

## ■ PREPARATION

### Materials

Thermometer	
Balance	
100 mL graduated cylinder	
Scoops	2-8 oz polystyrene cups with wooden cover or lid

3 salts from list of 6 possible

### 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,  $\Delta T$ . The relationship is expressed mathematically in the equation  $q = mc\Delta T$ . The specific heat capacity of water is  $4.184 \text{ J/}^\circ\text{C}\cdot\text{g}$ .

- 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  $1 \text{ g/mL}$ . 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  $1 \text{ g/mL}$ . 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

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  $\Delta T$  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  $\Delta H_{soln}$ . Calculate  $\Delta H_{soln}$  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  $20^{\circ}\text{C}$ . 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  $\Delta H_{soln}$  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  $\text{NaOH}(s) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$ ,  $\Delta H_{soln} = -44.2\text{kJ/mol}$ .

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  $\Delta H_{soln}$  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  $\Delta H_{soln}^{\circ}$  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