

LAB #9- CALORIMETRY/THERMOCHEMISTRY TO THE RESCUE

Objectives

- Determine whether a reaction is endothermic or exothermic.
- Determine the best ionic compound to use in a heat pack for treating frostbite based on your experimental results.
- Calculate the average heat capacity of your calorimeter.
- Calculate the specific heat of a metal.

Introduction

The human body works best within a very narrow temperature range. A temperature drop of as little as 2 °C in the body's core causes classic hypothermia symptoms such as mental difficulties and loss of physical coordination. Much more extreme temperature drops in the extremities may be survived, but can lead to frostbite if the flesh freezes. Victims of hypothermia require immediate treatment, and in outdoor situations the treatment is often warmth provided by portable heat sources, such as heat packs.

Heat packs are available that produce warmth through various chemical reactions. Such heat packs are convenient because they only release heat when triggered. One common heat pack contains an internal pouch of water and a solid powder. Once the pouch of water is broken open, there is an exothermic reaction between the water and the powder. These heat packs have limitations. For example, they do not work well in extreme cold because the water within the pack will freeze.

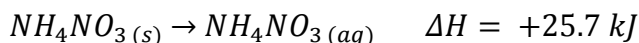
In the coldest environments, heat packs are available that contain only the powder in a resealable waterproof sack. When heat is required, the sack can be opened and any aqueous solution poured inside. The sack is resealed and the reaction produces heat.

Reaction Enthalpy

An important part of chemistry involves studying energy changes that occur during chemical reactions. These energy changes are of fundamental importance in understanding the “driving force” of a chemical reaction. The most common way energy is exchanged between a chemical system and the environment is by evolution or absorption of heat (q). The change in heat energy accompanying a chemical reaction is known as **enthalpy change, ΔH** . By convention, reactions in which heat is absorbed are labeled endothermic and have positive values of ΔH ; reactions in which heat is released are labeled exothermic and have negative values of ΔH .

Cold packs used in athletics are familiar to many sports enthusiasts. In order to derive coldness from the pack, a plastic packet of water is broken inside another packet containing a solid salt such as NH_4NO_3 . In this case, the enthalpy of solution, i.e., heat absorbed when a substance

dissolves, is endothermic indicating that heat is absorbed as the salt dissolves. Thus, the enthalpy of solution is designated with a positive sign since energy is absorbed or added.



On the other hand, gas stoves produce heat through the combustion (burning in oxygen) of methane, equation 1. Since heat is produced by the combustion, the reaction is exothermic and enthalpy of combustion (ΔH), i.e., heat released during combustion, must be negative. In fact, ΔH for the methane combustion is -890.4 kJ (see below). It is important to realize that ΔH is related to the coefficients in the balanced equation. Thus, -890.4 kJ of heat is released per every 1 mole of CH_4 that reacts but per every two moles of H_2O that is formed.



The negative sign serves to reinforce the fact that heat is produced rather than absorbed or that this is an exothermic reaction. We can use stoichiometry to calculate the enthalpy of any amount of a reactant.

Heat and Temperature

If an object (such as a pot of water) is positioned to absorb the heat given off during a combustion reaction, then the temperature of the object will change as follows

$$q = mc\Delta T \quad \text{Equation 2}$$

where

q = the amount of heat absorbed by the object in Joules

m = the mass of the object being heated in grams

c = the specific heat of the object being heated

ΔT = the change in temperature of the object = final

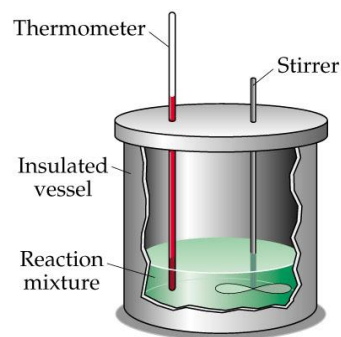
temperature minus the initial temperature = $T_f - T_i$

The specific heat is different for different substances. The specific heat or heat capacity is low for objects which are easily heated or cooled with minimal energy input. Metals usually have low specific heats because their temperature changes very quickly. Water has a high specific heat because it takes a lot of energy to change the temperature of a sample of water. For example:

<u>Substance</u>	<u>Specific Heat ($J g^{-1} ^\circ C^{-1}$)</u>
water	4.18
air	1.01
aluminum	0.897
granite	0.790

Calorimetry

Many experiments in thermochemistry involve a calorimeter. A calorimeter, like the one to the right, is simply a container that insulates a reaction from the surrounding environment so change in temperature as the reaction proceeds place can be measured accurately measured as independent of the environmental temperature. In lab, you will use two styrofoam cups as your calorimeter.



Ideally any calorimeter would be able to completely maintain the system without losing heat to the surroundings. Unfortunately, no calorimeter is perfect and heat is always lost to the surroundings. It is possible to determine how much heat is lost by the calorimeter by mixing hot water and room temperature water. The heat lost by the hot water is equal to the heat gained by the room temperature water.

$$|q_{\text{lost by hot water}}| = |q_{\text{gained by room temperature water}}| + |q_{\text{gained by calorimeter}}| \quad \text{Equation 3}$$

In order to calculate the heat lost to the calorimeter you will calculate q for both the hot water and the room temperature water (using Equation 2) and then subtract to find the difference.

It is also possible to view how heat can be transferred from one object to another. In a perfect system (where no heat is lost to the calorimeter) the heat lost by the hot object would be gained by the cold object such as

$$q_{\text{lost}} = -q_{\text{gained}} \quad \text{Equation 4}$$

which could be rearranged to give

$$q_{\text{hot water}} = -q_{\text{cold water}} \quad \text{Equation 5}$$

as well as replacing q with Equation 3 to give

$$mc(T_f - T_i)_{\text{hot object}} = -mc(T_f - T_i)_{\text{cold object}} \quad \text{Equation 6}$$

Although usually heat is lost to the calorimeter, as we showed in Equation 3, we can usually assume this value is exceptionally small and omit it from the equation as we did in Equations 4-6.

In this lab, we will be investigating the endothermic and exothermic qualities of salt solutions by dissolving various salts into water and monitoring the temperature. Then you will calibrate a coffee cup calorimeter by determining its heat capacity. Finally, you will calculate the specific heat of copper by dropping hot pennies into water and monitoring the heat exchange.

Materials

- Student tray containing the following:
 - 1-600 mL beaker
 - 1-250ml beaker
 - 1-100mL beaker
 - 1 50 mL graduated cylinder
 - 4-styrofoam cups
 - 2 Cardboard lid with hole
 - Stirring rod
 - 3 small test tubes
 - Container of NaCl
 - Container of KCl
 - Container of CaCl₂
 - ~10-15 pennies
 - 1 DI water bottle
- Student balance
- Logger Pro
 - Stainless steel temperature probe

Safety and Notes

It is important to follow the safety guidelines below while performing this lab.

- Remember to dry the calorimeter and probe between each trial.
- Use appropriate procedures for hot plate use.
- Dispose of waste in the CHM 111 Waste container.

Experimental Procedure and Data

Part A: Exothermic and Endothermic Dissolution of Salts

1. In this section of the procedure, you will observe temperature changes as various salts are dissolved in water. The first salt is NaCl, and the corresponding dissolution reaction is



2. You will be recording the temperature using the Logger Pro. Insert the end of the temperature probe into the bottom of a clean, dry test tube. (Use the small test tubes).
3. Record the mass of the empty test tube prior to beginning.
4. Fill the test tube approximately 2 cm with distilled water. Record the mass of the test tube and water to determine the mass of water.
5. Record the initial temperature of the water.
6. Record the mass of a second (clean and dry) test tube.
7. Fill the second test tube approximately 1 cm with solid NaCl.
8. Pour the solid NaCl into the water and stir gently with the temperature probe.
9. Monitor the temperature. The temperature will increase or decrease away from the initial temperature. Eventually the final temperature will be reached before it begins to return to room temperature. Record the final temperature.
10. The solution should be disposed of in the CHM 111 Waste container in the back hood. Rinse and dry the test tube.
11. Repeat Steps 2-7 for CaCl₂ and (if there is time) KCl.

Part B: Calculating the Heat Capacity of a Calorimeter

1. Obtain a hot plate and plug it in.
2. Stack the two Styrofoam cups together and place inside a 400 mL beaker
3. Place the assembly on a balance. Record the mass.
4. Place 50mL of tap water in the cup assembly and record mass. Subtract to find the mass of water.
5. Place the cardboard lid on top the Styrofoam cups.
6. Set up the Logger Pro with the stainless steel temperature probe. Insert probe through hole in lid. Obtain the initial temperature of the "cold" water.
7. Attach a thermometer clamp to the probe so that the probe is not sitting on bottom of cup.
8. This is our calorimeter.
9. Tare a 150mL beaker.
10. Add ~50mL of water to beaker on scale and record weight as the weight of "hot" water.
11. Place the ~150 mL beaker on the hot plate and heat to ~90 degrees.
12. Using the temperature probe, record the temperature of the hot water.
13. Tip the lid of the calorimeter up and using beaker tongs immediately pour the hot water into the calorimeter.
14. Immediately replace the lid and begin recording temperature on the Logger Pro.
15. Gently swirl the calorimeter and measure temperature every 10 seconds until the temperature is constant for 3 readings. (The temperature will increase to the final temperature before beginning to decrease back to room temperature.) Record the final temperature in the data section.
16. Repeat steps 9-14 for a second trial.

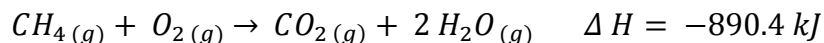
Part C: Calculating the Specific Heat of Copper

1. Obtain the mass of ~ 10-20 pennies. Record the mass in your data table.
2. Add pennies to a large (clean and dry) test tube. Place the test tube in a 400 mL beaker containing ~ 150 mL water. The water level should be above the level of the pennies to ensure they are adequately heated.
3. Heat water to ~ 95-100 degrees C. Record this as the initial temperature of the pennies. Allow the pennies to heat for ~2-3 minutes.
4. Add your calorimeter to the balance and tare it. Add ~40 mL of water. Record the mass of the water in the table.
5. Set up your calorimeter with temperature probe as described earlier.
6. Record the initial temperature of the water.
7. Swiftly but carefully remove the test tube from the hot water bath with a pair of test tube tongs. Dump the pennies into the calorimeter and immediately cover with the lid. *Be careful that the hot water on the outside of the test tube does not drip onto your hand or into the calorimeter.
8. Monitor the temperature by taking temperature readings every 10 seconds. When the temperature is consistent for 3 straight measurements (or the temperature begins to cool) record the final temperature of the pennies and water.
9. Repeat steps 1-8 for a second trial. (Be sure to dry the pennies between trials).
10. Use the data from your table to calculate the specific heat of copper (pennies).

Pre-lab Assignment and Questions

- *Note – This pre-lab must be completed before you come to lab.*

1. Given the balanced equation for the combustion of methane, calculate the amount of heat (q) produced by the combustion of 4.05 g CH_4 using equation 2. ($\Delta H_{\text{comb}} = -890.4 \text{ kJ}$). The formula for the combustion of methane.



2. Consider that the 4.05 g methane is burned and all of the heat from this combustion is absorbed by $1.0 \times 10^3 \text{ g}$ of 20.0°C water (which has a specific heat of $4.18 \text{ J g}^{-1}^\circ\text{C}^{-1}$). What would be the final temperature of the water?
3. A 62.5 gram sample of iron (with a heat capacity of $0.450 \text{ J/g}^\circ\text{C}$) is heated to 100.0°C . It is then transferred to a coffee cup calorimeter containing 52.7 g of water (specific heat of $4.184 \text{ J/g}^\circ\text{C}$) initially at 20.63°C . If the final temperature of the system is 29.59 , what was the heat capacity of the calorimeter?
4. A 17.5 g sample of metal heated in a test tube submerged in 100.0°C water. It was then placed directly into a coffee cup calorimeter holding 49.5 g of water at 21.6°C . The temperature of the water increased to 24.3°C , determine the specific heat capacity of the metal.
5. In the above problem, if the calorimeter's heat capacity is 21.3 J and we factored in this quantity to our calculations, what would have been the specific heat of the metal?

Name _____ Date _____
 Lab Partner's Name _____ Bin # _____

Experimental Data and Results

Part A: Exothermic and Endothermic Dissolution of Salts

	NaCl	CaCl₂	KCl
Equation for Dissolution of Salt			
Mass of Test Tube			
Mass of Test Tube and Water			
Mass of Water			
Initial Temperature of water			
Mass of Test Tube			
Mass of Test Tube and Salt			
Mass of the Salt			
Final Temperature			
Is this Exothermic or Endothermic Dissolution?			
Calculate q for the Dissolution of the Salt			
Calculate the J/g salt			
Would this salt be most useful for ice or heat packs?			

* Show your calculations.

Name _____ Date _____
 Lab Partner's Name _____ Bin # _____

Part B: Calculating the Heat Capacity of a Calorimeter

	Trial 1	Trial 2
Mass of Calorimeter (Coffee Cups, and Lid)		
Mass of "Cold" Water		
Initial Temperature of "Cold" Water		
Mass of Hot Water		
Initial Temperature of Hot Water		
Final Temperature of the System		
Specific Heat of Water	4.184 J/g°C	4.184 J/g°C
q Lost by Hot Water		
q Gained by Cold Water		
Heat Capacity of Calorimeter		
Average Heat Capacity of Calorimeter		

**Show your Calculations*

Trial 1				Trial 2			
Time	Temperature	Time	Temperature	Time	Temperature	Time	Temperature
0		40		0		40	
10		50		10		50	
20		60		20		60	
30		70		30		70	

Name _____ Date _____

Lab Partner's Name _____ Bin # _____

Part C: Calculating the Specific Heat of Copper

	Trial 1	Trial 2
Mass of Pennies Used		
Initial Temperature of Pennies		
Final Temperature of the System		
Mass of Water in Calorimeter		
Initial Temperature of Water in Calorimeter		
Specific Heat of Water	4.184 J/g°C	4.184 J/g°C
q Gained by Water		
Specific Heat of Copper (Assuming No Heat Lost to Calorimeter.)		
Average Specific Heat of Copper.		
Specific Heat of Copper (Using Calorimeter Specific Heat from Part B.)		
Average Specific Heat of Copper.		
If the specific heat of Cu is 0.386 J/g°C, what is the % error from the above row?		

* Show your calculations.

Trial 1				Trial 2			
Time	Temperature	Time	Temperature	Time	Temperature	Time	Temperature
0		40		0		40	
10		50		10		50	
20		60		20		60	
30		70		30		70	

Name _____ Date _____
Lab Partner's Name _____ Bin # _____

Results, Discussions and Post –Lab Questions

1. According to your results, what salt in Part A would have been the best choice for use in a heat pack? In a cold pack? Defend your choice.
2. Why is it not possible to reuse a heat pack in term of the chemistry?
3. Compare the specific heat of water to the specific heat of metal in the table provided. Which would heat up faster (with less energy required)?
4. Why would metal make a poor ingredient in a heat pack?
5. If there was a delay between measuring the initial temperature of a hot object and its transfer to the calorimeter, how would the heat capacity of the object be affected? How would the calculation of the heat capacity of the calorimeter be affected (too high, too low, or no affect)?
6. If hot water from the test tube in Part C had accidentally dripped into the calorimeter, how would the calculation of the specific heat of the calorimeter be affected (too high, too low, or no affect)?