

Energy Changes, Calorimetry and Specific Heat

Experiment #3

Objective: To measure heat energy emitted or absorbed in either a chemical process, such as dissolving a solid chemical in water, or in a purely physical process where a hot metal transfers heat to water.

Introduction

Calorimetry is the measure of a quantity of heat exchanged in chemical and physical processes. Heat is a form of energy associated with the motion of atoms or molecules of a substance and is measured in units of joules (SI units) or calories (1 calorie = 4.184 Joules). You are probably familiar with the term Calorie used by nutritionists, which is used for the energy content of foods, where 1 Calorie (with a capital C) = 1 kilocalorie (kcal) = 1000 calories (with a small c). In other words, nutritionists often ignore the correct terminology of kcal and use Calorie instead. Check the Nutrition Facts label on some foods or beverages to see what terminology they use. We will use calorie or kcal units in this experiment that applies calorimetry in water. One calorie is defined as the amount of heat needed to raise the temperature of one gram of water one degree Celsius.

In the first part of this experiment you will measure the amount of heat given off by a mass of hot metal when it is placed in cool water. This is much like the old western films showing the blacksmith beating a red hot piece of iron on an anvil and then putting it in a bucket of water to cool it. The specific heat of water is by definition 1 calorie per gram per degree Celsius (1 cal/g°C or 4.184 J/g°C). In this experiment you will determine the specific heat of at least two different metals and determine the type of metal based on its specific heat. The amount of heat a given mass of a metal absorbs when it is heated or releases as it cools is known as the specific heat of the metal, which varies inversely with the atomic mass of the metal. We will see how accurate your measurements are in the laboratory.

When an ionic compound dissolves in water, the water molecules interact with each ion and heat may be absorbed (endothermic) or released (exothermic). The heat exchanged when a substance dissolves in water is known as the heat of solution. Another term for this is the enthalpy of solution. In the second part of this experiment you will measure the amount of heat absorbed in an endothermic reaction and the amount of heat given off in an exothermic reaction by measuring heats of solution for two different solids. The solids you will be using are ammonium nitrate (NH_4NO_3), calcium chloride (CaCl_2).

The dissolution of ammonium nitrate in water takes place in cold packs that are used in sports medicine. The dissolution of calcium chloride is used to melt ice in the winter, which is more effective than the dissolution of regular salt (NaCl), which has a small endothermic heat of solution. These salts, when dissolved in water will lower the melting point of ice, so the ice will melt at a cooler temperature than normal.

When salts dissolve in water, the ionic bonds holding the salt molecules must first be broken, which requires energy being added to the molecules and absorbed from the surrounding (water). The ions can then bind with water, which is called solvation, and release energy. Each ion may bind with multiple water molecules, which can result in the release of much energy for some ions, such as calcium. The ammonium ion, on the other hand, does not bind as strongly to

water, so more energy is required to break the ionic bonds in ammonium nitrate than the amount of energy released when water molecules associate with those ions. When heat is given off in a process or reaction it is exothermic (heat going out to the surroundings) and the convention is to give the change in heat a negative sign, since heat is lost from the system and the surroundings become warmer. When heat is absorbed in a process it is endothermic and the change in heat has a positive sign because heat is added to the system and the surroundings become cooler.

Sample Calculations

A metal sample weighing 46.05 g was heated in boiling water to 99.0°C and quickly transferred to a styrofoam calorimeter cup containing 100.22 g of distilled water at a temperature of 28.7°C. The temperature of the water in the calorimeter increased and reached a maximum of 36.5°C after a couple of minutes. Calculate the specific heat of the metal in cal/g·°C.

The change in temperature of the water is: $\Delta T = 36.5^{\circ}\text{C} - 28.7^{\circ}\text{C} = 7.8^{\circ}\text{C}$

Heat gained by the water: $q_w = (m_w)(\text{sp ht}_w)(\Delta T_w) = (100.22 \text{ g})(1.00 \text{ cal/g}^{\circ}\text{C})(7.8^{\circ}\text{C})$
 $= 780 \text{ cal} = \text{heat lost by metal (2 sig figs)}$

Specific heat of metal: $\text{sp ht}_m = (\text{cal lost by metal})/(\text{g of metal})(\Delta T_m)$
 $= (780 \text{ cal})/(46.05 \text{ g})(99.0^{\circ}\text{C} - 28.7^{\circ}\text{C}) = 0.24 \text{ cal/g}^{\circ}\text{C} \text{ (2 sig figs)}$

If we look in Table 3.1, we see the specific heat of aluminum is 0.212 cal/g°C, which is closest to our answer, and so the metal is most likely aluminum.

When 5.12 g of anhydrous calcium chloride is dissolved in 50.15 g of water at 21.5°C in a styrofoam calorimeter cup, the temperature of the water increases to 35.8°C. Calculate the heat of solution of calcium chloride in water.

Moles of calcium chloride in 5.12 g $\text{CaCl}_2 = (5.12 \text{ g})(1 \text{ mole}/111 \text{ g}) = 0.0461 \text{ moles (3 sig figs)}$

The change in temperature of the water is: $\Delta T = 35.8^{\circ}\text{C} - 21.5^{\circ}\text{C} = 14.3^{\circ}\text{C} \text{ (3 sig figs)}$

Heat gained by the water: $q_w = (14.3^{\circ}\text{C})(1.00 \text{ cal/g}^{\circ}\text{C})(50.15 \text{ g} + 5.12 \text{ g}) = 792 \text{ cal (3 sig figs)}$
 Since the water gained heat when CaCl_2 was dissolved in it, the heat given off during the dissolution process would have the opposite sign, or a negative sign.

Heat of solution for $\text{CaCl}_2 = -792 \text{ cal}/0.0461 \text{ moles} = -17,200 \text{ cal/mol} = -17.2 \text{ kcal/mol}$

If we look in Table 3.2, we see the heat of solution for CaCl_2 is -17.4 kcal/mol, which is relatively close to the value measured. What might account for the small difference between the experimental result and the accepted value?

Materials

Double styrofoam cups for calorimeter, thermometer, unknown metal sample with wire

attached, glass stirring rod, 250 mL beaker, 400 mL beaker, 50 mL graduated cylinder, solid anhydrous calcium chloride, solid anhydrous ammonium nitrate, deionized water.

Procedure

A. Measuring the specific heat of a metal

1. Add about 125 mL of water to a 250 mL beaker and put it on a hot plate to boil. When the water begins to boil, you can set the dial of the hot plate high enough to keep the water boiling gently.
2. Obtain an unknown metal sample with a wire wrapped around it to suspend it in the water bath. Accurately weigh the metal sample to two decimal places (0.01 g) and record the mass on the report sheet.
3. After weighing it, wrap the wire around a glass stirring rod so the length of wire is appropriate to suspend the metal in the boiling water bath. Place the glass rod on the rim of the beaker to suspend the metal in the water with the piece of metal completely submerged in the water but not in contact with the bottom of the beaker.
4. Record the temperature of the boiling water as accurately as possible with your thermometer after it has begun boiling. Do not allow the thermometer to make contact with the bottom of the beaker, which may be hotter than the water in the beaker, since it is in contact with the hot plate. If the temperature is not close to 100°C (at least 98°C) you should notify the instructor.
5. Obtain a double styrofoam cup with a lid to use as your calorimeter. Weigh the empty cups (keep them together throughout the experiment) with lid on it and record the mass to two decimal places (the nearest 0.01 g) on the report sheet.
6. Use your 50 mL graduated cylinder to measure exactly 100 mL of deionized water into the styrofoam cup and reweigh the cup with 100 mL water and lid on it. Record the mass to two decimal places on the report sheet. Note the exact mass of the 100 mL of water you just measured with the graduated cylinder. If the mass of water is more than 1 gram different from 100 g, you did not measure the water very carefully with the graduated cylinder.
7. Place the double styrofoam cup in a 400 mL beaker to make it more stable and place your thermometer in the cup with the lid on it to measure the temperature of the deionized water. The temperature should remain stable (no change) for at least 2 minutes. Record the initial temperature of the water in the calorimeter on the report sheet to the nearest 0.1°C (initial means before adding the hot metal).
8. If the hot water bath with the metal suspended in it has been boiling for at least 2 minutes, you can quickly transfer the metal to the styrofoam cup by holding the glass stirring rod and sliding the wire off. **Do not put the glass rod in the styrofoam cup**, but the thin wire on the metal will be in there. The thermometer should be in the cup with the lid on it and watch the temperature of the water in the cup as it absorbs the heat from the metal. Stir the water with the thermometer but

keep the lid on it to avoid heat loss.

9. When the temperature does not change any more, record the maximum final temperature as accurately as possible on the report sheet. It should take a few minutes for the temperature of the metal and water to equilibrate.
10. Calculate the specific heat of the metal from the data you have collected and determine which of the metals you probably have from the specific heat values in Table 3.1 at the end of the Procedure section.
11. Remove the metal from the calorimeter cup and return it to the boiling water bath to repeat the data collection for this metal. You do not have to change the water in the styrofoam cup, but you should reweigh the cup with water in it to make sure it has the same mass of water and record the mass on the report sheet for trial 2.
12. Allow the metal to remain in the boiling water bath for several minutes (3 to 5 min) and get an accurate measure of the water temperature in the hot water bath and in the styrofoam cup just before transferring the metal back to the calorimeter and record the two temperatures on the report sheet for trial 2 to the nearest 0.1°C.

B. Measuring heats of solution for two salts

1. Weigh the empty styrofoam calorimetry cups (two together) with lid and record the mass on the report sheet to two decimal places.
2. Add exactly 50.0 mL of deionized water to the styrofoam cup and reweigh it with lid, recording the mass on the report sheet to two decimal places. How does the mass of water measured compare to the volume of water added to the cup?
3. Record the initial temperature of the water in the calorimeter after the thermometer has stabilized for at least a minute.
4. Accurately weigh about 15 grams (± 2 g) of ammonium nitrate (NH_4NO_3) to two decimal places and record the mass on the report sheet.
5. Transfer the ammonium nitrate to the calorimeter and place the lid on it with the thermometer in it. Swirl the cup gently (DO NOT SPILL ANY WATER FROM THE CUP) or stir the solution with the thermometer until the ammonium nitrate is completely dissolved. Keep an eye on the thermometer and record the final minimum temperature the solution reaches on the report sheet. After you record the final temperature, you may want to stick your finger into the solution to see how cold it got.
6. Calculate the number of calories consumed in the dissolution of this salt in water using 0.85 cal/g°C as the specific heat of the ammonium nitrate solution. You will also have to use the sum

of the mass of water and mass of ammonium nitrate for the total mass of the solution and ΔT is the change in temperature as the salt dissolves. Record the answer in the table on the report sheet.

$$\text{cal} = (\Delta T)(m_{\text{water}} + m_{\text{salt}})(0.85 \text{ cal/g}^\circ\text{C})$$

7. Calculate the number of moles of ammonium nitrate used for this dissolution and write the answer in the table on the report sheet.:

$$\text{moles NH}_4\text{NO}_3 = \frac{\text{grams NH}_4\text{NO}_3}{\text{molecular weight of NH}_4\text{NO}_3}$$

8. Calculate the heat of solution for ammonium nitrate in water and record it on the report sheet:

$$\text{Heat of Solution} = \text{cal/mol} \quad \text{or divide the answer in step 6 by the answer in step 7}$$

9. How does your answer compare with the value in Table 3.2 at the end of the Procedure on the next page? Note that the values in the table are in kcal/mole and not cal/mole.

You can dispose of the solution in the styrofoam cup in the sink.

10. Repeat steps 1 through 9 for calcium chloride, except you will accurately weigh about 10 g (± 2 g) of calcium chloride to two decimal places. You will still use 50 mL of clean deionized water in the calorimeter to dissolve the calcium chloride salt. In this case the temperature will increase and you will record the maximum temperature the solution reaches. Stick your finger in the solution after you record the final temperature.
11. For step 7 you will use the grams of CaCl_2 added to the water and the molecular weight of the CaCl_2 to calculate the moles of calcium chloride.

Table 3.1. Specific Heat of Selected Metals

Metal	Specific Heat (cal/g $^\circ$ C)
Aluminum	0.212
Copper	0.090
Lead	0.030
Steel	0.110

Table 3.2. Heat of Solution of Two Salts

Salt	Heat of Solution
NH_4NO_3	+6.2 kcal/mol
CaCl_2	-17.4 kcal/mol

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Energy Changes, Calorimetry and Specific Heat**Experiment #3****Prelab Exercise**

1. Calculate the molecular weight of calcium chloride (CaCl_2) (show work). Enter this number at the top of the table in the lab report for part B.
2. Calculate the molecular weight of ammonium nitrate (NH_4NO_3) (show work). Enter this number at the top of the table in the lab report for part B.
3. A piece of metal weighing 35.5 g was heated to 90°C and transferred to a calorimeter cup containing 80.0 g of deionized water at a temperature of 22.5°C . The temperature of the water in the calorimeter increased to 28.0°C . Calculate the specific heat of this metal.

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4. When 2.0 g of NaOH was dissolved in 50.0 mL of water at 22.2°C in a calorimeter cup, the temperature of the water increased to 32.8°C. Calculate the heat of solution for sodium hydroxide in kcal/mole. Assume the density of the water is 1.00 g/mL and the specific heat of the solution is 1.00 cal/g. Notice that the mass of the solution is the sum of the masses of water and sodium hydroxide.

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Energy Changes, Calorimetry and Specific Heat**Experiment #3****Data & Report Sheet****Part A. Measuring Specific Heat of a Metal**Mass of **empty** styrofoam calorimeter cup and lid: _____ g

Use this mass of the empty cup to determine the mass of water in the cup in the table below. Use the proper number of significant figures throughout the report sheet.

	Unknown Metal	
a) Mass of unknown metal	g	
	Trial 1	Trial 2
b) Temperature of boiling water bath	°C	°C
c) Mass of cup + lid + 100 mL of water	g	g
d) Mass of water in cup	g	g
e) Initial temperature (of calorimeter)	°C	°C
f) Final temperature (of calorimeter)	°C	°C
g) Change in temperature (of calorimeter) [f - e]	°C	°C
h) Calories released to water in calorimeter [see Introduction: Sample Calculation]	cal	cal
i) Change in temperature of the metal [b - f]	°C	°C
j) Specific heat of metal in cal/g°C [h/(a · i)]		
Average Specific Heat	cal/g°C	

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A-1. Using the data for specific heats of selected metals in Table 3.1 at the end of the procedure, indicate what metal the unknown sample you tested would be.

Unknown Metal is:

A-2. If your measurements do not match very closely the specific heat for any of the metals listed in Table 3.1, indicate why you selected one metal vs another for your answer above.

A-3. If your values for specific heat in the two trials for either metal are much different from one another, try to explain what may have caused the discrepancy.

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Part B. Measuring heats of solution for two saltsMass of **empty** styrofoam calorimeter cup and lid: _____ g

You will use the mass of the empty cup to determine the mass of water in the cup for each trial below. Use the proper number of significant figures for all data and answers.

	NH ₄ NO ₃	CaCl ₂
(a) Molecular weight (MW)	g/mole	g/mole
(b) Mass of salt used	g	g
(c) Moles of salt used (g/MW) [b/a]	moles	moles
(d) Mass of cup + lid + 50 mL water	g	g
(e) Mass of water in cup	g	g
(f) Initial Temperature	°C	°C
(g) Final Temperature	°C	°C
(h) Change in Temperature [f–g]	°C	°C
(i) Calories (consumed or released) [Introduction: Sample Calculation]	cal	cal
Heat of Solution* (Note cal)	cal/mole	cal/mole
Heat of Solution* (Note kcal)	kcal/mole	kcal/mole

* Be sure to give the proper sign (+ or –) for the heat of solution, depending on whether the dissolution of the salt is releasing heat to the surroundings (and going to a lower energy for the dissolved salt) or absorbing heat from the surroundings and going to a greater energy for the dissolved salt).

B-1. Compare your values for heat of solution with the accepted values given in Table 3.2 at the end of the Procedure section. Are any of your values more than 10 percent different from the accepted value? If so, which one(s)?

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B-2. You should discuss what factors may account for any large (more than 10 percent) differences between your measured value and the accepted values. [Note: the salts of NH_4NO_3 and CaCl_2 may absorb moisture from the atmosphere while standing, which would mean part of the dissolution reaction has already taken place].

B-3. How would you explain to a friend why dissolving ammonium nitrate in water causes the water to get cold, whereas dissolving calcium chloride in the water causes the water to get warm. You should discuss this in terms of breaking bonds that hold the salt molecules together and the formation of new bonds when water dissolves the salt. [See introduction]