Lecture to Lab Problem

In a coffee cup calorimeter, 1.60 g ammonium nitrate was mixed with 75.0 g of water at an initial temperature 25.00 ^{o}C. After dissolution of the salt, the final temperature of the calorimeter contents was 23.34^{o}C.

a) If the solution had a heat capacity of 4.184 J/g^{o}C, and no heat was lost to the calorimeter, calculate the enthalpy of solution (ΔH_{soln}) for the dissolution of ammonium nitrate in kJ/mol.

b) If the enthalpy of hydration for ammonium nitrate is -630 kJ/mol, calculate the lattice energy for ammonium nitrate.

Lecture to Lab Problem -2

In a coffee cup calorimeter, 1.60 g ammonium nitrate was mixed with 75.0 g of water at an initial temperature 25.00 ^oC. After dissolution of the salt, the final temperature of the calorimeter contents was 23.34^oC.

a) If the solution had a heat capacity of 4.184 J/g^oC, and no heat was lost to the calorimeter, calculate the enthalpy of solution (ΔH_{soln}) for the dissolution of ammonium nitrate in kJ/mol.

b) If the enthalpy of hydration for ammonium nitrate is -630 kJ/mol, calculate the lattice energy for ammonium nitrate.

(A) Step 1: calorimetry means $q=ms\Delta T$ so we must first calculate the total mass of products that we are measuring. 75.0 g of water + 1.60 g of salt makes 76.6 g of solution product. ΔT = f-i =23.34 – 25 = -1.66 and the solution is very dilute so the specific heat is the same as the solvent, water, 4.184 J/g^oC.

$$q_{Rx} = - q_{soln} = ms\Delta T = 76.6 \text{ g } (4.184 \text{J/g}^oC)(-1.66^oC) = -(-532.0 \text{ J})$$

Step 2: to calculate ΔH_{soln} we need the measurement in kJ/mol so…

$$\Delta H_{soln} = q_{Rx} / n$$

$$n = 1.60 \text{ g } NH_4NO_3 \text{ (1 mol/80 g)} = 0.02 \text{ moles}$$

$$\Delta H_{soln} = 0.532\text{kJ} / 0.02 \text{ mol} = \mathbf{26.6 \text{ kJ/mol}}$$

(B) $\Delta H_{soln} = \Delta H_{hyd.} – \Delta H_{lat}$ therefore $\Delta H_{lat} = \Delta H_{hyd.} – \Delta H_{soln}$

ΔH_{lat} = -630 kJ/mol – 26.6 kJ/mol = -657 kJ/mol

Lecture example

In a coffee cup calorimeter a different student made the following measurements, 0.988 g ammonium nitrate was mixed with 35.0 g of water at an initial temperature 20.00 ^{o}C. After dissolution of the salt, the final temperature of the calorimeter contents was 16.82^{o}C.

a) If the solution had a heat capacity of 4.184 $J/g^{o}C$, and no heat was lost to the calorimeter, calculate the enthalpy of solution (ΔH_{soln}) for the dissolution of ammonium nitrate in kJ/mol.

b) If the enthalpy of hydration for ammonium nitrate is -630 kJ/mol, calculate the lattice energy for ammonium nitrate.

Expressing Solution Concentration

Solution Concentration

Earlier we learned about one method for expressing solution concentration—molarity. We defined molarity as moles of solute/liters of solution. Now we will describe several other methods of expressing concentration, each of which serves a different purpose.

A. mass percentage

$$\textbf{Mass \% of component} = \frac{\text{Mass of component in solution}}{\text{total mass of solution}} x100$$

B. parts per million (for dilute solutions)

$$\textbf{ppm of component} = \frac{\text{Mass of component in solution}}{\text{total mass of solution}} x10^6$$

C. parts per billion (for even MORE dilute solutions!)

$$\textbf{ppb of component} = \frac{\text{Mass of component in solution}}{\text{total mass of solution}} x10^9$$

D. mole fraction (denoted with the symbol χ)

$$\textbf{Mole fraction of component} = \frac{\text{Moles of component}}{\text{total moles of components}}$$

E. molality (denoted with the lowercase letter *m*)

$$\textbf{molality} = \frac{\text{moles solute}}{\text{kilograms of solvent}}$$

The molality of a given solution does NOT vary with temperature because masses do not vary with temperature. Molarity, however, changes with temperature because the expansion or contraction of the solution changes its volume. Thus, molality is often the concentration unit of choice when a solution is to be used over a range of temperatures.

Lecture Problem - Concentration

1) An aqueous antifreeze solution is 40.0% ethylene glycol ($C_2H_6O_2$) by mass, the density of the solution is 1.06 g/mL. Calculate the molarity, molality, and mole fraction of ethylene glycol.

STEP 3: mole fraction – since χ is independent of volume, you can start with density units g/mL.
the d_{soln} is 1.06 g/mL and only 40% of it is due to the solute, the other 60% is from the solvent. Ignoring IMF issues, 40% of 1.06 g/mL = 0.424 g/mL and that leaves 0.636 g/mL for the solvent.

So to calculate the χ of solute – simply make the volume 1 mL therefore 0.424 g solute/62 g/mol = 0.0068387 mol of solute & 0.63594 g /(18 g/mol) of solvent makes 0.03533 moles of solvent!

χ = 0.0068387 mol / (0.03533 + 0.0068387mol)
= **0.162 mole fraction!**

Lecture Problem – Concentration p2

1) An aqueous antifreeze solution is 40.0% ethylene glycol ($C_2H_6O_2$) by mass, the density of the solution is 1.06 g/mL. Calculate the molarity, molality, and mole fraction of ethylene glycol.

STEP 4: **molality** – Again list all relevant knowledge/equations.

$\mathbf{d_{soln} = m_{soln} / V_{soln}}$ $\quad$ $\% = (m_{solute} / m_{total})100$ $\quad$ $\mathbf{M = n_{solute} / V_{soln}}$

$\mathbf{MM = m_{solute} / n_{solute}}$ $\quad$ $\boldsymbol{m} \mathbf{= n_{solute} / kg_{solv}}$

STEP 2: molality = n_{solute} / kg_{solv}

$= 0.0068387\text{ mol} / 0.63594 g_{solv}\ (1\text{kg}/1000\text{g}) = \mathbf{10.8}\ \boldsymbol{m}$

Practice/lecture Problem - Concentration

1) A 1.37 M solution of citric acid ($H_3C_6H_5O_7$) in water has a density of 1.10 g/mL. Calculate the mass percent, molarity, & mole fraction of citric acid. Citric acid is triprotic.

Lecture to LAB – Redox titration

1) The vanadium in a sample of ore is converted to VO^{2+}. The VO^{2+} ion is then titrated with potassium permanganate in an acidic solution to form $V(OH)_4^+$ and manganese(II) ion. The unbalanced reaction is:

$$MnO_4^- + VO^{2+} \rightarrow V(OH)_4^+ + Mn^{2+}$$

To titrate the solution, 26.45 mL of 0.02250 M permanganate was required. If the mass percent of vanadium in the ore was 58.1% what was the mass of the ore in the sample?

Lecture to LAB – Redox titration answer

STEP 1: Balance the redox reaction using the steps learned in chapter 4

$$3H_2O + VO^{2+} \rightarrow V(OH)_4^{+} + 2\ H^{+} + 1e^-$$
$$5e^- + 8\ H^{+} + MnO_4^{-} \rightarrow Mn^{2+} + 4\ H_2O$$
$$\mathbf{11\ H_2O + 5\ VO^{2+} + MnO_4^{-} \rightarrow 5\ V(OH)_4^{+} + 2\ H^{+} + Mn^{2+}}$$

STEP 2: Volumetric Analysis!

$$\mathbf{11\ H_2O + 5\ VO^{2+} + MnO_4^{-} \rightarrow 5\ V(OH)_4^{+} + 2\ H^{+} + Mn^{2+}}$$

V=26.45 mL

M=0.02250 mol/L

n = 5.95125 x10^{-4} (5 mol VO^{2+} / 1 mol MnO_4-) =

2.9756 x 10^{-3} mol VO^{2+} (1mol V / 1mol VO^{2+}) (54.94 g / mol) = 0.16348 g V

58.1 % = ($mass_V$ / $mass_{ORE}$) 100 =

0.581 = 0.16348g / $mass_{ORE}$

Mass of ore = 0.2814 g

<u>"Proof of Knowledge" on Concentration</u>

1. A solution of hydrochloric acid contains 36% HCl by mass.

A. Calculate the mole fraction of HCl in the solution.

B. Calculate the molality of HCl in the solution.

2. A solution contains 5.0 g of toluene (C_7H_8) and 225 g of benzene and has a density of 0.876 g/mL. Calculate the molarity of the solution.

3. Calculate the mol of CO_2 that will dissolve in enough water to form 900 mL of solution at 20 °C if the partial pressure of CO_2 is 1.00 atm. *<u>NOTE</u>:* k for CO_2 in water at 20 °C = 2.3×10^{-2} mol $L^{-1}atm^{-1}$