Designing a Hand Warmer -- AP* Chemistry

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.

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 ($\Delta H_1 = + C \text{ kJ/mole}$), the energy required to disrupt intermolecular forces between water molecules ($\Delta H_2 = + D \text{ kJ/mole}$), and the energy released when the dissociated (free) ions form ion-dipole attractive forces with the water molecules ($\Delta H_3 = - F \text{ kJ/mole}$). The overall process can be represented by the following equation.

$$\text{MaXb(s) \rightarrow aM^+b(aq) + bX^-a(aq)}$$

$$\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3 = (+ C + D - F) \text{ kJ/mole}$$

If the amount of energy released in the formation of hydrated ions ($\Delta H_3$) 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:

$$q = m \times c \times \Delta T \quad \text{Equation 1}$$

where $m$ is the total mass of the solution (solute plus solvent), $c$ 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.

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 ($q_{\text{soln}}$) for the reaction (the system) then becomes

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

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, $T_{\text{mix}}$. 
The average temperature $T_{\text{avg}}$ of the initial hot ($T_H$) and cool water ($T_C$) is also calculated:

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

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:

$$q_{\text{water}} = (\text{mass of water}) \times (\text{specific heat of water}) \times (T_{\text{mix}} - T_{\text{avg}})$$  \hspace{1cm} \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:

$$C_{\text{cal}} = q_{\text{calor}} / (T_{\text{mix}} - T_{\text{initial}})$$  \hspace{1cm} \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 and Purpose
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, each student group will be provided three different solids, along with their costs. 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, $\text{CrCl}_2$, 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 $\text{CrCl}_2$ 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, $q_{\text{soln}}$, 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, $q_{\text{soln}}$?
4. 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 (kJ/mole). First determine the heat released per gram of solid.
5. What data is needed to calculate the enthalpy change for a reaction?
6. What variables should be controlled (kept constant) during the procedure?
7. What calculations would be performed to determine the heat of the solutions? (Don’t forget to include the heat capacity of the calorimeter that you just determined)
Part A: Heat Capacity of the Calorimeter (how much heat escapes when you use your instrument)

1. Working in pairs, set up a calorimeter consisting of two nested polystyrene cups in a beaker.
2. Turn on your interface to get a live feed reading of temperature.
3. Measure 100.0 mL of distilled water in a 100-mL graduated cylinder and transfer the water into the calorimeter.
4. Put thermometer into the water. (You will use the thermometer as a stir rod...carefully)
5. Measure and record the initial temperature of the water.
6. Measure 100.0 mL of hot water in a 100-mL graduated cylinder. When it gets to approximate 70°C record the exact temperature.
7. Immediately pour the hot water into the room temperature water in the calorimeter. Insert the thermometer, and stir the water.
8. Record the mixing temperature $T_{\text{mix}}$ after 20 seconds.
9. Empty the calorimeter and dry the inside with a paper towel.
10. Calculate the calorimeter constant, $C_{\text{cal}}$, using $T_{\text{mix}}$ and Equations 3 and 4 from the Background section. Read and follow this introduction CLOSELY to do these calculations. Show all calcs!

Part B:
In this section you will be answering some questions and preparing procedures to run three different coffee cup calorimetry experiments on three different salts to find the enthalpy change of dissolving.

Design three procedures to determine the heat of solution for each of the three solids. Find out from your teacher if you’re Set A or B and RECORD WHAT SET YOU HAVE in the data table. (Helpful hints: Hand warmers must contain between 10 and 15 g of an ionic solid and an inner pouch filled with 40-50 mL of water). As you develop your procedure try and keep in mind the mixing time used in Part A to maintain consistency.

Create a data table to house all necessary information.

Make sure you and your partner have the same procedures written out and are ready to roll.

Run your procedures and collect data necessary to calculate the enthalpy change of each reaction.

Record all data in your created table.

Post Lab Questions:
1. Calculate the enthalpy change for each of the salts in Joules. (Show all calcs)
2. Convert the calculated Joules value to kJ/kg of salt.
3. Search the internet for and list the environmental hazards of all the salts you tested.
4. Under this question #4, organize the answers to questions 1-3 in a table for ease of comparison.
5. Review the cost information shown below and select the salt that you think would be the most cost-effective hand warmer that is least toxic to the environment. Explain your reasoning in full.

<table>
<thead>
<tr>
<th>Solid</th>
<th>NH₄NO₃</th>
<th>CaCl₂</th>
<th>LiCl</th>
<th>NaCH₃CO₂</th>
<th>Na₂CO₃</th>
<th>NaCl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cost ($)/Kilogram</td>
<td>18.50</td>
<td>10.80</td>
<td>68.30</td>
<td>27.30</td>
<td>5.95</td>
<td>4.25</td>
</tr>
</tbody>
</table>