules per mole of magnesium.

$$\Delta H_{rxn} = q_{rxn} / (\text{moles magnesium})$$

The balanced chemical equation for this reaction is:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The enthalpy of reaction for your experiment would be based on the heat lost or gained per 1 mole of magnesium, because magnesium is the limiting reagent. The enthalpy, ΔH, will be negative for an exothermic reaction and positive for an endothermic reaction.

ΔH_{rxn}: ___________

There are no additional questions for this experiment.

Experiment 5: Prelab Worksheet

(Submit via Brightspace before the start of your lab session.)

Name: ______________________________ Date: ________ Section: ________ Grade: ________

All information needed to complete this worksheet can be found in the pre-lab information and calculations sections of the lab manual. Read this introductory material first!

- Record all values with the correct number of significant figures and units.
- Place all answers on the line when provided.
- Show calculations for any numerical answers; **work must be shown to receive credit**.
- See any 102 TA in the help office before your prelab is due if you have any questions.
- Each question is worth 2 points.

1. What temperature (approximately), in °C, will you use for the initial temperature of the metal when determining specific heat?

____

2. Which calorimeter do you use for determining the specific heat of the metal, the thermos or the Styrofoam cup?

____

3. How much energy, in joules, does it take to raise the temperature of 255 g of water by 12.5°C ? (s_w = 4.184 J/g°C)

____

4. Calculate the heat capacity of the water in the previous problem. Include units in your answer.

____

Use the following information to answer questions 5-10. Remember to show all work.

Sodium metal reacts with water to produce hydrogen gas and sodium hydroxide according to the chemical equation shown below. When 1.4g of sodium metal is added to 100.00 g of water, the temperature of the resulting solution rises from 25.00 °C to 32.8 °C. The density of the final solution is 1.000g/mL. The specific heat of the final solution is 4.18 J/g°C.

$$2Na\ (s) + 2H_2O(aq)\ +\ \rightarrow\ 2NaOH\ (aq)\ +\ H_2\ (g)$$

5. Calculate the mass of the final solution in grams.

____

6. Calculate the heat (in Joules) gained by the solution. Be careful with the sign!

____

7. What is the amount of heat (in Joules) generated from the reaction? (Be careful of the sign in your answer!)

____

8. Calculate the moles of Na used.

____

9. Calculate the enthalpy change, ΔH, in kilojoules per mole of Na used.

____

10. Is this reaction exothermic or endothermic?

____