

Designing a Hand Warmer Lab

AP Chemistry

PROBLEM

The purpose of this advanced inquiry lab is to design an effective hand warmer using one of seven different solids that is inexpensive, nontoxic, and safe for the environment.

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* (ΔH_{soln}). At constant pressure, this enthalpy change, ΔH_{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 (ΔH_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 (ΔH_{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:

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

where m is the **total mass of the solution** (solute plus solvent), C is the **specific heat** of the solution, and ΔT 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 \cdot °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_{soln}) for the reaction (the system) then becomes:

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

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

$$T_{\text{avg}} = (T_H + T_C)/2 \qquad \text{Equation 3}$$

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Annelies Wilder-Smith



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