Lab 9

Enthalpy and Specific Heat

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Learning Objectives

- Assemble and employ a calorimeter to determine the change in enthalpy
- Measure the specific heat of an unknown metal to determine its identity
- Observe the kinetic energy of molecules by comparing the temperature changes in a calorimeter

Introduction

Every substance has the ability to absorb heat. If enough heat is absorbed, the temperature of that substance will rise. The amount of heat required to raise the temperature varies by the type and amount of a substance because of the different chemical and physical properties that exist. **Specific heat capacity** is the amount of heat per unit mass required to raise the temperature of a substance by one degree Kelvin (Figure 1).

Figure 1: ***Water has a high specific heat capacity, so it requires more thermal energy to raise its temperature.***

The SI unit used to measure specific heat capacity is joules per kilogram Kelvin (J/kg K), although you may also see it expressed as calories per gram degrees Celsius (cal/g °C). The relationship between heat and temperature change can be described by the equations:

$$q = C_p \times \Delta T$$

$$q = m \times C_s \times \Delta T$$

where:

q = the heat being transferred to (or from) a substance

C (or C_p) = the heat capacity at a constant pressure

C_s = the specific heat capacity

m = the mass

ΔT = the temperature change

Every substance has a different specific heat capacity. For example, the specific heat of water is 1 calorie/gram °C, or, 4.186 joules/gram °C. This value is higher than most other common substances. As a result, water plays a very important role in temperature regulation. This is why water is used as a medium in calorimetry. Table 1 lists the specific heat capacity of different materials.

Table 1. Specific Heat Capacity

Substance	Specific Heat Capacity C_{sp} (J/g °C)
H_2O (l)	4.184
Ice at 0 °C	2.010
Steam at 100 °C	2.010
Aluminum	0.900
Chromium	0.448
Copper	0.385
Iron	0.444
Lead	0.160
Manganese	0.479
Magnesium	1.017
Tin	0.213
Zinc	0.388

Calorimetry

Calorimetry is a way to measure the heat that is generated or consumed by a substance during a chemical reaction or physical change. If heat is absorbed, it is an endothermic process. If heat is generated, it is an exothermic process. Most reactions involve some amount of heat transfer. Therefore, calorimetry has industrial applications in pharmacy, physiology, and other biological fields.

Modern calorimetry typically uses a **calorimeter**. Student labs often use a Styrofoam® calorimeter, due to its ability to minimize heat exchange with the outside environment and reliable insulation. This device is also used to contain the reaction and provide an environment with either constant pressure or constant volume. The heat capacity (C) of a calorimeter (the amount of heat required to raise the temperature of a calorimeter by one Kelvin) should be determined prior to the experiment. This can be done by transferring heat into a calorimeter and measuring how much the internal temperature increases. The formula for determining this value is:

$$C = \frac{\text{Energy input (joules)}}{\text{Increase in internal temperature (K)}}$$

Enthalpy

Enthalpy, abbreviated H, is the heat content of a chemical system. It is difficult to measure the total enthalpy. Therefore, the enthalpy change (ΔH) is more commonly calculated. Enthalpy change is the amount of heat generated or consumed when a chemical reaction occurs at a constant pressure. The change in enthalpy for a system is calculated by:

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

Note that ΔH is also equal to q (the energy added to or released from a system as heat).

The change in enthalpy, ΔH, is specified per mole of substance in the balanced chemical equation for the reaction. The units are typically given as kJ mol^{-1} (kJ/mol) or kcal mol^{-1} (kcal/mol). Energy changes are measured under standard laboratory temperature and pressure, defined as 25 °C and 1 atm.

Enthalpy change is positive if a reaction is endothermic (energy consuming) and negative if a reaction is exothermic (energy generating). In the experiment "specific heat of an unknown metal," the enthalpy of a reaction and the fusion of water will be determined (Figure 2).

Figure 2: ***An unknown type of metal will be provided for you to perform your first calorimetry experiment.***

Heat of Fusion

Similar to specific heat capacity, the amount of heat required to convert a substance from a solid to a liquid varies by the type and amount of the substance. However, the amount of heat required for a known type and amount of substance is a constant, calculated value known as the **heat of fusion**. Heat of fusion is often expressed as the amount of heat per gram or per mole of a certain substance. When it is referred to per mole, it is termed the molar heat of fusion. An example of a simplified equation for the molar heat of fusion using ice can be expressed as:

$$H_2O\ (s) + \text{Molar Heat of Fusion} \rightarrow H_2O\ (l)$$

To calculate molar heat of fusion, use the following equation:

$$q = \Delta H_{fus} m$$

where:

q = The amount of heat involved

ΔH_{fus} = The molar heat of fusion

m = The number of moles of a substance.

Imagine that a pre-weighed amount of metal is heated to a known temperature and is then quickly transferred into a calorimeter that contains a measured amount of water at a known temperature. Energy in the form of heat flows from the metal to the water, and the two eventually equilibrate at some temperature between the initial temperatures of the water and metal. Assuming that no heat is lost from the calorimeter to the surroundings, and that a negligible amount of energy is absorbed by the calorimeter walls, the amount of energy that flows from the metal as it cools is equal to the amount of energy absorbed by the water. In other words, the energy that the metal loses is equal to the energy that the water gains.

As discussed, when heat energy flows into a substance, the temperature of that substance will increase. The quantity of heat energy (q) required to cause a temperature change in any substance is equal to the specific heat capacity (C_{sp}) of that particular substance times the mass (m) of the substance times the temperature change (ΔT), as is given in the equation:

$$q = C \times m \times \Delta T$$

Since the metal loses energy (T_{final} is less than $T_{initial}$; therefore ΔT is negative because $\Delta T = T_{final} - T_{initial}$), q_{metal} is negative. The water in the calorimeter gains energy (T_{final} is greater than $T_{initial}$; therefore ΔT is positive); therefore q_{water} is positive. The following equation can be written, since the total energy is always conserved:

$$q_{metal} + q_{water} = 0$$

Rearranging this equation gives (note the negative sign):

$$q_{metal} = -q_{water}$$

Experiment 1 requires you to measure the mass of the water in the calorimeter, the mass of the unknown metal, and their initial and final temperatures (Figure 3). Using the equation:

$$q = C \times m \times \Delta T$$

the heat energy gained by the water and lost by the metal can be written as:

$$q_{water} = m_{water} \times C_{sp\,water} \times \Delta T_{water}$$

$$q_{metal} = m_{metal} \times C_{sp\,metal} \times \Delta T_{metal}$$

(Note that $\Delta T_{metal} < 0$ and $\Delta T_{water} > 0$, since $\Delta T = (T_{final} - T_{initial})$)

q_{water} can be calculated from the experimental data. The m_{water} and ΔT_{water} measurements will be measured, and the specific heat capacity of water is constant ($C_{sp\ water}$ = 4.184 J/g °C). The heat energy lost by the metal is equal (but opposite) to the heat energy gained by the water. To determine the specific heat capacity of your unknown metal ($C_{sp\ metal}$), substitute q_{metal} with $-q_{water}$. This equation can be written as:

$$C_{sp\,metal} = \frac{-q_{water}}{m_{metal} \times \Delta T_{metal}}$$

Three of the four variables are known: q_{water} is calculated from experimental data, and m_{metal} and ΔT_{metal} are measured in the experiment. Thus, you can solve for C_{metal}.

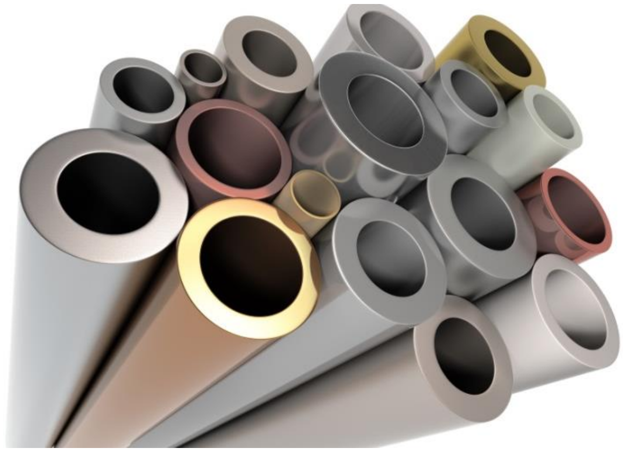

Figure 3: ***Heated metals will lose heat when put in a calorimeter. The lost heat is primarily absorbed (gained) by the water, while a small amount may also be absorbed by the calorimeter.***

Pre-Lab Questions

1. When 3.0 kg of water is warmed from 10 °C to 80 °C, how much heat energy is needed?

2. A calorimeter contains 100 g of water at 39.8 ºC. A 8.23 g object at 50 ºC is placed inside the calorimeter. When equilibrium has been reached, the new temperature of the water and metal object is 40 ºC. What type of metal is the object made from?

Lab 9 Enthalpy and Specific Heat

Experiment 1: Determination of Specific Heat of a Metal

In this experiment, you will measure and record the temperature over time to determine the specific heat and identity of an unknown metal.

Materials

(1) 500 mL Glass Beaker

(1) 100 mL Graduated Cylinder

Hot Pad

Scale

(2) 8 oz. Styrofoam® Cups

1 Styrofoam® Cup Lid

Test Tube Clamp

Thermometer

30 g Unknown Metal

*Camera/Smart phone is Sufficient

*Pot

*Stopwatch or Timer

*Stovetop or microwave

*Tap Water

* Wooden Toothpick (if using a microwave)

*You Must Provide

Procedure

1. Put on your safety glasses and gloves (provided in your safety box).

2. Use the 500 mL beaker to measure and pour 500 mL of tap water into a cooking pot. Place the pot on a stovetop and bring the water to a boil. Set a stopwatch or timer for three minutes. Once it has boiled for three minutes, reduce the heat from a boil to a simmer (Figure 4).

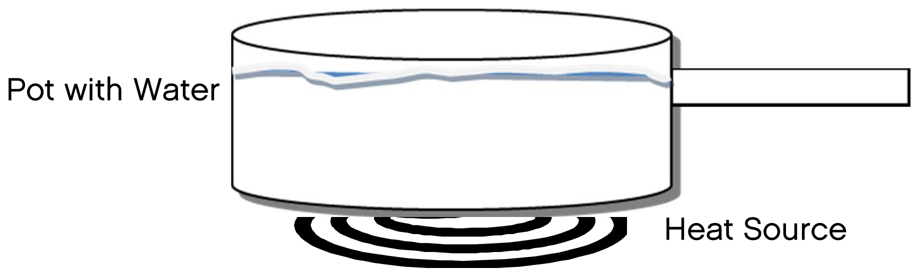

Figure 4. ***Sample set-up for Step 1. A pot with 500 mL of water rests on a stove top burner.***

Note: If you do not have a stove, heat the 500 mL of water to boiling in a microwave safe container. To avoid the risk of superheating water (water that does not look boiling, but explodes when agitated), place a microwave safe object, such as a wooden toothpick, in the water and heat the water at 1 minute increments. At the end of every minute, tap the microwave-safe container before removing the container from the microwave to test if your water is superheated. Superheated water can be extremely dangerous and can result in hospitalization. Use extreme caution if using the microwave method to boil water. Once boiling, carefully remove the container using the provided hot pad.

3. Using your 100 mL graduated cylinder, measure 50 mL of room temperature tap water into one of the Styrofoam® cups.

Note: The density of water is 1 g/mL. Therefore, 50 mL of water contains 50 g of water.

4. Record the mass of the water in Table 2.

5. Put the two Styrofoam® cups together by placing the Styrofoam® cup containing 50 mL of room temperature tap water inside the empty Styrofoam® cup.

6. Place your cups upright inside the 500 mL beaker. The beaker is used to provide vertical support for the Styrofoam® cups. See Figure 5 for reference.

Figure 5. ***Sample set-up for Step 5.***

7. Turn on the scale by pressing the button labeled 0/T. If your scale does not turn on, you may have to remove the battery cover and remove a small strip of plastic from the battery housing. Once the scale is on, press the 0/T button a second time to zero the scale. Make sure that the units are in grams (g). If not, press the M button until the units displayed are in grams.

8. Measure and record the mass of the unknown metal strip in Table 2.

9. Using the hot pad, pick up the unknown metal strip with the test tube clamp and hold the clamp and unknown metal in the simmering water for five minutes to ensure that the clamp and metal reach the same temperature as the water.

   Note: Do not drop the unknown metal into the pot of hot water and leave it there. It will be extremely difficult to retrieve the heated metal from the hot water, and any attempt to do so could cause severe burns.

10. Cover the cup containing 50 mL of tap water with the Styrofoam® cup lid.

11. Record the temperature of the water (this is the initial temperature) in Table 3 by inserting the thermometer through the hole in the lid. Use your camera (or smart phone) to take a picture of the thermometer at the initial temperature. Remove the thermometer after recording the temperature.

12. After five minutes have passed, quickly transfer the hot, unknown metal into the calorimeter.

Remember to quickly replace the lid. The metal strip will most likely stick out the side of the lid. This will not significantly affect your results.

13. Record the temperature of the water every minute for five minutes. Gently swirl the contents of the cup right before recording the temperature. Be careful to ensure that the thermometer is not touching the metal, only the water.

14. After your last measurement, use your camera (or smart phone) to take a picture of the thermometer. Be sure to correctly label your pictures and send them to your instructor with the post-lab questions.

15. Repeat Steps 1 - 13 two more times. Calculate the average heat capacity of the unknown metal. Begin by calculating q for water then use that information to calculate the heat capacity of the metal.

Table 2. Mass

	Mass (g)
Water	
Unknown Metal Strip	

Table 3. Specific Heat Data

Time (minutes)	Temperature (°C)		
	Trial 1	Trial 2	Trial 3
Initial			
5 minutes			
6 minutes			
7 minutes			
8 minutes			
9 minutes			
10 minutes			
Average Specific Heat Capacity of the Unknown Metal:			

Post-Lab Questions

1. Why is $\Delta T_{metal} < 0$?

2. Why is $\Delta T_{water} > 0$?

3. A metal sample weighing 43.5 g and at a temperature of 100.0 °C was placed in 39.9 g of water in a calorimeter at 25.1 °C. At equilibrium, the temperature of the water and metal was 33.5 °C. Determine the specific heat capacity of the metal.

4. What is the average specific heat capacity of the unknown metal in this experiment?

5. What is the unknown metal? Use Table 1 for reference.

Experiment 2: Cold Packs vs. Hand Warmers

In this experiment, you will observe the kinetic energy of molecules by comparing the temperature changes in a calorimeter. Specifically, you will observe the temperature changes for cold packs and hand warmers. Since temperature is defined as the average kinetic energy of the molecules, changes in temperature indicate changes in energy. You will use a Styrofoam® cup as a calorimeter to capture the energy. The customary lid will not be placed on the cup, since ample oxygen from the air is needed for the hand warmer ingredients to react within a reasonable amount of time.

Materials

Calorimeters (2 Styrofoam® cups)

1/4 Contents of a Cold Pack (ammonium nitrate), NH_4NO_3

Entire Contents of a Hand Warmer

(1) 10 mL Graduated Cylinder

Scale

Stir Rod

Spatula

Thermometer

Weigh Boat

*Access to Graphing Program

*Computer/Internet Access

*Distilled Water

*Paper Towels

*Scissors

*Stopwatch or Timer

*You Must Provide

Procedure

Part 1: Cold Pack

1. Put on your safety glasses and gloves (provided in your safety box).
2. Measure 10 mL of distilled water into a 10 mL graduated cylinder.
3. Turn on the scale by pressing the button labeled "0/T."
4. Place the weigh boat on the scale and press the 0/T button a second time to zero the scale. Make sure that the units are in grams (g). If not, press the M button until the units displayed are in grams.
5. Using your scissors, cut off the top of your cold pack and weigh out 10 g of the ammonium nitrate (NH_4NO_3) crystals found in the inner contents of the pack.
6. Place the 10 g of the ammonium nitrate into a Styrofoam® cup. The Styrofoam® cup will be used as a calorimeter.
7. Place a thermometer and a stir rod into the calorimeter (Styrofoam® cup).
8. Pour the 10 mL of distilled water into the calorimeter containing the ammonium nitrate (NH_4NO_3) taken from the cold pack.
9. Immediately record the temperature and time in Table 4.
10. Quickly begin stirring the contents in the calorimeter.
11. Use your stopwatch or timer to record the temperature at 30 second intervals in Table 4. You will need to stir the reaction the entire time you are recording data.
12. Collect data for at least five minutes, or until the temperature reaches its minimum and begins to rise. This should take approximately five to seven minutes.
13. Record the overall minimum temperature in the appropriate place in Table 4.

Part 2: Hand Warmer

1. The iron within the hand warmer mixes with the oxygen in the air and oxidizes in a process known as an oxidation reaction: $4\,Fe\,(s) + 3\,O_2\,(g) \rightarrow 2\,Fe_2O_3\,(s)$
2. Wash the thermometer and dry it with paper towels. Remember to rinse it with distilled water before drying. Carefully place the thermometer in the other Styrofoam® cup.
3. Cut open the inner package of the hand warmer and quickly transfer all of its contents into the calorimeter.
4. Immediately record the initial temperature of the contents and begin timing the reaction.

   Hint: Data collection should start quickly after the package is opened, because the reaction will be activated as soon as it is exposed to air.
5. Begin stirring the contents in the calorimeter with your thermometer.
6. Continue stirring, recording the temperature at thirty second intervals in Table 5. You will need to stir the reaction at a consistent rate the entire time you are recording data.
7. Set a stopwatch or timer for five minutes. Let the reaction continue for at least five minutes, or until the temperature has reached its maximum and then fallen a few degrees. This should take approximately 5 to 7 minutes.
8. Record the overall maximum temperature in the appropriate place in data Table 5.

Table 4. Cold Pack Data

Time (sec)	Temperature (˚C)
Initial	
30	
60	
90	
120	
150	
180	
210	
240	
270	
300	
330	
360	
390	
420	
450	

Table 5. Hand Warmer Data

Time (sec)	Temperature (˚C)
Initial	
30	
60	
90	
120	
150	
180	
210	
240	
270	
300	
330	
360	
390	
420	
450	

Post-Lab Questions

1. Graph the data from Tables 4 and 5 as two separate lines on the same graph and send the graph to your instructor. Construct your graph on a computer program such as Microsoft Excel®. If you do not have a graphing program installed on your computer, you can access one on the internet via the following links: http://nces.ed.gov/nceskids/createagraph/ or http://www.onlinecharttool.com

2. Calculate the overall temperature change for the cold pack (T_{max} - T_{min}). Show your work.

3. Calculate the overall temperature change for the hot pack. Show your work.

4. Which pack works by an exothermic process? Use experimental data to support your answer.

5. Which pack works by an endothermic process? Use experimental data to support your answer.

6. Write the balanced reaction between ammonium nitrate and water.

7. Which pack had the greatest change in enthalpy? How do you know?

8. Describe what would happen to the maximum temperature if the experiment used double the amount of hand warmers. Justify your answer.

9. Describe if and how the rate of temperature change would be affected if you crushed the cold pack crystals.

10. Explain how the law of conservation of energy was observed in this experiment.