HOT PACKS, COLD PACKS, AND HEATS OF SOLUTION

CCLI Experiments
Computers in Chemistry Laboratory Instruction

LEARNING OBJECTIVES

The objectives of this experiment are to . . .

- use the ***MicroLab*** interface to measure the heats of solution ($\Delta \mathbf{H}_{soln}$, in kJ/mol) of several salts.
- calculate the heats of formation ($\Delta \mathbf{H}_f$, in kJ/mol) of single aqueous ions.
- predict the heats of solution of additional salts.
- design a hot pack and a cold pack for specified temperature changes.

BACKGROUND

All chemical reactions and many physical processes are accompanied by the absorption or release of heat. Your body temperature of 37 °C is maintained by heat released from chemical reactions occurring inside of cells. Sometimes it is medically desirable to temporarily increase or decrease the temperature of a certain body part, such as in the treatment of sprains or bruises. These temperature changes can be achieved with commercially available hot packs or cold packs in which the dissolving of a salt in water either generates or absorbs heat. In this experiment you will measure the heat changes which occur when various salts are dissolved in water and use that information to design both a hot pack and a cold pack capable of producing specified temperature changes.

Measuring heats of solution

The heat of solution, ΔH_{soln}, is the heat that is generated or absorbed when a salt is dissolved in water and is expressed in kJ/mol solute. ΔH_{soln} is negative for an exothermic solution process and positive if it is endothermic. Thermochemical equations for the solution process are commonly shown as follows:

$$KCl\,(s) \longrightarrow K^+(aq) + Cl^-(aq) \qquad \Delta H_{soln} = +17.5 \text{ kJ/mol} \qquad (1)$$

The experimental measurement of ΔH_{soln} is done in a calorimeter. For our experiment, the calorimeter will consist of two nested styrofoam cups containing 100 grams of water, a magnetic stirring bar, and a temperature probe inserted through a styrofoam lid. Five grams of a salt will be added and the temperature change of the resulting solution measured.

Consider the dissolving of KCl, which is endothermic. Since the calorimeter is assumed to be an adiabatic system (which means that it is a constant temperature system), the heat absorbed in the solution process (q_{soln}) must be equal but opposite in sign to the heat lost by the rest of the calorimeter (q_{cal}). The relationship for these heat changes can thus be written as follows.

$$q_{soln} = -q_{cal} \tag{2}$$

The heat change for the calorimeter (q_{cal}) equals the measured temperature change ($\Delta T = T_f - T_i$) times the heat capacity (C_p) of the calorimeter.

$$q_{cal} = C_p \times \Delta T \tag{3}$$

The heat capacity (C_p) for the calorimeter used in this experiment has been determined and has a value of 463 J/°C. Finally, ΔH_{soln} is calculated by dividing q_{soln} by the moles of salt and expressing this final result in kJ/mol.

ΔH_f of single aqueous ions and predicting ΔH_{soln}

Another aspect of this experiment will be the use of Hess's Law to evaluate ΔH_f for single aqueous ions. To illustrate this process, consider the dissolving of KCl(s) as shown in equation (1). ΔH_{soln} can be expressed as follows.

$$\Delta H_{soln} = \Delta H_f(K^+) + \Delta H_f(Cl^-) - \Delta H_f(KCl) \tag{4}$$

ΔH_f values for the aqueous ions can be determined if ΔH_{soln} and ΔH_f for the solid are known. For this exercise, ΔH_f values for the appropriate solids are given in Table 1, and the ΔH_{soln} values will be obtained experimentally. This leaves two unknowns in equation (4), the ΔH_f values for the two aqueous ions. It is impossible to have only a single ion type in solution since there must always be both cations and anions. Therefore, the absolute value of ΔH_f for a single ion cannot be known. The accepted approach is to arbitrarily assign $\Delta H_{f\,(H+)}$ a value of 0 kJ/mol. The ΔH_f's for all other aqueous ions are then expressed relative to that for $H^+_{(aq)}$. For this experiment, you will use the literature value for $\Delta H_{f\,(K+)}$ of - 251.2 kJ/mol as your reference. ΔH_f values for other aqueous ions, such as $Cl^-_{(aq)}$, can then be calculated.

$$\begin{aligned} \Delta H_f(Cl^-) &= \Delta H_{soln} + \Delta H_{f\,(KCl)} - \Delta H_{f\,(K+)} \\ &= 17.5 + (-436.7) - (-251.2) \\ &= -168.0 \text{ kJ/mol} \end{aligned} \tag{5}$$

Finally, ΔH_{soln} values can be predicted from ΔH_f data for ions and solid salts. This involves the application of equation (4) as applied to any salt.

Table 1 Heats of Formation, ΔH_f^o (kJ/mol)			
KCl (s) $MgSO_4$ (s) K_2SO_4 (s) KNO_3 (s)	- 436.7 - 1284.9 - 1433.4 - 492.7	NH_4NO_3(s) $Mg(NO_3)_2$(s) $(NH_4)_2SO_4$(s)	- 365.6 - 789.5 - 1179.5

SAFETY PRECAUTIONS

As usual, any skin contacted with chemicals should be washed immediately. Safety goggles must be worn in the lab at all times.

BEFORE PERFORMING THIS EXPERIMENT . . .

...you will need a ***MicroLab*** program capable of measuring and displaying temperature versus time. See your instructor if you don't know how to do this.

EXPERIMENTAL PROCEDURE

Temperature probe calibration and experiment program
Calibrate your temperature probe using an ice/water mixture for the lower temperature calibration and hot water for the upper temperature calibration. When doing experimental runs, readings should be taken every five seconds.

Running the experiment
Obtain two styrofoam cups (one nested in the other for extra insulation), a cup lid, a temperature probe, a magnetic stir bar, and a 400 ml beaker or a three or four inch ring clamp and ring stand to support the styrofoam cups for increased stability. Cut a notch into the lid so that you can clamp the temperature probe through the cup lid to a depth such that its tip goes deep into the cup but misses the stir bar.

Heat of solution for $MgSO_4$(s)

1. Obtain 100 ml of distilled water in a graduated cylinder. Pour the water into the cup, add the stir bar, position the cups and supporting beaker on the magnetic stirrer and adjust the stirring rate too vigorous but without splashing. Finally, insert the temperature probe and allow several minutes for temperature equilibration.

2. Tare a weighing boat on the balance pan, then weigh approximately five grams of $MgSO_{4(s)}$ on a top-loading balance. Weigh to the nearest 0.01 g. Record this mass in your lab notes.

3. Start the experiment program. You will be performing four experiments and thus creating four data files. Be sure to give them different but distinguishable file names. As the experiment starts, you should see successive, constant temperature readings of near room temperature. Allow these readings to continue for at least one to two minutes to establish an accurate temperature baseline value.

4. Add the $MgSO_{4(s)}$ continuously and smoothly into the cup, and reposition lid assembly. As the dissolution occurs, you should observe the temperature climb. Continue taking data until the temperature plateau is well established, or a well established downward slope is obtained. The experiment can be stopped at that point.

5. Repeat steps one through four for each of the remaining salts: $K_2SO_{4(s)}$, $KNO_{3(s)}$, and $NH_4NO_{3(s)}$.

DATA ANALYSIS

1. Import your first data file onto **EXCEL** and construct the temperature vs. a time graph. From examination of both the raw data and the temperature vs. a time graph, decide on the most accurate values of initial temperature and final temperature. If a slight downward or upward trend appears on the final temperature plateau, use the maximum value achieved. Record these temperatures (to the nearest 0.1 °C) in your report. Repeat this procedure for the three remaining data files. (Your instructor may want to teach you how to do a graphical method that will produce more accurate values for ΔT. If so, this can be done using the ***MicroLab*** analysis tools, then use this data for the calculations.)

2. Calculate the ΔH_{soln} for each of the four salts and record the results in your report. Show your calculations for each salt.

3. Calculate the ΔH_f for Mg^{2+}, NH_4^+, SO_4^{-2}, and NO_3^-. Show your calculations for each ion.

   Calculate the ΔH_{soln} for both $Mg(NO_3)_{2(s)}$ and $(NH_4)_2SO_{4(s)}$. Show your calculations.

4. Based on your results, design a cold pack using $NH_4NO_{3(s)}$. It should have a total mass of 400 g (salt plus water) and be able to achieve a temperature of 10 °C. Assume a specific heat for the solution of 4.18 J/g °C (same as water) and that the water in the cold pack is initially 23 °C. Show your calculations.

5. Using the same guidelines given for the cold pack, design a hot pack (using $MgSO_{4(s)}$ that will reach a temperature of 60 °C. Show your calculations.