Arnaldo Zelaya
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INVESTIGATION 12

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

■Central Challenge: Use chemistry to design an effective, safe, environmentally benign, and inexpensive hand warmer. 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. 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 engage in 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.

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. Students 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 the normal procedures for chemical waste disposal at your institution. If it is allowed in your location, solutions may be safely diluted and poured down the drain.

■Prelab Guiding Questions/Simulations–

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

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 the 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 absorb energy from the surroundings, while forming new bonds and particulate attractions release 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. 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.

Procedure

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

Hand Warmer Design Procedure

Use the following steps to design a procedure for this investigation:

1. Safety and Environmental Impact: Obtain the 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 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

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 calorimeter. 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 cool 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 cool 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.

Cup1 = 3.5g
Cup2 = 3.5 g

Cold cup = 93.04g = 20°C
Hot cup = 93.04g = 50°C
Hot&cold cup = 36.8°C
$\Delta T = 13.2°C$

Temp°C	Na_2CO_3	LiCl	NH_4NO_3	
Cup -3.5g	19.6°C	20.2°C	20.2°C	← water
NH_4CO_3- 5.04g	15.2°C	27.7°C	22.7°C	← salts in water
Na_2CO_3 5.14				