NAME________________________________ **SECTION________________**

PARTNER(S)__________________________ **DATE_____________________**

THE INS AND OUTS OF ENERGY IN SYSTEMS

Physical and chemical processes always occur with the transfer of energy. This activity explores the absorption and release of energy during a physical and chemical change.

Pre-Lab Queries

1. You have just made yourself a cup of hot tea or hot coffee and get called away for 15 minutes. What happens to the coffee or tea during that time? Why?

   Would you call this a physical or chemical change? Explain.

   List an example of a physical change in a system where energy is:
   a. gained

   b. lost

2. Recall the variety of chemical changes that you encountered in *Investigations of Chemical Reactions I.*

   What did you have to do to make $CuCO_3$ convert to CuO? Was energy absorbed or released?

   For the following chemical reactions from that activity, decide whether each change involved the release or absorption of energy based on your actions and observations. Explain your answers.

   $2\ Mg + O_2 \rightarrow 2\ MgO$ ____________________________

   $2\ Fe + O_2 \rightarrow 2\ FeO$ ____________________________

Introduction

You have experienced the exchange of heat energy in physical and chemical systems in the laboratory and in your personal life. From what you know, which way does heat flow?

Generally we say that heat energy flows from a warmer object to a colder object when the process is a physical one. When we speak of energy transfer in chemical or physical changes, it is important to note in which direction energy is flowing. We need to define a **system** (the materials and processes we are studying) and its **surroundings**. In the example of the cup of coffee or tea, the coffee/tea is the system and the air and tabletop would be parts of the surroundings. Energy can be transferred from the coffee (if hot) to the air or from the air to the coffee (if iced). When heat is transferred from a system to the surroundings the process is said to be **exothermic**. Another way to express this is to say that the system loses or releases energy. When the heat is moved from the surroundings into the system the process is termed **endothermic**. In an endothermic process the system gains or absorbs heat.

In this activity you will have a chance to investigate heat exchange in a physical and chemical system.

Part I: Heat Transfer in a Physical System

In this part you will be studying heat transfer between metal cylinders and water. You will be heating the cylinders in boiling water and then placing them in colder water.

What will happen when you put the hot cylinder into the colder water?

How can you monitor the amount of heat transferred?

Will the amount of heat exchanged be the same if the identical hot cylinders are placed in different amounts of water? Explain.

Procedure

Be sure to record all data and results in the appropriate table.

1. As one group, obtain two cylinders (~6 cm long) made of brass, iron or copper. Be sure both cylinders are made from the same material. Place both of the cylinders in a beaker of water and heat the cylinders and water to boiling on a hot plate. Allow the beaker contents to **boil** for at least 5 minutes to be sure each cylinder reaches the same temperature.

2. While the cylinders are heating, obtain a large Styrofoam cup, lid, CBL, temperature probe, and TI-83 calculator. Take the Styrofoam cup and mass it as precisely as possible. Place 75 mL of distilled water in the cup and mass the cup and water.

Why do you think you will be using Styrofoam cups?

3. Attach the temperature probe to the CBL in channel 1 (CH1). Attach the calculator to the CBL by using the black link cable. Be careful inserting the cables but make sure they are firmly pressed into the calculator and CBL.

Suspend the temperature probe through the hole in the cup lid so that the probe extends into the water. Be sure to make any adjustments so that the probe does not touch the sides or bottom of the cup if possible. You need to set up the commands required on the calculator.

❑Turn the calculator ON.
❑Press the **[PRGM]** button. You will see a list of programs. Select the one called CHEMBIO by pressing the **[ENTER]** button if the cursor is on 1. You should now see "pgrmCHEMBIO". Press **[ENTER]** again to get the Vernier software screen. Press **[ENTER]** again to get the "MAIN MENU"
❑Select 1:Set-up Probes and hit **[ENTER]**.
❑When the screen asks how many probes, press 1 and **[ENTER]**.
❑Select the temperature probe by pressing **[ENTER]** when the cursor is on 1.
❑Enter 1 when asked for the channel number.
❑Move the cursor down to COLLECT DATA on the MAIN MENU and press **[ENTER]**.
❑When you get the DATA COLLECTION screen, select TIME GRAPH and **[ENTER]**.
❑For time between sample put in 30 and press **[ENTER]**.
❑When asked for number of samples, put in 30 and press **[ENTER]**. The display will tell you how many second the total run will be. Press **[ENTER]**.
❑Select USE TIME SETUP and **[ENTER]**.
❑For Ymin type in (-) 5 and **[ENTER]**, for Ymax type in 120 and **[ENTER]**, for Yscl type in 10 and **[ENTER]**. This sets the axes and scale for the final graph of your data.

When you press **[ENTER]** the CBL and calculator will record and store the temperatures of the water every 30 seconds for 15 minutes

Two minutes before you are going to add the cylinder to the water in the cup, go back to the calculator and press **[ENTER]**. If the screen is blank hit the **[ON]** key and the screen image will return.

4. Remove one of the metal cylinders from the beaker with crucible tongs and quickly transfer it to one of the cups. Be careful not to splash (and lose) any of the water. Cover the cup. Be sure the probe is not touching the rod and is completely submerged.

Periodically agitate the cup gently to mix the water. Be sure the second rod is covered in boiling water.

5. Once the data collection is complete, the calculator screen will indicate where the data has been stored in the calculator (L_1, L_2). Press **[ENTER]** and the graph of your data will be displayed. You can use the trace function to determine the values for the initial and final temperatures from the graph or locate the values in the calculator list.

6. Repeat the procedure starting at step 2 but add 150 mL of water instead of 75 mL and use the second hot metal rod.

Did the temperature change the same amount in each cup? Why or why not?

Water, like most pure substances, will change its temperature by a certain amount for each unit of heat energy that it absorbs or releases. We express the relationship among heat energy, mass and temperature the following way:

$$\text{Specific Heat Capacity} = \frac{\text{amount of heat absorbed or released}}{(\text{mass of sample}) \bullet (\text{change in temperature})} = \frac{\text{calories or joules}}{\text{g} \bullet {}^{o}\text{C}}$$

The specific heat for water is 4.184 J/ g • ^{o}C.

7. Using the formula for the specific heat capacity and the value for water, calculate the number of joules transferred from the metal cylinder to the water in each cup. Record and compare the values.

How do the values for heat transferred compare?

Is this result expected? Why or why not?

Data Table

Data	Trial 1	Trial 2
Mass of cup (g)		
Mass of cup and water (g)		
Initial Temperature ^oC		
Final Temperature ^oC		

Results Table

Results	Trial 1	Trial 2
Mass of water (g)		
Change in temperature (^oC)		
Heat transferred (J)		

Part II: Heat Transfer in a Chemical System

To investigate heat transfer in chemical systems we will use a reaction you may already be familiar with, the dehydration and hydration of copper (II) sulfate, $CuSO_4$. Normally, copper (II) sulfate is found in the form of bright blue or blue-green crystals. In this form the compound is a **hydrate**, that is, $CuSO_4$ is loosely bound to several molecules of water. The most common form is copper (II) sulfate pentahydrate, $CuSO_4 \bullet 5H_2O$. When heated, this compound decomposes, releasing its water molecules:

$$CuSO_4 \bullet 5H_2O\ (s) \quad + \quad heat \quad \rightarrow \quad CuSO_4\ (s) \quad + \quad 5\ H_2O\ (g)$$

The dehydrated form of copper (II) sulfate is white and is termed **anhydrous** (meaning "without water").

Is the dehydration process exothermic or endothermic?

Since the dehydration is endothermic, the reverse process (called rehydration) must be exothermic. It happens very rapidly.

$CuSO_4$ (s) + 5 H_2O (l) → $CuSO_4 \bullet 5H_2O$ (s) + heat
(white) (blue)

If we do this reaction in an insulated container (calorimeter) containing more water than is required for the actual hydration reaction, we can measure the heat released by this exothermic reaction by monitoring the rise in the temperature of the water when copper (II) sulfate is added. From Part I you know that water has a specific heat capacity of 4.184 J/ g • ^{o}C. We can use this, the mass of the water in the calorimeter, and the temperature change to measure the amount of heat released.

The amount of heat absorbed or released in a chemical reaction is called the **heat of reaction,** given the symbol ΔH. Typical units for ΔH are cal/g or J/kg. If J/mol or cal/mol are used it is called the molar heat of reaction.

Your task in Part II will be to calculate the molar heat of reaction for the hydration of copper (II) sulfate.

Do*es the molar heat of reaction depend on the amount of* $CuSO_4$ *used?*

Procedure

Be sure to record all data and results in the appropriate table

1. Mass out approximately 5 grams of copper (II) sulfate pentahydrate ($CuSO_4 \bullet 5\ H_2O$) in a massing boat. You do not need to record this mass.

2. Transfer the $CuSO_4 \bullet 5\ H_2O$ to a large **pyrex** test tube. Clamp the tube onto a ring stand and position the tube so that it is at a slight angle above horizontal and the mouth is not directed at anyone. GENTLY heat the crystals with a Bunsen burner until the crystals begin to lose their blue color:

   Warning: If you heat too strongly not only is the water lost but $CuSO_4$ converts to black CuS. If you see gray or black at any point, stop heating immediately!

   Moisture may condense at the mouth of the test tube. Why?

   Heat the region where the condensation occurs gently to vaporize the water.

4. Continue heating until the sample appear white throughout. Allow the tube and contents to cool to room temperature (5-10 minutes). Mass a massing boat while you are waiting.

5. When the sample is cool, empty the contents into the pre-massed boat and mass the boat and sample.

   Why did the mass of the solid decrease?

6. Mass your empty Styrofoam cup calorimeter. Add about 30 mL of distilled water to the cup and remass the cup and water.

6. Set up the CBL/probe/calculator as you did before.

7. Make sure that the anhydrous copper (II) sulfate is at about the same temperature as the water. Start the program and allow it to run for about a minute or so. Quickly pour the copper (II) sulfate into the water. Cover and mix gently.

8. When the process is complete, dispose of the copper (II) sulfate solution in the proper waste container. Clean and return all materials including the cup.

9. Using the data and results you collected, calculate the heat transferred to the water from the hydration reaction. From the heat transferred and the moles of anhydrous $CuSO_4$, calculate the molar heat of reaction.

10. Obtain the results for the molar heat of reaction from two other groups and compare your results with theirs.

Data Table

Data	
Mass of boat (g)	
Mass of boat and anhydrous $CuSO_4$ (g)	
Mass of cup (g)	
Mass of cup and water (g)	
Initial Temperature ^oC	
Final Temperature ^oC	

Results Table

Results	
Mass of water (g)	
Mass anhydrous $CuSO_4$ (g)	
Moles of anhydrous $CuSO_4$ (mol)	
Change in temperature (^oC)	
Heat transferred (J)	
D**H (J/mol)**	

D**H (J/mol) for other groups:** __________________ __________________

Post-Lab Questions

1. Explain any factors that many have caused your results to differ between your subgroups (from Part I) or from another group (Part II). Massing errors and spills don't count!

2. A chemist mixes two reactants in a beaker and frost forms on the outside of the beaker. Is the reaction endothermic or exothermic? Why?

3. Copper has a specific heat capacity of 0.385 J/g• ^{o}C. What will the final temperature of a copper sample be if it has a mass of 5.45 kg at 25^{o}C and absorbs the energy released when 100 g of water cools from 100^{o}C to 25 ^{o}C?