Designing a Hand Warmer

Introduction

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 of 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 tempe4rature 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), 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 (ΔH1 = +C kJ/mole), the energy required to disrupt the intermolecular forces between water molecules, (ΔH2 = +D kJ/mole) and the energy released with the dissociated (free) ions form ion-dipole attractive forces with the water molecules(ΔH3 = -F kJ/mole) The overall process can be represented by the following equation:

**MaXb(s) 🡪 aM+(aq) + bX-(aq) ΔHsoln = ΔH1 + ΔH2 + ΔH3= (+ C + D –F) kJ/mole**

If the amount of energy released in the formation of the hydrated ions (ΔH3) is greater than the amount of energy required to separate the solute and solvent particles (ΔH1 + ΔH2), then the sum (ΔHsoln) of the energy changes will be negative and the solution process exothermic (releases heat). *The solution temperature will rise*. If the amount of energy released in the formation of the hydrated ions (ΔH3) 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). *The solution temperature will fall*.

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 = mcΔT 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˚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 (qaq). A small amount of the heat will be absorbed by the calorimeter itself (qcal). The overall heat transfer (qsoln) for the reaction (the system) then becomes:

**qsoln = -(qaq + qcal) Equation 2**

In order to determine the correction factor qcal 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/g˚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, Tmix. The average temperature ,Tavg of the initial hot (TH) and cool water (Tc) is also calculated.

**Tavg = (TH + Tc)/2**

The difference between Tavg and Tmix is due to the heat lost by the water and absorbed by the calorimeter. The heat lost by the water, qwater is:

**qwater = (mass of water) X (specific heat of water X (Tmix – T avg) Equation 3**

where the mass is the total mass of hot and cool water. The heat gained by the calorimeter, qcalor, is equal to that lost by the water, but opposite in sign. The calorimeter constant, Ccal, is calculated as follows:

**Ccal = qcalor Equation 4**

**(Tmix – Tinitial)**

Where Tinitial is the initial temperature of the calorimeter containing cool water.

To calculate the correction factor qcal 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. When calculating the qsoln for your experiment.

**qcal = [ΔT(˚C)]Experiment  X [Ccal(J/g˚C)]**