

## Enthalpy of Dissolving and Solvation

### Using $\Delta H_f$ values to Calc $\Delta H_{\text{dissolve}}$ and Compare to Experimental Values

#### Introduction

Enthalpy of formation (heat of formation,  $\Delta H_f^\circ$ ) is the enthalpy change associated with the reaction for the formation of a substance from its constituent elements. A set of values has been tabulated for a large number of compounds at a defined set of conditions called the standard state. The standard state conditions are 1 atm pressure and 25°C.

Thus  $\Delta H_f^\circ$  (standard enthalpy of formation) is the change in enthalpy for the reaction that forms one mole of a particular compound from its elements, with all substances in their standard states.

We can use values to calculate *theoretical* heats of particular reactions. ( $\Delta H$  values are located in your text pg 1041 or you can find a set of tables on chap 5 document page.)

*Experimentally* we can calculate energy transferred by measuring temperature changes. Remember the heat equation below:

$$q = c * m * \Delta T$$

heat lost or gained = specific heat capacity \* mass \* change in temperature

For a chemical reaction (or dissolving or phase change)

$$q = \Delta H * \text{moles}$$

If you put these two equations together:

$$c * m * \Delta T = \Delta H * \text{moles}$$

In this LAD, an ionic compound will be dissolved into solution. There is an energy change associated with this process. We will determine the molar value for this process and then compare that experimental value with the calculated theoretical value.

#### PreLAD

- Write out the equation that represents the dissolving process of ammonium nitrate.
- Write out the equation that represents the dissolving process of calcium chloride.
- Calculate the  $\Delta H_{\text{dissolve}}$  for the two equations above using the  $\Delta H_f^\circ$  values in your thermodynamic tables distributed in class and also on the chapter 5 documents page.

#### Procedure for determining the $\Delta H_{\text{dissolve}}$

- Set up your stirring plate apparatus.
- Measure a mass (around 100 ml or so, you need to know exactly) of tap water into your foam cup calorimeter.
- Using the plastic weighing boats, measure out a mass of ammonium nitrate (around 5+ g or so).
- Record the starting temperature of the water.
- Dump all of the ammonium nitrate into the water, cover, turn on the stirrer, then watch the temperature until thermal equilibrium is reached. Record the temperature at thermal equilibrium.
- The solution can be flushed down the drain.
- Repeat the procedure for the second chemical.

#### Process the Data

- Calculate the energy lost/gained by the water. Using the  $q_{\text{water}} = SHC_{\text{water}} * m_{\text{water}} * |T_{\text{f mix}} - T_{\text{start}}|$ 
  - What sign should be attached to the  $q$  for the water?
- Assume that the energy lost/gained by the water, is the same (though lost vs gained or gained vs lost) by the dissolving compound.
  - What sign should be attached to the  $q$  for the dissolving process?
- The energy calculated in number 3 is for a very particular mass of ammonium nitrate (calcium chloride), convert this energy value in number 3 to a per mole value.
- Compare your experimental  $\Delta H_{\text{dissolve}}$  (in #4) to the theoretical  $\Delta H_{\text{dissolve}}$  that you calculated in the preLAD and calculate your percent error.

**Questions to Ponder**

1. Imagine the same procedure using twice the mass of salt to be dissolved.
  - a. How would  $q_{water}$  change?
  - b. How would  $\Delta H_{dissolve}$  change?
  
2. Imagine the same procedure using the same amount of salt, but instead using twice the volume of water.
  - a. How would  $q_{water}$  change?
  - b. How would  $\Delta H_{dissolve}$  change?
  
3. Imagine the same procedure using the same amount of salt, and same amount of water, but if the water were 10°C warmer than during your procedure.
  - a. How would  $q_{water}$  change?
  - b. How would  $\Delta H_{dissolve}$  change?
  
4. Suppose that during your procedure for calcium chloride that a significant amount of heat were lost out the top of the calorimeter, what effect would this have on your calculated molar enthalpy of dissolving,  $\Delta H_{dissolve}$ .