Dissolving salts

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INTRODUCTION:

We know that spontaneous chemical processes are those that produce total entropy. There are two ways to produce entropy: 1) in the system or 2) in the surroundings. If you heat the surroundings, it will produce entropy in the surroundings. Therefore, exothermic reactions (negative values of the reactive enthalpy change) tend to be spontaneous. Spontaneous dissolution processes are interesting because, they may be either exo- or endothermic. If they are endothermic, they decrease the entropy in the surroundings, thus they must increase the entropy in the system by a larger magnitude to be overall spontaneous. One would expect that dissolution is inherently going to increase the entropy of the system, but it turns out that the ions in solution make a partially ordered cage of water molecules around them, which acts to counteract the increase in entropy upon breaking up the solid's crystal lattice. Thus the entropy change can be either positive or negative.

In this demonstration laboratory, you will measure the heat of dissolution of a series of reactions that vary from exothermic to endothermic.

INSTRUCTIONS:

Please read through the entire instructions before you start the experiment.

Each calorimetry experiment will consist of placing 100ml of water in a set of two nested styrofoam cups, and recording the water's initial temperature, Ti. Then weigh out enough salt to make a 1 Molar solution in this volume of water. Add this salt to the water, stir and record the most extreme (highest or lowest) temperature reached as the final temperature, Tf. Place the solution in the appropriate waste container. Use the heat capacity of liquid water (at constant pressure, which is 418 J / K for this 100g water sample) to calculate the heat of dissolution for the salt.

Be careful with the NaOH (it is corrosive), and the $NaNO_3$ (causes skin irritation). Rinse off immediately if you contact either of these substances. Dispose of both of these compounds in the labeled waste containers. The NaCl is table salt; it can go down the drain.

Do the experiment with 4.0 g NaOH (makes a 1M solution in 100ml)

Do the experiment with 5.8 g NaCl (makes a 1M solution in 100ml)

Do the experiment with 8.5 g $NaNO_3$ (makes a 1M solution in 100ml)

QUESTIONS:

1) The table below lists the $\Delta_r S_m$ for these reactions:

Compound	$\Delta_r S_m$ / (J mol^{-1} K^{-1})
NaOH	-4.8
NaCl	+43.2
$NaNO_3$	+90.6

Combine this entropy change with your measured enthalpy change to predict the Gibbs energy of reaction for these processes. Are they all predicted to be spontaneous?

2) Calculate $\Delta_r S_m$ (Surr) = - $\Delta_r H_m$ / T and then the total entropy change. Does this agree with the Gibbs energy method for prediction of spontaneous reaction?

3) Silver Chloride (AgCl) has a very endothermic dissolution reaction, $\Delta_r H_m$ = +66.0 kJ mol^{-1}. It also produces entropy, $\Delta_r S_m$ = +32.9 J mol^{-1} K^{-1}. Will it dissolve?

4) For a dissolution reaction that has $\Delta_r S_m$ = +100 J mol^{-1} K^{-1}, what is the maximum heat (pre mole) that could be taken up in a spontaneous process?