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

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Science In Your Mittens: The Chemistry Of Hand Warmers

By RYAN | *Published:* DECEMBER 15, 2010

'Tis the season for cold fingers. If you're stuck out in the cold for a few hours, your mittens can only do so much. You may need to bring in some chemical reinforcements.

There are two types of chemical **hand warmers** that I'm going to look at here. The first are the more common disposable variety. Like a little pouch filled with something powdery. You shake it up and it starts to warm, and later on as it wears down, more shaking will squeeze out more heat. Let's find out what's going on inside the pouch.

The heat is generated from plain old iron that plain old rusts. That's right. If you could fit grandpa's Plymouth in your mittens, that would do the trick. As iron rusts (the fancy word being ***oxidizes***) it generates heat. That's **exothermic**, daddio! So inside the pouch there is a lot of iron powder. When you remove that pouch from its airtight plastic wrapper, the oxygen begins the oxidization process and it begins putting out heat.

The rest of the ingredients in the pouch serve to control the oxygen-iron reaction. Some things get the reaction started quickly for instant heat, while others work to keep the heat lasting as long as possible. Carbon is in there to help spread the heat evenly, just the same as in a barbeque. When you shake the pouch after a while, it exposes new iron to oxygen to give the rusting another boost.

Now the second kind of hand warmer is more nifty because you can see it happening. It's a clear plastic sealed pouch filled with a liquid. Floating around inside is a little metal disc about the size of a nickel. When you bend that disc, the liquid instantly "freezes" and begins generating heat. Once the heat is done you can boil the pouch to reverse the reaction and be ready to warm another day.

So let's examine that cool "freezing" part. What's going on there?

The clear liquid in the pouch is **sodium acetate**. That's sodium salt dissolved in acetic acid. It's critical that this is a super-saturated solution, meaning there is more sodium than the acid can actually hold. Plus, it's a **super-cooled** solution at room temperature. It should be "frozen" (crystallized), but it's not. The crystals have nothing to grow on. (you know, like how every snowflake is formed around a speck of dust) Until you snap that metal disc.

Adapted from AP® Chemistry Guided-Inquiry Experiments: Applying the Science Practices, Investigation 12 and Advanced Chemistry with Vernier , as well as Laboratory Experiments for Advanced Placement Chemistry by Sally Ann Vonderbrink, Ph. D. Article from "Learn Something New Every Day" http://lsned.com/

With a snap, you create a small bit of solid **sodium acetate trihydrate** that lets those first crystals form. Then further crystals grow from there, and so forth across the entire pouch in a second. The change from liquid to crystal releases heat energy. That's exothermic, dude!

After 20 minutes the heat will be done, but by putting the pouch in warm water, the crystals dissolve. In fact, the sodium acetate is soaking up heat energy that it will store until the next crystallization.

Fun bonus fact: You've probably eaten sodium acetate, as it is used to make "salt and vinegar flavored" potato chips.

- Source: Chemical & Engineering News: hand warmers
- Source: Science Buddies: sodium acetate hand warmers

http://lsned.com/facts/hand-warmer-chemistry/

Where Does the Heat Come From?

The article mentions two types of hand warmers that are commonly sold in stores. In this experiment you will investigate the possibility of yet another type of hand warmer using different ionic solids dissolved in water. When a solution forms between an ionic solid and water several events must occur. First, the ionic bonds in the solid must be broken and new attractions between water molecules and the ions of the solid must form. Remember that breaking bonds is *always* an endothermic process while forming new bonds and particulate attractions (such as between ions and water) is an *exothermic* process.

When the amount of energy required to break the bonds is greater than the amount of energy released when forming the new attractions the overall solution process absorbs energy in the form of heat (endothermic, having a positive enthalpy value). The overall solution process will release heat when the amount of energy released from the new attractions is greater than the amount of energy needed to break the ionic bonds of the solid (exothermic, having a negative enthalpy value). The entropy change of solution formation is often positive because when the pure solid and liquid are mixed there are many more positional microstates (more disorder). Also note, there are some cases where the solution actually becomes more ordered!

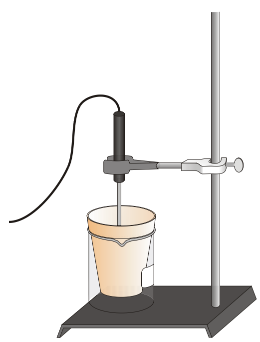

Figure 1

The solution formation is conducted within a calorimeter using two polystyrene cups nested in a beaker, as shown in Figure 1. By measuring the initial temperature of the water and the final temperature of the solution, the enthalpy of solution can be determined.

In this experiment, you will mix a known amount of solid with a known volume of water at constant pressure.

The following relationships apply: $q = mc_p\Delta T$ and $\Delta H_{solution} = \dfrac{q}{moles\ \ solute}$

OBJECTIVES

In this experiment, you will

- Calculate the calorimeter constant.
- Measure the temperature change of the dissolution of a variety of salts in water.
- Calculate the heat of dissolution $\Delta H_{solution}$, in kJ/mol_{rxn} solute of each salt.
- Determine which substance would make the best hand warmer.
- Compare your calculated enthalpies of solutions with the accepted values.

MATERIALS

Data Collection Mechanism
computer or handheld
Temperature Probe
Polystyrene cup calorimeter (2 cups per group)
lid for calorimeter
magnetic stirrer or glass stir rod
ring stand
utility clamp
hot plate

calcium chloride solid
sodium carbonate solid
sodium chloride solid
ammonium nitrate
distilled water
two 250 mL beakers or cups
100 mL graduated cylinder

PROCEDURE

Part I: Determining the Calorimeter Constant

1. Obtain and wear goggles.

2. Set up the data collection system. Connect the interface to the computer or handheld with the proper interface cable.

   a. Connect a Temperature Probe to the interface.
   b. Start the data collection program and set up a Time Graph to gather data for ten minutes. The time between samples need not be shorter than 5 seconds.

3. Nest two polystyrene cups inside each other and place them into a 250 mL beaker as shown in *Figure 1*.

4. Use a graduated cylinder to measure 50.0 mL of distilled water. Place the water into the calorimeter you have constructed.

5. Use a graduated cylinder to measure another 50.0 mL of distilled water. Place this sample into a 250 mL beaker and heat to *approximately* 50.0°C with occasional stirring.

6. Record the initial temperature of *both* the cold water and the hot water *separately*.

7. Start the data collection program and allow the program to graph a few initial temperature readings, *and then* quickly pour the hot water from the beaker into the water in the calorimeter.

8. You may stop data collection soon after a maximum temperature is reached. Record the highest temperature of the water mixture and save or print a copy of the graph to use for data and analysis.

9. Repeat this procedure if time allows.

DATA TABLE PART I

Record all temperatures in °C	Trial 1	Trial 2
Initial temperature cold water		
Initial temperature hot water		
Maximum temperature of the mixture		
Temperature change, $\Delta T_{\text{hot water}}$		
Temperature change $\Delta T_{\text{cold water}}$		

Part II: Determining the Heat of Solution

1. Use the same data collection setup as Part I, but modify the time graph to gather data for ten minutes.

2. Measure out 50.0 mL of distilled water and place into the calorimeter.

3. Measure out approximately 4 grams of the ionic solid assigned by your teacher. *Record the exact mass*.

4. Start the data collection and allow the program to graph a few initial temperature readings and quickly pour the solid into the calorimeter. Stir using either a glass rod or the stir bar and magnetic stirrer as provided by your teacher.

5. Record the most extreme temperature, high or low, obtained during solution formation as the final temperature and save or print a copy of the graph to use for data and analysis.

6. Dispose of the solution as instructed by your teacher. If time allows run a second trial.

7. Repeat this procedure with a different ionic solid. If time allows run a second trial.

8. Share your results and obtain class data with respect to each solute tested as instructed by your teacher.

DATA TABLE PART II

Name of Ionic Solid: _________	Trial 1	Trial 2
Volume of water (mL)		
Mass of solid (g)		
Initial temperature of water (°C)		
Extreme temperature of solution (°C)		
Temperature change, ΔT		

Class Data

NaCl							
GROUP NUMBER:		1	2	3	4	5	6
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
GROUP NUMBER:		7	8	9	10	11	12
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						

Class Data

Na_2CO_3							
	GROUP NUMBER:	1	2	3	4	5	6
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	ΔT (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	ΔT (°C)						
	GROUP NUMBER:	7	8	9	10	11	12
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	ΔT (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	ΔT (°C)						

Class Data

NH$_4$NO$_3$							
	GROUP NUMBER:	1	2	3	4	5	6
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
	GROUP NUMBER:	7	8	9	10	11	12
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						

Class Data

$CaCl_2$							
	GROUP NUMBER:	1	2	3	4	5	6
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
	GROUP NUMBER:	7	8	9	10	11	12
Trial 1	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						
Trial 2	Volume of water (mL)						
	Mass of solid (g)						
	Δ*T* (°C)						

PRE LAB QUESTIONS

Review the article from *Learn Something New Every Day* and view the animation located at the following website: http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/thermochem/solutionSalt.html

1. After observing the animation of an ionic solid dissolving in water, **describe** the interactions between the *anion* and the water molecule. Illustrate this interaction by sketching a diagram. Be sure to include charges where appropriate.

2. **Describe** the interactions between the *cation* and the water molecule. Illustrate this interaction by sketching a diagram. Be sure to include charges where appropriate. What type of intermolecular attractive force exists between ions and the water molecule?

3. Two types of hand warmers were discussed in the article each using a different process for generating heat. Identify each process as either a chemical or physical change and justify your answer.

4. For each term in the equation $q = mc_p\Delta T$, identify what it measures or signifies as well as its unit of measure.

5. What determines whether a given solute will form a solution in a given solvent?

6. Describe the three steps of solution formation. Be sure to include the details regarding dissolving of the substance, solution formation and the intermolecular interactions involved in each step.

7. Identify the hand warmer prepared in this experiment as a physical or chemical process. Justify your choice.

8. Why it is thermodynamically possible for some ionic solids to dissolve even though the solution process is endothermic?

9. A student conducts an experiment to determine the enthalpy of solution for lithium chloride dissolved in water. The student combines 5.00 grams of lithium chloride with 100.0 mL of distilled water. The initial temperature of the water is 23.0°C and the highest temperature after mixing reaches 33.0°C. Assume a density of 1.00 g/mL and a specific heat of 4.18 $\frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}$. The calorimeter constant was found to be 20.0 $\frac{\text{J}}{{}^\circ\text{C}}$.
   (a) Is the reaction endothermic or exothermic? Justify your answer.
   (b) Calculate the temperature change of the solution in the calorimeter.
   (c) Calculate the total mass of the solution in the calorimeter.
   (d) Calculate the energy change of the solution.
   (e) Calculate the energy change of the calorimeter.
   (f) Calculate the total $\Delta H_{solution}$ in kJ/mole

POST LAB QUESTIONS AND DATA ANALYSIS

Part I

1. Describe the energy transfer occurring in the polystyrene cup when the hot and cold water are mixed.

2. Calculate the quantity of energy, q of the hot water. Is this energy absorbed or released?

3. Calculate the energy (q) for the cold water. Is this energy absorbed or released?

4. According to the Law of Conservation of Energy, the energy released and the energy absorbed should be equal in magnitude ($q_{\text{hot}} = -\,q_{\text{cold}}$). Explain the difference in energy between the two values calculated.

5. Calculate the calorimeter constant, C, in J °C^{-1}.

Part II

1. Calculate the total amount of heat energy, q, for each trial of each solution. Just as in the hot and cold water, q_{solution} and q_{rxn} are assumed to be equal in magnitude but opposite in sign, $\boldsymbol{q_{\text{rxn}} = -\,q_{\text{solution}}}$. Assume the density of the solutions to be that of water. Also assume the specific heat of each solution to be that of water.
   In calculating total enthalpy of the solution remember to include the amount of heat energy contributed by the calorimeter. $\boldsymbol{q_{\text{solution}} = -\,(q_{\text{rxn}}} + C\Delta T)$ Be sure to identify each salt used.

2. Calculate the number of moles of salt used for each trial.

3. Use the heat energy calculated in question 1 along with the moles calculated in question 2 to determine the enthalpy change, $\Delta H_{solution}$ in kJ/mol for each ionic solid that you tested. Calculate the average enthalpy change for each solid if more than one trial was performed.

4. Write the balanced net ionic equation for the dissolution of each ionic solid that you tested. Include the $\Delta H_{solution}$ calculated in question 3 on the appropriate side (reactant or product) of your equation.

5. Using the class data arrange the ionic solids in order of least to best possible solutions for a hand warmer. Justify your ranking.

6. Calculate the percent discrepancy between the accepted values (published values) of the ΔH of solution for each salt you tested with your experimental values. Cite possible reasons for any discrepancies.

7. A student conducts two trials of each experiment, but neglects to rinse the calorimeter cup between trials. What effect, if any, would this error have on the calculated value of the ΔH of solution?