The Hand Warmer Design Challenge: Where Does the Heat Come From?

■ CENTRAL CHALLENGE

The ideal hand warmer increases in temperature by 20°C (but no more) as quickly as possible, has a volume of about 50 mL, costs as little as possible to make, and uses chemicals that are as safe and environmentally friendly as possible. You will carry out an experiment to determine which substances, in what amounts, to use in order to make a hand warmer that meets these criteria.

■ CONTEXT FOR THIS INVESTIGATION

Have your fingers ever been so cold they felt numb? Wouldn't it be great if you could generate heat to warm your hands up anytime you want to? That's exactly what a "hand warmer" does. Hand warmers are small packets that people put inside gloves or mittens on cold days to keep their fingers warm. They are very popular with people who work outside in winter or do winter sports. One type of hand warmer contains water in one section of the packet and a soluble substance in another section. When the packet is squeezed the water and the soluble substance are mixed, the solid dissolves, and the packet becomes warm. In this experiment you will learn how a hand warmer works and use chemistry to design an effective, safe, environmentally benign, and inexpensive hand warmer.

■ PRELAB GUIDING QUESTIONS/SIMULATIONS

Your teacher may assign you the following animation to watch and associated prelab questions to strengthen your mental models of the particulate nature of matter.

An animation showing the dissolution of an ionic compound on the particulate level can be found on the website Chemistry Experiment Simulations and Conceptual Computer Animations: *http://group.chem.iastate.edu/Greenbowe/ sections/projectfolder/flashfiles/thermochem/solutionSalt.html*

Watch the animation, and then work with your lab partner to answer the following questions.

1. Describe the changes you observe in the animation, including changes in the bonds and particulate attractions and changes in the amount of disorder in the system.

2. When sodium chloride is dissolved in water, the temperature of the resulting solution is lower than the temperature of the water before the salt dissolves. How can this result be explained based on bond breaking and bond making that is occurring?

3. Why do some salts, such as sodium chloride, dissolve spontaneously even though the process is endothermic overall?

4. When some ionic salts are dissolved in water the temperature of the resulting solution is higher than the temperature of the water before the salt dissolves. What do you think determines whether the resulting solution is cooler or warmer than the starting water?

■ EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

Breaking bonds and particulate attractions absorbs energy from the surroundings, while forming new bonds and particulate attractions releases energy to the surroundings. When an ionic solid dissolves in water, ionic bonds between cations and anions in the ionic solid and hydrogen bonds between water molecules are broken, and new attractions between water molecules and anions and water molecules and cations are formed. The amount of energy required to break these bonds and form new ones depends on the chemical properties of the particular anions and cations. Therefore, when some ionic solids dissolve, more energy is required to break the cation-anion bonds than is released in forming the new water-ion attractions, and the overall process absorbs energy in the form of heat. When other ionic compounds dissolve, the converse is true, and the bond making releases more energy than the bond breaking absorbs, and therefore the process overall releases heat. When heat is absorbed, the enthalpy change, q, is endothermic, and the enthalpy change is positive. When heat is released, the change is exothermic, and the value of q is negative. The entropy change of solution formation is always positive, regardless of whether it is endothermic or exothermic, because solutions are much more disordered than are the pure solute and solvent from which they are made. This positive entropy change is thermodynamically favorable.

■ PREPARATION

Materials

Thermometer	Sodium chloride (NaCl)
Balance	Anhydrous calcium chloride ($CaCl_2$)
100 mL graduated cylinder	Anhydrous sodium carbonate *Na_2CO_3)
Scoops	2–8 oz polystyrene cups with wooden cover or lid

Dixie cups or weighing boats	Anhydrous sodium acetate ($NaC_2H_3O_2$)
Magnetic stirrer with stir bars or stirring rod	Lithium chloride (LiCl)
Magnesium sulfate ($MgSO_4$)	Ammonium nitrate (NH_4NO_3)

Safety and Disposal

The solids and resulting solutions in this investigation are potential eye and skin irritants. Calcium chloride can cause skin burns. Ammonium nitrate is a powerful oxidizer that must be kept away from ignition sources and is quite toxic on ingestion. You must wear gloves and splash-proof safety goggles and must use caution throughout the experiment and adhere to all safety guidelines. If solutions are spilled on skin, wash with copious amounts of water. Follow your teacher's directions for the disposal of solutions and any unused chemicals.

■ PRACTICE WITH INSTRUMENTATION AND PROCEDURE

In this experiment, you will collect data that will allow you to calculate the change of enthalpy of dissolution (also called the "heat of solution," with symbol ΔH_{soln}, and units of kJ/mol solute) occurring in aqueous solution. The data necessary to calculate the heat of solution can be obtained using a device called a calorimeter.

A calorimeter is a container used to determine the enthalpy change that occurs during a process. Calorimetry is an important technique in chemistry, and chemists often work with devices called bomb calorimeters. For home or classroom experiments, however, a coffee cup calorimeter is sufficient to make rough measurements. This exercise will give you practice assembling and using a calorimeter so that you can use one to help you determine which solid is best to use in a hand warmer. It will also allow you to calibrate your calorimeter with a process that supplies a known amount of heat. This calibration process allows you to determine the amount of heat the calorimeter itself absorbs as the temperature of the materials inside it change, a value known as a calorimeter constant.
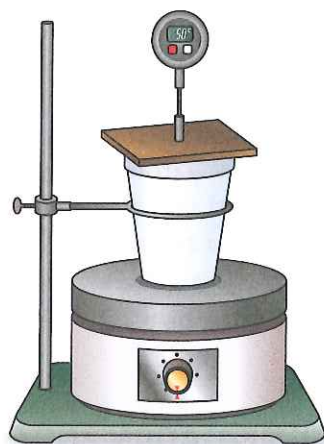
Procedure

Part 1 (calorimetry practice)

Assemble your calorimeter on a magnetic stirrer as shown in Figure 1. Measure out exactly 100.0 mL water in a graduated cylinder and pour into your calorimeter. Measure and record the temperature of the water. Add a stir bar and turn on the stirrer so the water is stirring without splashing. Measure 5.00 g magnesium sulfate anhydrate solid in a plastic cup. While monitoring the temperature of the water, quickly add all of the magnesium sulfate to the calorimeter. Record the highest temperature reached. Dilute the resulting solution with water and dispose of it per your teacher's instructions. Repeat the experiment. For each trial find the temperature change of the water per gram magnesium sulfate, and average this

value for the two trials. Check with your instructor to make sure that your average temperature change is within 10 percent of the expected value. If it is not, repeat the experiment, measuring and monitoring the temperature carefully and ensuring that all of the solid is dissolved, and recheck your results with your instructor.

Figure 1. Calorimeter assembly

Part 2 (calorimeter calibration procedure)

Place a 100.0 mL sample of water in a clean, dry 150 mL beaker. Heat with occasional stirring to approximately 50°C. Remove the beaker from the hot plate and place on the lab bench. Meanwhile, place exactly 100.0 mL of cool water (approximately 20°C) in the clean, dry calorimeter.

Measure the temperature of the hot water and the cold water and record, then immediately pour the entire hot water sample into the calorimeter and quickly put on the cover. Wait 15 seconds then take a temperature reading. Repeat this determination twice.

■ INVESTIGATION

In today's experiment you will design and execute an experimental procedure to determine which of three ionic compounds is most suitable for use in a hand warmer. Review the criteria for an ideal hand warmer from the central challenge. For each solid you need to consider safety, cost, and environmental impact as well as the amount of heat released or absorbed.

Procedure

1. Safety and environmental impact: Obtain the Material Safety Data Sheets (MSDS) for your three solids from your teacher. Review each one, making notes about safety concerns, necessary precautions, and disposal.

2. Cost: Rank the solids you are given from least to most expensive.

Substance	2012 Cost per 500 g ($)
NaCl	3.95
$CaCl_2$	6.55
$NaC_2H_3O_2$	12.90
Na_2CO_3	6.15
LiCl	32.75
NH_4NO_3	9.05

3. Heat of solution: Work with your group to design a procedure to compare the solids in terms of the heat released or absorbed when they dissolve and include what materials and equipment you will use. You must include the safety precautions you will take.

Important Procedure Tips:

1. Be sure to keep detailed records of the amounts of substances used and the starting and ending temperature as you will need it later to determine the amount of solid to use in your hand warmer.

2. You will receive a maximum of 10 g of each solid for this part.

Data Collection and Computation

Calorimeter constant determination: According to the law of conservation of energy, energy cannot be created or destroyed, only changed from one form to another or transferred from one system to another. The temperature change observed when water or any substance changes temperature can be a result of a transfer of energy from the substance to the surroundings (in which case the temperature of the substance decreases) or the surroundings to the substance (in which case the temperature of the substance increases). When hot and cold water are mixed the hot water transfers some of its thermal energy to the cold water. The law of conservation of energy dictates that the amount of thermal energy lost (or the enthalpy change) by the hot water, q_{hot}, is equal to the enthalpy change of the cold water, q_{cold}, but opposite in sign, so $q_{hot} = -q_{cold}$. The enthalpy change for any substance is directly related to the mass of substance, m; the specific heat capacity (a substance-specific constant), c; and the temperature change, ΔT. The relationship is expressed mathematically in the equation $q = mc\Delta T$. The specific heat capacity of water is 4.184 J/°C•g.

1a. Calculate the enthalpy change of the cold water using the equation $q_{hot} = m_{hot}c\Delta T_{hot}$. Assume that the density of water is exactly 1 g/mL. Is this an endothermic or exothermic process? Explain.

1b. Calculate the enthalpy change of the hot water using the equation $q_{cold} = m_{cold}c\Delta T_{cold}$. Assume that the density of water is exactly 1 g/mL. Is this an endothermic or exothermic process? Explain.

1c. These amounts are not equal because the calorimeter (the coffee cups) absorbs some of the thermal energy transferred by the hot water. Thus under the real conditions

observed in the laboratory the law of conservation of energy equation becomes $q_{hot} = -(q_{cold} + q_{cal})$, where q_{cal} is the enthalpy change of the calorimeter. Use this equation to calculate the enthalpy change of the calorimeter.

1d. The calorimeter constant, C, is the heat absorbed by the calorimeter per degree of temperature change, $C = q_{cal}/\Delta T_{cal}$. Assuming the starting temperature of the calorimeter is the same as that of the cold water, calculate the calorimeter constant in units of joules per degree Celsius.

2. The solid and water, considered together, have a certain amount of internal energy as a function of the bonds that exist in the solid and in the water. The solution that is produced as a result of the dissolving has a different amount of internal energy than the solid and water did because the arrangement of particles and the bonds and attractions between the particles in the solution are different bonds and particulate attractions than the arrangement of particles and the bonds and attractions between the particles in the solid and water. The difference in energy, q_{soln}, is the reason for the difference in the thermal energy of the two systems (solid and pure water versus solution), with symbol q_{rxn}. Just as with the hot and cold water in the calorimeter constant determination, q_{soln} and q_{rxn} are equal in magnitude and opposite in sign, $q_{rxn} = -q_{soln}$. And just as in that case of the cold and hot water mixing, the calorimeter will also experience an enthalpy change during the solution formation process. To account for this enthalpy change the relationship is adjusted to $q_{soln} = -(q_{rxn} + C\Delta T)$ where C is the calorimeter constant determined above.

This difference in thermal energy of the system before and after solution formation, q_{soln}, can be calculated using the relationship $q_{rxn} = mc\Delta T$, where m is the total mass of the solution and c is the specific heat capacity of the solution and ΔT is the temperature change of the solution. It is important to note that we will assume that the heat capacity of the solutions is the same as pure water but in reality the solutions do not have exactly the same heat capacity, and this assumption affects the accuracy of this determination.

Using this information, calculate q_{soln} and q_{rxn} for all three solids you tested for your hand warmer.

3. By convention, scientists report enthalpy changes for dissolution (and many other processes) in units of kilojoules per mole of solute dissolved. Using your values of q_{soln}, calculate the enthalpy in units of kilojoules per mole. This quantity has the symbol ΔH_{soln}. Calculate ΔH_{soln} for each of the three solids you tested.

4. Based on the cost information provided, and your experimental work and calculations, select which chemical you believe will make the most cost-effective hand warmer. The hand warmer you are designing needs to increase in temperature by 20°C. Calculate the amount of the compound you selected that would be required for a hand warmer that meets this requirement.

Argumentation and Documentation

Write a paragraph in which you describe all of the factors you considered and you explain your rationale for choosing one chemical and not each of the other chemicals studied in this experiment. Your paragraph should start with a claim sentence that clearly states your choice and the amount of substance to use. The claim should be followed by evidence from your experiment and cost and safety analysis. The paragraph should conclude with reasoning explaining how your evidence supports your claim.

POSTLAB ASSESSMENT

1. Are the dissolving processes you carried out endothermic or exothermic or neither? Explain your thinking.

2. Dissolving ionic compounds involves the separation of the solid ionic compound into cations and anions in water. This process can be represented by an equation showing the solid as a reactant and the aqueous ions as products. The heat of reaction ΔH_{soln} is written after the products, typically in units of kJ/mol.

Example: sodium hydroxide dissolves exothermically, releasing 44.2 kilojoules per mole dissolved. This process is represented as $NaOH(s) \rightarrow Na^+ (aq) + OH^- (aq)$, $\Delta H_{soln} = -44.2 kJ/mol$.

Write an equation to represent the dissolving process for each salt you studied. Include your calculated heat of reaction as in the example.

3. Changes in matter are generally classified as physical or chemical based on whether new substances are formed through the process. Does dissolving represent a physical change, a chemical change, or an intermediate change? Explain your reasoning, including evidence from the animation you viewed.

4. Share your calculated values of ΔH_{soln} with your classmates and obtain their values.

   a. Determine the class average value and standard deviation for each solid.

   b. Find the published value of ΔH°_{soln} for each solid and determine the percent error in the class average value.

5. What possible sources of error could affect the accuracy of your calculated value of the amount of solid in your hand warmer? List at least two and what effect they would have on the temperature change.

■ SUPPLEMENTAL RESOURCES

Links

"Calorimetry." Sparknotes SAT Chemistry Common Experiments. Accessed August 3, 2012.
http://www.sparknotes.com/testprep/books/sat2/chemistry/chapter11section4.rhtml

"Heats of Reaction." University of Massachusetts Boston. Accessed July 31, 2012.
http://alpha.chem.umb.edu/chemistry/genchem/103/files/103lab/7CoffeeCupCalorimeterRevised.pdf

References

Barlag, Rebecca E., Phyllis Arthasery, and Frazier Nyasulu. "Electrical Determination of the Heat Capacity of a Calorimeter in Approximately One Minute." *Journal of Chemical Education* 87, no. 9 (2010): 992–992.

Brouwer, Henry. "Small-Scale Thermochemistry Experiment." *Journal of Chemical Education* 68, no. 7 (1991): A178.

Marzacco, Charles J. "The Enthalpy of Decomposition of Hydrogen Peroxide: A General Chemistry Calorimetry Experiment." *Journal of Chemical Education* 76, no. 11 (1999): 1517.

Vannatta, Michael W., Michelle Richards-Babb, and Robert J. Sweeney. "Thermochemistry to the Rescue: A Novel Calorimetry Experiment for General Chemistry." *Journal of Chemical Education* 87, no. 11 (2010): 1222–1224.

Designing a Hand Warmer
AP* Chemistry Big Idea 5, Investigation 12
An Advanced Inquiry Lab

Introduction

Put your chemistry skills to commercial use! From instant cold packs to flameless ration heaters and hand warmers, the energy changes accompanying physical and chemical transformations have many consumer applications. The backbone of these applications is calorimetry—measuring heat transfer. Investigate the energy changes accompanying the formation of solutions for common laboratory salts, and then apply the results to design a hand warmer that is reliable, safe and inexpensive.

Concepts

- Enthalpy change
- Heat of solution
- Exothermic versus endothermic
- Calorimetry
- Specific heat
- System and surroundings

Background

Hand warmers are familiar cold weather gear used to quickly provide warmth to frigid fingers. Many commercial hand warmers consist of a plastic package containing a solid and an inner pouch filled with water. When the pack is activated, the solid dissolves in water and produces a large temperature change.

The energy or *enthalpy change* associated with the process of a solute dissolving in a solvent is called the *heat of solution* (ΔH_{soln}). At constant pressure, this enthalpy change, ΔH_{soln}, is equal in magnitude to the heat loss or gain, q, to the surroundings. In the case of an ionic solid dissolving in water, the overall energy change is the net result of three processes—the energy required to break the attractive forces between ions in the crystal lattice ($\Delta H_1 = +C$ kJ/mole), the energy required to disrupt intermolecular forces between water molecules ($\Delta H_2 = +D$ kJ/mole), and the energy released when the dissociated (free) ions form ion-dipole attractive forces with the water molecules ($\Delta H_3 = -F$ kJ/mole). The overall process can be represented by the following equations.

$$M_aX_b(s) \rightarrow aM^{+b}(aq) + bX^{-a}(aq) \qquad \Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3 = (+C + D - F) \text{ kJ/mole}$$

If the amount of energy *released* in the formation of hydrated ions (ΔH_3) is greater than the amount of energy *required* to separate the solute and solvent particles ($\Delta H_1 + \Delta H_2$), then the sum (ΔH_{soln}) of the energy changes will be negative and the solution process exothermic (releases heat). If the amount of energy released in the formation of hydrated ions is less than the amount of energy required to separate the solute and solvent particles, then the sum of the energy changes will be positive and the solution process endothermic (absorbs heat).

Heats of solution and other enthalpy changes are generally measured in an insulated vessel called a *calorimeter* that reduces or prevents heat loss to the atmosphere outside the reaction vessel. The process of a solute dissolving in water may either release heat into the resulting aqueous solution or absorb heat from the solution, but the amount of heat exchanged between the calorimeter and the outside surroundings should be minimal. When using a calorimeter, the reagents being studied are mixed directly in the calorimeter and the temperature is recorded both before and after the reaction has occurred. The amount of heat transfer (q) may be calculated using the heat energy equation:

$$q = m \times s \times \Delta T \qquad \textit{Equation 1}$$

where m is the **total mass of the solution** (solute plus solvent), s is the specific heat of the solution, and ΔT is the observed temperature change. The specific heat of the solution is generally assumed to be the same as that of water, namely, 4.18 J/g·°C.

When measuring the heat transfer for an *exothermic* heat of solution using a calorimeter, most of the heat released is absorbed by the aqueous solution (q_{aq}). A small amount of the heat will be absorbed by the calorimeter itself (q_{cal}). The overall heat transfer (q_{soln}) for the reaction (the system) then becomes:

$$q_{soln} = -(q_{aq} + q_{cal})$$ *Equation 2*

In order to determine the correction factor q_{cal} for heat of solution calculations, the heat capacity of the calorimeter, also called the calorimeter constant, must be determined experimentally. The calorimeter constant has units J/°C. This calibration experiment is done by mixing equal volumes of hot and cool water in the calorimeter and measuring the temperature after 20 seconds. The resulting value is assumed to be the instantaneous mixing temperature, T_{mix}. The average temperature T_{avg} of the initial hot (T_H) and cool water (T_C) is also calculated:

$$T_{avg} = (T_H + T_C)/2$$

The difference between T_{avg} and T_{mix} is due to the heat lost by the water and absorbed by the calorimeter. The heat lost by the water, q_{water}, is:

$$q_{water} = (\text{mass of water}) \times (\text{specific heat of water}) \times (T_{mix} - T_{avg})$$ *Equation 3*

where the mass is the total mass of hot and cool water. The heat gained by the calorimeter, q_{calor}, is equal to that lost by the water, but opposite in sign. The calorimeter constant, C_{cal}, is calculated as follows:

$$C_{cal} = \frac{q_{calor}}{(T_{mix} - T_{initial})}$$ *Equation 4*

where $T_{initial}$ is the initial temperature of the calorimeter containing cool water.

To calculate the correction factor q_{cal} for use in Equation 2 above—to determine the heat of solution or heat of reaction for any system—the calorimeter constant is multiplied by the change in temperature of that solution:

$$q_{cal} = \Delta T \,(°C) \times C_{cal} \,(\text{J/°C})$$

Experiment Overview

The purpose of this advanced inquiry lab is to design an effective hand warmer that is inexpensive, nontoxic and safe for the environment. The investigation begins with an introductory activity to become familiar with the principles of calorimetry and heat of solution calculations. The results provide a model for the guided-inquiry challenge, which is to design an optimum hand warmer for consumer applications. Working in groups of four, each student group will be provided six different solids, along with their costs and individual Material Safety Data Sheets (MSDS). Determine the heat of solution for each solid and analyze the cost and safety information to propose a design for the best all-around hand warmer.

Pre-Lab Questions

1. When chromium chloride, $CrCl_2$, is dissolved in water, the temperature of the water decreases.

   a. Is the heat of solution exothermic or endothermic?

   b. Which is stronger—the attractive forces between water molecules and chromium and chloride ions, or the combined ionic bond strength of $CrCl_2$ and intermolecular forces between water molecules? Explain.

2. A solution was formed by combining 25.0 g of solid A with 60.0 mL of distilled water, with the water initially at 21.4 °C. The final temperature of the solution was 25.3 °C. Calculate the heat released as the solid dissolved, q_{soln}, assuming no heat loss to the calorimeter (see Equation 1).

3. In Question 2 above, the calorimeter was found to have a heat capacity of 8.20 J/°C. If a correction is included to account for the heat absorbed by the calorimeter, what is the heat of solution, q_{soln}?

4. The solid in Question 2 was aluminum sulfate, $Al_2(SO_4)_3$. Calculate the molar heat of solution, ΔH_{soln}, for aluminum sulfate. *Hint:* The units for molar heat of solution are kilojoules per mole (kJ/mole). First determine the heat released per gram of solid.

Materials (for each lab group)

Ionic solids:

Set A: Ammonium chloride, NH₄Cl, 15 g

Calcium chloride, anhydrous, $CaCl_2$, 15 g

Sodium acetate, $NaCH_3CO_2$, 15 g

Set B: Sodium chloride, NaCl, 15 g

Lithium chloride, LiCl, 15 g

Sodium carbonate, Na_2CO_3, 15 g

Magnesium sulfate, anhydrous, $MgSO_4$, 5 g

Water, deionized or distilled

Balance, 0.01-g precision (shared)

Beaker, 250-mL

Calorimeter (two nested polystyrene cups)

Graduated cylinder, 100-mL

Heat-resistant gloves

Hot plate

Magnetic stirrer and stir bar, or stirring rod

Paper towels

Support stand and ring clamp

Thermometer, digital

Timer or stopwatch

Weighing dishes

Safety Precautions

Lithium chloride is moderately toxic by ingestion. Calcium chloride and ammonium chloride are slightly toxic. Magnesium sulfate is a body tissue irritant. Sodium acetate is a body tissue and respiratory tract irritant. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines.

Introductory Activity

Part A. Heat Capacity of the Calorimeter

1. Working in pairs, set up a calorimeter consisting of two nested polystyrene cups in a ring clamp attached to a support stand.

2. Place a magnetic stirrer below the calorimeter, then lower the ring clamp until the bottom of the cup just sits on the surface of the magnetic stirrer (see Figure 1).

3. Measure 100.0 mL of distilled water in a 100-mL graduated cylinder and transfer the water into the calorimeter.

Figure 1.

4. Add a magnetic stirring bar to the calorimeter, and set the bar spinning slowly. If a magnetic stirrer is not available, use a stirring rod. Do not remove the stirring rod from the calorimeter.

5. Measure and record the initial temperature of the water.

6. Heat approximately 125 mL of distilled water to 60–70 °C in a 250-mL beaker.

7. Using heat-resistant gloves, measure 100.0 mL of the hot water in a 100-mL graduated cylinder.

8. Measure and record the temperature of the hot water.

9. Immediately pour the hot water into the room temperature water in the calorimeter.

10. Insert the thermometer, and stir the water.

11. Record the mixing temperature T_{mix} after 20 seconds.

12. Empty the calorimeter and dry the inside.

13. Calculate the calorimeter constant, C_{cal}, using T_{mix} and Equations 3 and 4 from the *Background* section.

IN7654

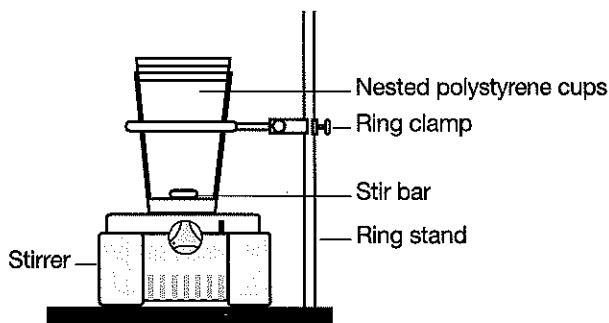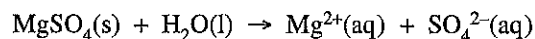### Part B. Calorimetry Procedure

Working in pairs, examine the heat energy change for the following solution.

$$MgSO_4(s) + H_2O(l) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$$

1. Measure 100.0 mL of distilled or deionized water in a 100-mL graduated cylinder and transfer to the calorimeter.

2. Measure and record the initial temperature of the water.

3. Measure 5.00 g of anhydrous magnesium sulfate in a weighing dish.

4. Put a magnetic stir bar or stirring rod into the calorimeter and slowly stir the water.

5. Quickly add the 5.00 g of anhydrous magnesium sulfate to the calorimeter and insert the thermometer.

6. Monitor the temperature and record the highest or lowest temperature reading.

7. Calculate the molar heat of solution for magnesium sulfate. Include the correction due to the heat capacity of the calorimeter.

Guided-Inquiry Design and Procedure

Form a working group with other students and discuss the following questions.

1. Review the calorimetry procedure:

   a. What data is needed to calculate the enthalpy change for a reaction?

   b. Identify the variables that will influence the experimental data.

   c. What variables should be controlled (kept constant) during the procedure?

   d. The independent variable in an experiment is the variable that is changed by the experimenter, while the dependent variable responds to or depends on the changes in the independent variable. Name the independent and dependent variables in a calorimetry experiment to determine the molar heat of solution.

   e. Discuss the factors that will affect the precision of the experimental results.

2. One pair of students in the group should study the three solids in Set A, while the other pair studies Set B.

3. Working collaboratively with the general procedure provided in the *Introductory Activity*, design and carry out experiments to determine the heat of solution for each solid. Be sure to review all safety precautions with your instructor before starting.

4. Extrapolating from the information collected, predict which solid(s) could be used in an effective hand warmer meeting the following requirements:

   • The hand warmer must contain 10 g of an ionic solid and an inner pouch filled with 40 mL of water.

   • Activating the hand warmer must increase the temperature of the resulting solution by at least 20 °C.

   • The solid should be nontoxic, safe for the environment, and economical.

5. Review the cost information shown below and consult the MSDS for each potential hand warmer. Propose the optimum design for the most cost-effective hand warmer that is nontoxic and least harmful to the environment.

Solid	NH_4Cl	$CaCl_2$	LiCl	$NaCH_3CO_2$	Na_2CO_3	NaCl
Cost ($)/Kilogram	21.90	10.80	68.30	27.30	5.95	4.25

6. With your instructor's permission, verify the design and demonstrate the use of your hand warmer.

Opportunities for Inquiry

Instant cold packs are used to treat sports and other injuries when ice is unavailable. Research the properties of commercial cold packs and select a "cold pack solid" from among the solids provided in this activity. Propose and test a design (quantity of solid and volume of water) for an effective instant cold pack.

AP Chemistry Review Questions

Integrating Content, Inquiry and Reasoning

Review the following data from a calorimetry experiment to determine the heat of fusion of ice. After shaking off any excess water, several ice cubes were added to 99 g of warm water contained in a calorimeter. The initial temperature of the warm water was 46.8 °C. The ice–water mixture was stirred until the temperature reached a stable, minimum value, which was 1.1 °C. Any *unmelted* ice remaining at this point was immediately and carefully removed using tongs and the mass of the water in the calorimeter was measured—154 g.

1. Use the heat energy equation to calculate the amount of heat in joules released by the warm water as it cooled.

2. Based on the law of conservation of energy, what amount of heat was absorbed by the ice as it melted?

3. Determine the amount of energy absorbed per gram of ice as it melted.

4. Calculate the heat of fusion (the heat required to melt ice) in units of kilojoules/mole.

5. The literature value for the heat of fusion of ice is 6.02 kJ/mole. What is the percent error for the experimentally determined heat of fusion?

6. When a mixture of ice and water originally at 0 °C is heated, the temperature remains constant (within experimental error) until all of the ice melts. Explain what happens to the heat energy that is absorbed during this time while the temperature does not change.

IN7654