



Coffee Cup Calorimetry

Measuring Heat Transfer

MATERIALS AND RESOURCES

EACH GROUP

balance	hydrochloric acid, 0.5 <i>M</i>
data collection device	sodium hydroxide, 0.5 <i>M</i>
graduated cylinder, 100 mL	2 cups, 8-oz Styrofoam [®] , with lids
sensor, temperature, stainless steel	spoon, plastic
weigh boat	water, distilled
ammonium chloride	
calcium turnings	

ABOUT THIS LESSON

In this lesson, students will evaluate a sample investigation to review concepts and techniques required for quantifying heat exchange in a high school laboratory. With some guidance, students will design calorimetry investigations for the physical process of an endothermic dissolution and two exothermic chemical reactions.

OBJECTIVES

Students will:

- Review and evaluate sample procedures and data analysis for the heat exchange between metal and water in a coffee cup calorimeter
- Design calorimetric investigations using parameters and assumptions for physical and chemical systems
- Organize and analyze data to determine enthalpy changes for the processes carried out in their coffee cup calorimeters

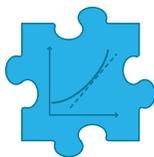
LEVEL

Chemistry

NEXT GENERATION SCIENCE STANDARDS



PLANNING/CARRYING OUT
INVESTIGATIONS



ANALYZING AND
INTERPRETING DATA



CONSTRUCTING EXPLANATIONS
DESIGNING SOLUTIONS



ENERGY AND MATTER



SCALE, PROPORTION,
AND QUANTITY



PS3: ENERGY

ASSESSMENTS

The following types of formative assessments are embedded in this lesson:

- Assessment of student understanding through discussion of Pre-Lab Exercises regarding the sample investigation
- Assessment of student-developed procedures and laboratory technique through active monitoring and questioning during investigation

ACKNOWLEDGEMENTS

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COMMON CORE STATE STANDARDS

(LITERACY) RST.9-10.4

Determine the meaning of symbols, key terms, and other domain-specific words and phrases as they are used in a specific scientific or technical context relevant to grades 9–10 texts and topics.

(LITERACY) W.1

Write arguments to support claims in an analysis of substantive topics or texts, using valid reasoning and relevant and sufficient evidence.

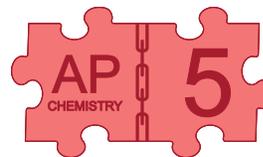
(MATH) A-CED.4

Rearrange formulas to highlight a quantity of interest, using the same reasoning as in solving equations.

(MATH) N-Q.1

Use units as a way to understand problems and to guide the solution of multi-step problems; choose and interpret units consistently in formulas; choose and interpret the scale and the origin in graphs and data displays.

CONNECTIONS TO AP*



- A.2** The process of kinetic energy transfer at the particulate scale is referred to in this course as heat transfer, and the spontaneous direction of the transfer is always from a hot to a cold body.
- B.1** Energy is transferred between systems either through heat transfer or through one system doing work on the other system.
- B.2** When two systems are in contact with each other and are otherwise isolated, the energy that comes out of one system is equal to the energy that goes into the other system. The combined energy of the two systems remains fixed. Energy transfer can occur through either heat exchange or work.
- B.3** Chemical systems undergo three main processes that change their energy; heating/cooling, phase transitions, and chemical reactions.
- B.4** Calorimetry is an experimental technique that is used to measure the change in energy of a chemical system.

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TEACHING SUGGESTIONS

Calorimetry	
Prerequisite Knowledge and Skills	
Understanding the relationship between mass, heat, and temperature	
Solving for variables using the equation,	
$q = mC\Delta T$	(Eq. A)
Differentiating between <i>exothermic</i> and <i>endothermic</i> processes	
Working with thermochemical equations to calculate heat or chemical amounts	
Relating heat transfer to the law of conservation of energy	

According to the law of conservation of energy,

$$0 = q_{\text{gained}} + q_{\text{lost}} \quad (\text{Eq. B})$$

Manipulating the equation, the relationship between heat lost by one component of a system and the heat gained by another becomes

$$-q_{\text{lost}} = q_{\text{gained}} \quad (\text{Eq. C})$$

This equation uses the same convention established in other lessons like “Mass, Heat, and Temperature” and “Burn, Baby, Burn.” From the perspective of the system, the amount of heat gained (q_{gained}) is always a positive value. When values are substituted into the equation the directionality of heat lost (q_{lost}), always a

negative value, is resolved and the magnitudes become equal.

For the purposes of vertical alignment, check with the physics teachers to make sure that students are prepared for other variations of the heat variable. It is often q in chemistry but Q in physics. Students are trained that capitalization changes the variable entirely, such as with t (time) and T (temperature). If there is a difference between disciplines, make note of it with your students.

Note that coffee cup calorimeters are not perfectly sealed and therefore maintain pressures equal to the barometric pressure in the room. The heat transfer, q , is equal to the change in enthalpy, ΔH , for processes at constant pressure. Textbooks that introduce *bomb calorimetry*, referred to as *constant-volume calorimetry*, equate heat transfer to an overall change in energy, ΔE . For the purposes of a first-year course, the approach to problem solving is the same for both types of calorimeters. The distinction between ΔE and ΔH , based on technical definitions for enthalpy, energy and work, is beyond the scope of the course.

In the lab, students are asked to design their own procedures for three calorimetry investigations. A sample investigation—complete with procedures, data, and analysis—is presented to scaffold understanding of the techniques used in coffee cup calorimetry. Pre-Lab Exercises highlight essential conceptual and procedural points.

TEACHING SUGGESTIONS (CONTINUED)

In discussing the sample investigation, the following ideas should surface:

- Heat is transferred between two objects at different temperatures in the direction of “hot” to “cold.”
- If the temperature increases for a sample of matter, it *gained* thermal energy from another sample of matter or exothermic process.
- If the temperature decreases for a sample of matter, it *lost* thermal energy to another sample of matter or endothermic processes.
- The copper metal is at thermal equilibrium with the boiling water before it is transferred to the calorimeter, therefore we can use the boiling temperature of water as the initial temperature of the copper.
- Two substances at different temperatures do not always reach thermal equilibrium by experiencing equal changes in temperature.
- When equating two heat values, the units must be the same.
- Equation A is useful for calculating the heat lost or gained by a sample of matter with a measurable temperature change.
- The change in temperature is defined as

$$\Delta T = T_f - T_i \quad (\text{Eq. D})$$

- The specific heat capacity for water is $4.184 \text{ J g}^{-1} \cdot ^\circ\text{C}^{-1}$ and the density is $1 \text{ g} \cdot \text{mL}^{-1}$.
- The hot plate was required in the sample investigation to create a temperature differential between the metal and the water because no physical or chemical change occurs when the metal is placed in the water. The other investigations do not require an external heat source.

As students design their investigations, monitor their progress and procedural designs. If it is not logistically feasible to check off the procedure for each group prior to investigation, encourage groups to compare procedures. If you notice that a group is struggling with design or experimentation, discuss the errors/issues and ask them to begin a new trial.

Although most calorimetry calculations can look very different in the details of numbers and units, it is useful for students to see that a consistent conceptual approach can be applied to the analysis of heat transfer. The example calculations shown emphasize the following recurring steps in solving calorimetry problems.

1. Always start with Equation C. This equation is the entire basis of heat transfer and, though seemingly unnecessary, grounds students in the conceptual framework of the problem.
2. Identify the components of the system that are losing and gaining energy. Rewrite Equation C, including the specific components in the problem.
3. In simple calorimetry problems, heat gained by the calorimeter is often neglected. If the heat absorbed by the calorimeter is included in the calculations, this value should be added to the heat absorbed by the system.

Establish the unknown variable, making sure to include the appropriate units, and attribute it to a component. The variable being solved for can range from a specific heat capacity to a change in enthalpy to a final temperature. In this investigation, the units for the variable are clearly stated.

TEACHING SUGGESTIONS (CONTINUED)

4. Solve for the amount of heat transferred for all components of the system individually. Dimensional analysis can clearly show how units cancel until only an energy unit is left. In the analyses shown, heat is always calculated in joules for consistency. The variable, x , should fit into the calculation.

The way students label their components can tell them about the nature of the heat calculations they must complete. Equation A is applied in situations like q_{water} , q_{solution} , q_{metal} and other discrete samples of matter. Process enthalpies are applied in situations like $q_{\text{dissolution}}$, $q_{\text{neutralization}}$, q_{reaction} , and for other terms describing a physical or chemical change. For all processes, a limiting reactant must be used in the calculation.

5. Plug the values into the equation established in Step 2, being sure to multiply the q_{lost} component by (-1) . Solve for x . Despite the unit cancelation shown in the dimensional analysis, the units of x do not change.

PART I: HEAT OF SOLUTION

Part I requires the students to write a balanced equation for the endothermic dissolution process of ammonium chloride. Some students may be tempted to represent a chemical reaction where water is a reactant. However, in this case the process represents a physical change where the solid dissociates into aqueous ions. If students are struggling with this equation direct them to Pre-Lab Exercise 9, which gives the dissolution equation for sodium hydroxide.

Notes:

- The solution experiences a decrease in temperature, and must be losing energy. Deductive reasoning states that the process of dissolution must be gaining that energy.
- There is a difference between the physical “solution” and the process of “dissolution” between which students must distinguish.
- The terms q_{solution} and $q_{\text{dissolution}}$ can help students struggling with appropriate notation.
- The variable being solved for is established as $\Delta H_{\text{soln}}^{\circ}$ with units of $\text{kJ}\cdot\text{mol}^{-1}$ of ammonium chloride.
- Be sure that students add the mass of the ammonium chloride to the mass of the water when determining the heat change for the solution, as specified in the assumptions section.

TEACHING SUGGESTIONS (CONTINUED)

PART II: HEAT OF NEUTRALIZATION

Part II deals with the exothermic neutralization reaction of a strong acid with a strong base. The procedures work best when one solution is already in the calorimeter. Soon after the start of data collection, the other solution can be poured into the calorimeter. The reaction continues to completion but swirling as described in the sample procedures is recommended to make sure the temperature change is uniform.

Students must be careful if they use different volumes of the equimolar solutions, as they will then need to clearly establish the limiting reactant.

Notes:

- The temperature increase of the solution indicates it is gaining energy. Deductive reasoning states that the neutralization process must be losing that energy.
- The terms $q_{\text{neutralization}}$ and q_{solution} can help students struggling with appropriate notation.
- The variable being solved for is established as $\Delta H_{\text{neut}}^{\circ}$ with units of $\text{kJ}\cdot\text{mol}^{-1}_{\text{rxn}}$.
- The sodium hydroxide is the limiting reactant, so it is used in the thermochemical calculation for $q_{\text{neutralization}}$.
- Some students may want to use both volumes to calculate $q_{\text{neutralization}}$. Redirect them to determine the limiting reactant. Although the volume of excess reactant is necessary for the total mass of solution, it is not appropriate for establishing the heat change.
- The coefficients in the balanced equation provide the ratio between moles of reactant and moles of reaction.
- Be sure students use the total mass of the solution in their calculation, as specified in the assumptions section.

PART III: HEAT OF REACTION

Part III involves the exothermic single replacement reaction of calcium with water. Static electricity builds up easily in the bottle of calcium turnings. Pour some calcium samples into a beaker or weigh boat to avoid students getting shocked when using a metal scoopula to obtain a sample. Students must obtain a mass for their calcium sample, as it is the limiting reactant in the process.

Notes:

- The temperature increase of the solution indicates it is gaining energy. Deductive reasoning states that the reaction must be losing that energy.
- The terms q_{reaction} and q_{solution} can help students struggling with appropriate notation.
- The variable being solved for is established as $\Delta H_{\text{rxn}}^{\circ}$ with units of $\text{kJ}\cdot\text{mol}^{-1}_{\text{rxn}}$.
- The coefficients in the balanced equation provide the ratio between moles of reactant and moles of reaction.
- Be sure students add the mass of calcium to the mass of water, as specified in the assumptions.

TEACHING SUGGESTIONS (CONTINUED)

CLAIM-EVIDENCE-REASONING (C-E-R)

Conclusion Questions 2–5 are designed as error-analysis questions. The *Claim-Evidence-Reasoning* (C-E-R) framework for crafting scientific explanations is appropriate for formulating responses to each question.

- The *Claim* is the student’s answer to the question in simple, straightforward terms. It is where most students stop if not prompted to go further. A claim states what the student believes to be true but does not examine the evidence or reasoning behind the statement.

In the Conclusion Questions, the claim regards the effect an error will have on a calculated value (as compared to a theoretical value) and will state either “too high,” “too low,” or “the same.”

- In the context of a laboratory investigation the *Evidence* generally cites numbers, calculations, observations, or other collected data that support the student’s claim. In most cases, the evidence is a subset of all the data collected and represents just the portion of the data that relates to the claim. Evidence statements might also be derived from background reading, the procedures, prior knowledge, or other citable sources of objective information. The evidence is always absent of inference.

Note that for this particular investigation the evidence component is expected to be less

substantial than the reasoning, as students must logic through propagation of error in their calculations.

- The *Reasoning* statement is typically the most difficult for students to articulate. Proficient reasoning requires the selection of appropriate and adequate evidence as a subset of all the data provided. The heart of the response, the reasoning, is the precision with which the evidence is connected to the claim. The reasoning statement should justify why the data presented counts as evidence for the claim and include applicable scientific principles and knowledge. The reasoning statement weaves together background information, facts, and theories that justify the claim and the selection of evidence.

Conclusion Question 6 requires students to look up standard heats of formation for all reactants and products in each investigation. Thermodynamic data exists for different phases of pure substances, including aqueous, so students should select their data carefully. As an example the calcium hydroxide formed in Part III is aqueous and not solid.

In the Going Further section, most students will find a procedure that requires ice. This connects latent heat of fusion, explored in “Heating Curves,” with calorimetry. Students may need clarification on the difference between *specific heat* and *heat capacity*.

PRE-LAB EXERCISES

1. Sketch a diagram for the setup described in the Sample Procedure.

See Figure A.

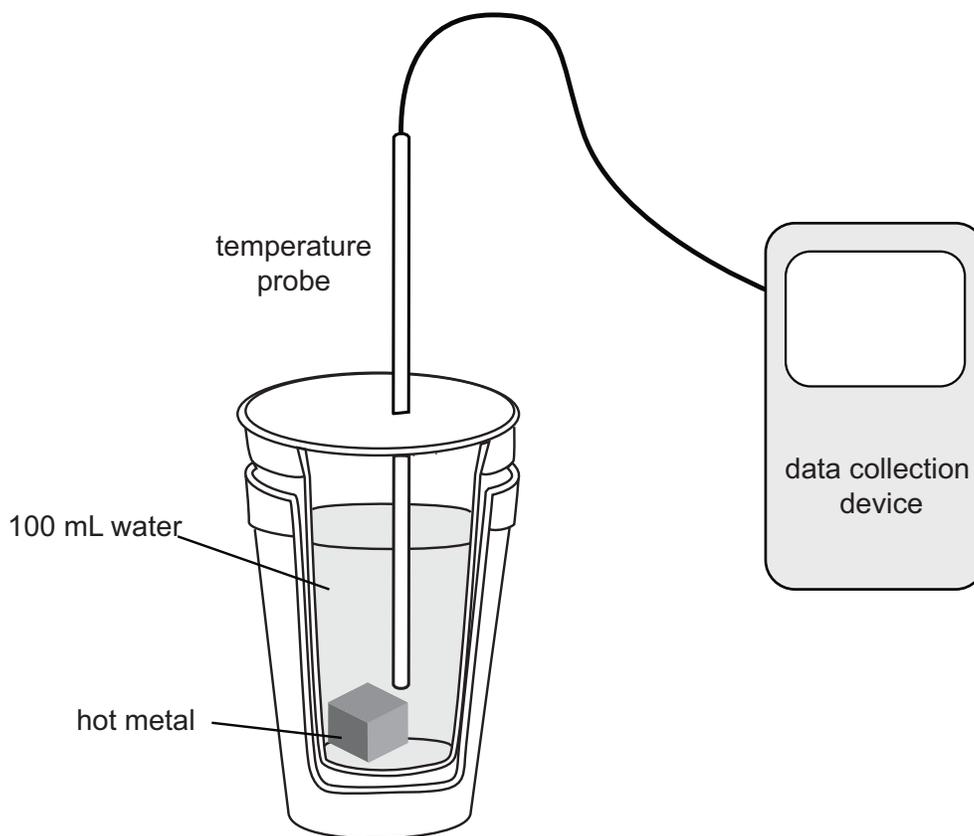


Figure A. Calorimeter setup

2. Describe the major safety concerns for the Sample Procedure.
 - Avoid touching the hot plate with hands or cords.
 - Place the metal sample carefully into the boiling water to avoid splashing hot water and to avoid cracking the glass beaker.
 - Remember that hot and cold glass look the same.

3. Consider the transition from Equation 2 to Equation 3 of the Sample Analysis. Explain how the experimenter can determine which system loses heat and which system gains heat based on the data.

The temperature changes for the copper and water indicate whether they lost or gained heat. The copper metal has an initial temperature equal to the boiling water, 99.8°C . The initial temperature of water was recorded to be 22.3°C . The final temperature of the system was recorded as 24.1°C . As the temperature of the copper went down, heat must have been lost. Therefore the water gains that heat, which increased its temperature.

PRE-LAB EXERCISES (CONTINUED)

4. State what the variable x represents in Equation 4 of the Sample Analysis, and report its units.

The x represents the value for the specific heat capacity of the metal (refer to Equation 4). The units are $\text{J/g}\cdot^{\circ}\text{C}$.

5. In Equation 4 of the Sample Analysis, why is the temperature of the boiling water used as the initial temperature of the copper?

The metal was allowed to sit in the boiling water for a period of time. Metals absorb heat quickly, so the copper would have reached thermal equilibrium with the water and have been at the same temperature initially.

6. Explain why the final temperature of the sample is closer to 22.3°C rather than 99.8°C . Why are the final temperatures of water and metal the same?

The final temperature for the copper and water is 24.1°C when they reach thermal equilibrium. The mass of the water in the sample is greater than the mass of the copper used, so the heat lost by the metal is spread out over more matter.

Also, the specific heat calculated for the metal is much smaller than the given specific heat for the water. The metal changes temperature more readily than the water when heat is lost or gained.

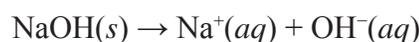
7. Justify why the units from Equation 4 and Equation 5 of the Sample Analysis must be the same.

Equation 4 represents the amount of energy lost by the copper and Equation 5 represents the amount of energy gained by the water. The energy lost equals the energy gained, as shown in Equation 6. If the units are not the same, then the equation does not work and variables cannot be calculated correctly.

8. Considering the actual data collected, state the assumption made in Equation 5 of the Sample Analysis.

The assumption is that the density of water is 1 g/mL , which allows us to use the volume of water as its mass.

9. The heat calculation, q , differs from the sample analysis when the system losing or gaining heat is a process for which a heat capacity and temperature change are not applicable. Calculate the heat transferred by the dissolution of 3.0 g of $\text{NaOH}(s)$ in water.



$$\Delta H^{\circ}_{\text{soln}} = -44.5\text{ kJ}\cdot\text{mol}^{-1}$$

$$3.0\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.01\text{ g NaOH}} \times \frac{1\text{ mol rxn}}{1\text{ mol NaOH}} \times \left(\frac{-44.5\text{ kJ}}{1\text{ mol rxn}} \right) = -3.34\text{ kJ}$$

DATA AND OBSERVATIONS

PART I: HEAT OF SOLUTION

BALANCED EQUATION FOR THE DISSOLUTION



SUMMARY OF PROCEDURES

- Measure out ammonium chloride. Record the mass in the data table.
- Obtain two stacked coffee cups and a fitted lid. Transfer a known volume of water into the cups, replace the lid, and record the initial temperature with the temperature probe.
- Set up the data collection device to record the temperature every second for 180 seconds.
- Add the solid ammonium chloride, stir, and continue recording the temperature until the data collection device stops.

DATA TABLE

Table A. Heat of Solution	
Mass of ammonium chloride	1.30 g
Volume of water	65.5 mL
T_i of water in calorimeter	22.1°C
T_f of water in calorimeter	21.0°C

ANALYSIS

$$-q_{\text{lost}} = q_{\text{gained}}$$

$$-q_{\text{solution}} = q_{\text{dissolution}}$$

$$q_{\text{solution}} = (65.5 + 1.30)\text{g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (21.0 - 22.1)^\circ\text{C} = -307.4 \text{ J}$$

$$q_{\text{dissolution}} = 1.30 \text{ g} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.5 \text{ g NH}_4\text{Cl}} \times \frac{x \text{ kJ}}{1 \text{ mol NH}_4\text{Cl}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 24.30x \text{ J}$$

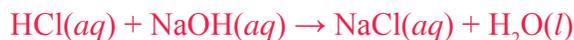
$$-(-307.4) = 24.30x$$

$$x = \frac{307.4}{24.30} = 13 \frac{\text{kJ}}{\text{mol NH}_4\text{Cl}}$$

DATA AND OBSERVATIONS (CONTINUED)

PART II: HEAT OF NEUTRALIZATION

BALANCED EQUATION FOR THE NEUTRALIZATION



or



DATA TABLE

Table B. Heat of Neutralization	
Volume of 0.5 M HCl	30.0 mL
Volume of 0.5 M NaOH	28.4 mL
T_i of solution	22.1°C
T_f of solution	25.1°C

SUMMARY OF PROCEDURES

- Rinse and dry the cups used in Part I.
- Transfer a known volume of hydrochloric acid into the cups, replace the lid, and record the initial temperature with the temperature probe.
- Set up the data collection device to record the temperature every second for 180 seconds.
- Add a known volume of sodium hydroxide, stir, and continue recording the temperature until the data collection device stops.

ANALYSIS

$$-q_{\text{lost}} = q_{\text{gained}}$$

$$-q_{\text{neutralization}} = q_{\text{solution}}$$

$$q_{\text{neutralization}} = 28.4 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.5 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol rxn}}{1 \text{ mol NaOH}} \times \frac{x \text{ kJ}}{1 \text{ mol rxn}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 14.2x \text{ J}$$

$$q_{\text{solution}} = (30.0 + 28.4) \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (25.1 - 22.1)^\circ\text{C} = 733 \text{ J}$$

$$-(14.2x) = 733$$

$$x = \frac{733}{-14.2} = -52 \frac{\text{kJ}}{\text{mol rxn}}$$

DATA AND OBSERVATIONS (CONTINUED)

PART III: HEAT OF REACTION

BALANCED EQUATION FOR THE REACTION



or



DATA TABLE

Table C. Heat of Reaction	
Mass of calcium	0.35 g
Volume of water	120.5 mL
T_i of water in calorimeter	22.1°C
T_f of water in calorimeter	29.3°C

SUMMARY OF PROCEDURES

- Rinse and dry the cups used in Part I.
- Measure out a small amount of calcium metal.
Record the mass in the data table.
- Transfer a known volume of water into the cups, replace the lid, and record the initial temperature with the temperature probe.
- Set up the data collection device to record the temperature every second for 180 seconds.
- Add the solid calcium metal, stir, and continue recording the temperature until the data collection device stops.

ANALYSIS

$$-q_{\text{lost}} = q_{\text{gained}}$$

$$-q_{\text{reaction}} = q_{\text{solution}}$$

$$q_{\text{reaction}} = 0.35 \text{ g} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \times \frac{1 \text{ mol rxn}}{1 \text{ mol Ca}} \times \frac{x \text{ kJ}}{1 \text{ mol rxn}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 8.73x \text{ J}$$

$$q_{\text{solution}} = (120.5 + 0.35) \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (29.3 - 22.1)^\circ\text{C} = 3641 \text{ J}$$

$$-(8.73x) = 3641$$

$$x = \frac{3641}{-8.73} = -420 \frac{\text{kJ}}{\text{mol rxn}}$$

CONCLUSION QUESTIONS

1. For each of the three investigations, determine whether each process is *endothermic* or *exothermic*. State your reasoning.

Part I: Endothermic. The temperature change was negative, indicating that the solution was cooler; the enthalpy change was positive.

Part II: Exothermic. The temperature change was positive, indicating that the solution was warmer; the enthalpy change was negative.

Part III: Exothermic. The temperature change was positive, indicating that the solution was warmer; the enthalpy change was negative.

2. In the sample experiment to determine the specific heat capacity of a metal, would the calculated value be reported as too large, too small, or no different than the theoretical value if the experimenter took too long to transfer the metal to the calorimeter? Explain your reasoning.

The specific heat capacity would be reported as too low. The metal is at a higher temperature than the air, so it will lose energy during the transfer.

The initial temperature of the metal would be recorded as the boiling temperature of the water but it is actually less. This would cause the calculated difference between the final and initial temperature to be too large.

Temperature is in the denominator of the specific heat ratio. Dividing by a larger change in temperature will result in a specific heat capacity that is smaller than it should actually be.

3. In Part I, would your calculated value for the heat of solution be reported as too large, too small, or the same as the theoretical value if some of the solid sample remained in the weigh boat instead of being transferred into the calorimeter? Explain your reasoning.

The calculated value for heat of solution would be reported as too low if some of the solid remained in the weigh boat.

The recorded mass of solid would be larger than the actual amount that dissociated. This means that the unit of moles, found in the denominator, would be larger than the actual value. When divided by a larger number, the amount of energy in the numerator yields a smaller value for heat of solution.

4. In these experiments, it is assumed that heat is neither transferred to nor from the calorimeter. This is not actually true. The heat transferred to the calorimeter or by the calorimeter is a part of the system. How would you decide whether the term $q_{\text{calorimeter}}$ belongs as a component of q_{lost} or q_{gained} in the analysis of the experiment? Explain.

The term $q_{\text{calorimeter}}$ can be a component of either q_{lost} or q_{gained} . The temperature change for the calorimeter will be the same as for the water or solution inside due to the thermal equilibrium. For exothermic processes happening in the calorimeter, the calorimeter will gain heat. For endothermic processes in the calorimeter, the calorimeter will lose heat.

CONCLUSION QUESTIONS (CONTINUED)

5. Calcium metal reacts readily with air. Consider the reaction shown in Equation 7:



Assume the calcium turning used in the experiment had a coating of calcium oxide.

- a. Would the recorded mass of calcium be greater, smaller, or the same as the actual value of the calcium turning? Explain.

If the calcium metal was actually calcium oxide, the reported mass of calcium would be recorded greater than the actual value. The molar mass of calcium is less than the molar mass of the calcium oxide that results from a reaction with oxygen in the air. This causes the mass of the sample to increase slightly.

- b. Would the calculated heat of reaction be reported as too large, too small, or the same? Explain.

The calculated heat of reaction would be reported as too low. The mass of calcium is reported as too great, so the number of moles would be as well. The moles of calcium are related to moles of reaction, which is in the denominator for the heat of reaction. If the denominator is artificially large, the calculation will be less than the actual value.

6. Research the standard heats of formation for all reactants and products in each of your calorimetry investigations.
- a. Determine the theoretical values for each of the three unknowns you solved for using Equation 8:

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_{f \text{ (products)}} - \sum \Delta H^\circ_{f \text{ (reactants)}}$$

Part I:

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= [1(-132) + 1(-167)] - [1(-314)] \\ &= +15 \text{ kJ/mol rxn} \end{aligned}$$

Part II:

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= [1(-286)] - [1(0) + 1(-230)] \\ &= -56 \text{ kJ/mol rxn} \end{aligned}$$

Part III:

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= [1(-543) + 1(-230) + 1(0)] \\ &\quad - [1(0) + 1(-286)] = -487 \text{ kJ/mol rxn} \end{aligned}$$

- b. Calculate the percent error for the calculated values in each investigation based on the theoretical values from (a).

$$\text{Part I: } \% \text{ error} = \left| \frac{13 - 15}{15} \right| \times 100 = 13.3\%$$

$$\text{Part II: } \% \text{ error} = \left| \frac{-52 - (-56)}{-56} \right| \times 100 = 7.1\%$$

$$\text{Part III: } \% \text{ error} = \left| \frac{-420 - (-487)}{-487} \right| \times 100 = 13.8\%$$

GOING FURTHER

Research a laboratory procedure to determine the heat capacity of the coffee cup calorimeter. If resources permit, determine the heat capacity in units of $\text{J} \cdot ^\circ\text{C}^{-1}$.

Student procedures will vary depending on the source but will likely center around the mixing of hot and cold water.



Coffee Cup Calorimetry

Measuring Heat Transfer

Chemical reactions and physical processes are always accompanied, if not driven by, a change in energy. **Heat**, the energy transferred due to a difference in temperature, is measured in calorimetry experiments. The analysis of a calorimetry experiment changes depending on the nature of the calorimeter and processes involved but is always based on the law of conservation of energy.

As the unit for heat, q , is attributed with a positive or negative sign to indicate the direction of the transfer, we can say

$$-q_{\text{lost}} = q_{\text{gained}} \quad (\text{Eq. 1})$$

In your investigations, you will be using a coffee cup calorimeter similar to the one shown in Figure 1. Read through the sample procedures, data, and analysis for a calorimetry experiment for a physical process of heat transfer. Answer the questions in the Pre-Lab Exercises that follow.

MATERIALS

balance

data collection device

*graduated cylinder,
100 mL*

*sensor, temperature,
stainless steel*

weigh boat

ammonium chloride

calcium turnings

hydrochloric acid, 0.5 M

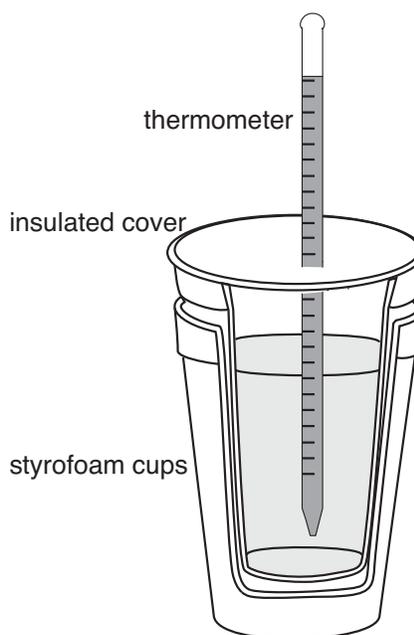
*sodium hydroxide,
0.5 M*

*2 cups, 8-oz Styrofoam®,
with lids*

spoon, plastic

water, distilled

Figure 1. Coffee cup calorimeter



DETERMINING THE SPECIFIC HEAT OF METAL**SAMPLE PROCEDURE**

1. Fill a 400-mL beaker halfway with water. Bring the water to boil on a hot plate and record the boiling temperature using a temperature sensor.
2. Obtain two stacked coffee cups and a fitted lid. Add approximately 100 mL of water and record the exact volume of water used.

Place the lid on the coffee cups and insert the temperature probe through the lid into the water.

3. Obtain a metal block. Record the identity and mass of the sample. Using crucible tongs, transfer the metal to the boiling water. Be sure that the metal is completely submerged in the water.
4. Set up a data collection device to record the temperature every second for 180 seconds.
5. Place the temperature probe in the water and begin data collection.
Open the lid, allowing the temperature probe to briefly come out of the water. With extreme care, use crucible tongs to transfer the metal from the boiling water to the water in the calorimeter. Replace the lid and the temperature probe.
6. Keeping the base of the coffee cups on the table, swirl the water by moving the cup in small circular motions. Continue data collection until time has elapsed.

DETERMINING THE SPECIFIC HEAT OF METAL (CONTINUED)

Table 1. Sample Data	
Identity of the Metal	Copper
Mass of metal	25.00 g
Volume of water	98.5 mL
T of boiling water	99.8°C
T_i of water in calorimeter	22.3°C
T_f of water in calorimeter	24.1°C

SAMPLE ANALYSIS

$$-q_{\text{lost}} = q_{\text{gained}} \quad (\text{Eq. 2})$$

$$-q_{\text{Cu}} = q_{\text{water}} \quad (\text{Eq. 3})$$

$$q_{\text{Cu}} = 25.0 \text{ g} \times \frac{x \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (24.1 - 99.8)^\circ\text{C} = -1892.5x \text{ J} \quad (\text{Eq. 4})$$

$$q_{\text{water}} = 98.5 \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (24.1 - 22.3)^\circ\text{C} = 741.8 \text{ J} \quad (\text{Eq. 5})$$

$$-(-1892.5x) = 741.8 \quad (\text{Eq. 6})$$

$$x = \frac{741.8}{1892.5} = 0.39 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$$

PROCEDURE**PART I: HEAT OF SOLUTION**

Design an experiment to determine the heat of solution, $\Delta H_{\text{soln}}^{\circ}$, for the dissolution of ammonium chloride in water. Solve for units of $\text{kJ}\cdot\text{mol}^{-1}$ of ammonium chloride. The rate of dissolution is increased with agitation, so be sure to adequately swirl the solution.

PARAMETERS

- Use between 1.0 g and 2.0 g of solid ammonium chloride.
- Use between 50.0 mL and 80.0 mL of distilled water.

ASSUMPTIONS

- The total mass of solution includes the mass of ammonium chloride.
- The specific heat of the solution is the same as that for water.

PART II: HEAT OF NEUTRALIZATION

Design an experiment to determine the heat of neutralization, $\Delta H_{\text{neut}}^{\circ}$, for the reaction of sodium hydroxide and hydrochloric acid. Solve for units of $\text{kJ}\cdot\text{mol}^{-1}_{\text{rxn}}$ for the balanced equation using the lowest whole-number coefficients.

PARAMETERS

- The solution concentrations are all 0.5 *M*.
- The total volume of the solution must not exceed 60.0 mL.

ASSUMPTIONS

- Volumes are additive.
- Both solutions have the same starting temperature.
- The density and specific heat of the solutions are the same as those for water.

PROCEDURE (CONTINUED)**PART III: HEAT OF REACTION**

Design an experiment to determine the heat of reaction, $\Delta H_{\text{rxn}}^{\circ}$, for the reaction of calcium metal with water. Solve for units of $\text{kJ}\cdot\text{mol}^{-1}_{\text{rxn}}$ for the balanced equation using the lowest whole-number coefficients.

PARAMETERS

- The amount of distilled water in the calorimeter can range from 80.0 mL to 150.0 mL.
- Use only one (1) calcium turning.

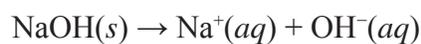
ASSUMPTIONS

- The total mass of the solution includes the mass of added calcium.
- The specific heat of the solution is the same as that for water.

PRE-LAB EXERCISES (CONTINUED)

8. Considering the actual data collected, state the assumption made in Equation 5 of the Sample Analysis.

9. The heat calculation, q , differs from the sample analysis when the system losing or gaining heat is a process for which a heat capacity and temperature change are not applicable. Calculate the heat transferred by the dissolution of 3.0 g of NaOH(s) in water.



$$\Delta H_{\text{soln}}^{\circ} = -44.5 \text{ kJ}\cdot\text{mol}^{-1}$$

DATA AND OBSERVATIONS

PART I: HEAT OF SOLUTION

BALANCED EQUATION FOR THE DISSOLUTION

SUMMARY OF PROCEDURES

DATA TABLE

ANALYSIS

DATA AND OBSERVATIONS (CONTINUED)

PART II: HEAT OF NEUTRALIZATION

BALANCED EQUATION FOR THE NEUTRALIZATION

SUMMARY OF PROCEDURES

DATA TABLE

ANALYSIS

DATA AND OBSERVATIONS (CONTINUED)

PART III: HEAT OF REACTION

BALANCED EQUATION FOR THE REACTION

SUMMARY OF PROCEDURES

DATA TABLE

ANALYSIS

CONCLUSION QUESTIONS (CONTINUED)

6. Research the standard heats of formation for all reactants and products in each of your calorimetry investigations.
- a. Determine the theoretical values for each of the three unknowns you solved for using Equation 8:

$$\Delta H^{\circ}_{\text{rxn}} = \sum \Delta H^{\circ}_{f}(\text{products}) - \sum \Delta H^{\circ}_{f}(\text{reactants}) \quad (\text{Eq. 8})$$

- b. Calculate the percent error for the calculated values in each investigation based on the theoretical values from (a).

GOING FURTHER

Research a laboratory procedure to determine the heat capacity of the coffee cup calorimeter. If resources permit, determine the heat capacity in units of $\text{J}\cdot\text{C}^{-1}$.