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# 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* ( $\Delta H_{\text{soln}}$ ). At constant pressure, this enthalpy change,  $\Delta H_{\text{soln}}$ , 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.



If the amount of energy *released* in the formation of hydrated ions ( $\Delta H_3$ ) is greater than the amount of energy *required* to separate the solute and solvent particles ( $\Delta H_1 + \Delta H_2$ ), then the sum ( $\Delta H_{\text{soln}}$ ) 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:

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

where  $m$  is the **total mass of the solution** (solute plus solvent),  $s$  is the specific heat of the solution, and  $\Delta T$  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.

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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 ( $q_{\text{aq}}$ ). A small amount of the heat will be absorbed by the calorimeter itself ( $q_{\text{cal}}$ ). The overall heat transfer ( $q_{\text{soln}}$ ) for the reaction (the system) then becomes:

$$q_{\text{soln}} = -(q_{\text{aq}} + q_{\text{cal}}) \quad \text{Equation 2}$$

In order to determine the correction factor  $q_{\text{cal}}$  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  $\text{J}/^\circ\text{C}$ . 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,  $T_{\text{mix}}$ . The average temperature  $T_{\text{avg}}$  of the initial hot ( $T_{\text{H}}$ ) and cool water ( $T_{\text{C}}$ ) is also calculated:

$$T_{\text{avg}} = (T_{\text{H}} + T_{\text{C}})/2$$

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

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

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

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

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

To calculate the correction factor  $q_{\text{cal}}$  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:

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

## 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

- When chromium chloride,  $\text{CrCl}_2$ , is dissolved in water, the temperature of the water decreases.
  - Is the heat of solution exothermic or endothermic?
  - Which is stronger—the attractive forces between water molecules and chromium and chloride ions, or the combined ionic bond strength of  $\text{CrCl}_2$  and intermolecular forces between water molecules? Explain.
- A solution was formed by combining 25.0 g of solid A with 60.0 mL of distilled water, with the water initially at  $21.4^\circ\text{C}$ . The final temperature of the solution was  $25.3^\circ\text{C}$ . Calculate the heat released as the solid dissolved,  $q_{\text{soln}}$ , assuming no heat loss to the calorimeter (see Equation 1).
- In Question 2 above, the calorimeter was found to have a heat capacity of  $8.20 \text{ J}/^\circ\text{C}$ . If a correction is included to account for the heat absorbed by the calorimeter, what is the heat of solution,  $q_{\text{soln}}$ ?
- The solid in Question 2 was aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$ . Calculate the molar heat of solution,  $\Delta H_{\text{soln}}$ , for aluminum sulfate. *Hint:* The units for molar heat of solution are kilojoules per mole ( $\text{kJ}/\text{mole}$ ). First determine the heat released per gram of solid.

## Materials (for each lab group)

Ionic solids:

Set A: Ammonium chloride,  $\text{NH}_4\text{Cl}$ , 15 g

Calcium chloride, anhydrous,  $\text{CaCl}_2$ , 15 g

Sodium acetate,  $\text{NaCH}_3\text{CO}_2$ , 15 g

Set B: Sodium chloride,  $\text{NaCl}$ , 15 g

Lithium chloride,  $\text{LiCl}$ , 15 g

Sodium carbonate,  $\text{Na}_2\text{CO}_3$ , 15 g

Magnesium sulfate, anhydrous,  $\text{MgSO}_4$ , 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.
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  $T_{\text{mix}}$  after 20 seconds.
12. Empty the calorimeter and dry the inside.
13. Calculate the calorimeter constant,  $C_{\text{cal}}$ , using  $T_{\text{mix}}$  and Equations 3 and 4 from the *Background* section.

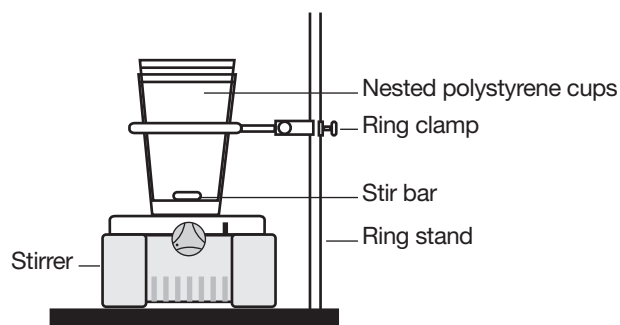
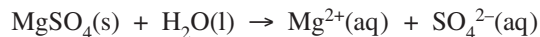


Figure 1.

## Part B. Calorimetry Procedure

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



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.

<b>Solid</b>	NH <sub>4</sub> Cl	CaCl <sub>2</sub>	LiCl	NaCH <sub>3</sub> CO <sub>2</sub>	Na <sub>2</sub> CO <sub>3</sub>	NaCl
<b>Cost (\$)/Kilogram</b>	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.

**Please contact Flinn for the Teacher Notes and Sample Data**