AP CHEMISTRY

TOPIC 12: SOLUTIONS, REVIEW, PART I

Day 136:

1. Lead(II) bromide dissociates in water and has a K_{sp} = 4.6 x 10^{-6} at 25^0C
   a) Write the chemical dissociation equation for the above dissociation.

$$PbBr_2 \leftrightarrow Pb^{+2} + 2\ Br^{-1}$$

   b) Write the equilibrium expression

$$K_{sp} = [\ Pb^{+2}\ ]\ [\ Br^{-1}\ ]^2$$

   c) Calculate the concentration, in mol L^{-1} of the bromide ions in a saturated solution of lead(II) bromide at 25^0C.

	$PbBr_2$	$\leftrightarrow$	Pb^{+2}	+	$2\ Br^{-1}$
I	-		0		0
C	-		+ x		+ $2x$
E	-		x		$2x$

$$4.6 \times 10^{-6} = [\ Pb^{+2}\ ]\ [\ Br^{-1}\ ]^2$$

$$4.6 \times 10^{-6} = [\ x\ ]\ [\ 2x\ ]^2 = 4x^3$$

$$x = 1.05 \times 10^{-2}\ M = [\ Pb^{+2}\ ]$$

$$[\ Br^{-1}\ ] = 2x = (\ 2\ )\ (1.05 \times 10^{-2}\ M\ ) = 2.10 \times 10^{-2}\ M$$

   d) Calculate the maximum mass, in grams, of lead(II) bromide that can dissolve in 4297 mL of water at 25^0C.

$$\frac{4297\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{1.05\times10^{-2}\ mol\ Pb^{+2}}{L} \times \frac{1\ mol\ PbBr_2}{1\ mol\ Pb^{+2}} \times \frac{367\ g}{1\ mol\ PbBr_2} = 16.6\ g$$

2. In a saturated solution of silver phosphate at 25^0C, the concentration of PO_4^{-3} (*aq*) is 1.6069 x 10^{-5} *M*. The equilibrium constant expression for the dissolving of silver phosphate in water is shown below:

$$K_{sp} = [\ Ag^{+1}\ ]^3\ [\ PO_4^{-3}\ ]$$

   a) Write the balanced equation for the dissolving of silver phosphate in water.

$$Ag_3PO_4 \leftrightarrow 3\ Ag^{+1} + PO_4^{-3}$$

   b) Calculate the value of K_{sp} for silver phosphate at 25^0C.

	Ag_3PO_4	$\leftrightarrow$	$3\ Ag^{+1}$	+	PO_4^{-3}
I	-		0		0
C	-		+ $3x$		+ x
E	-		$3x$		x

$$[\ PO_4^{-3}\ ] = 1.6069 \times 10^{-5}\ M = x$$

$$[\ Ag^{+1}\ ] = \frac{1.6069\times10^{-5}\ mol}{L} \times \frac{3\ mol\ Ag^{+1}}{1\ mol\ PO_4^{-3}} = 4.821\times10^{-5}\ M = 3x$$

$$K_{sp} = (\ 4.821 \times 10^{-5}\ )^3\ (\ 1.6069 \times 10^{-5}\ ) = 1.8 \times 10^{-18}$$

3. A solution is prepared by dissolving 144 grams barium hydroxide ($K_{sp} = 5.0 \times 10^{-3}$) in 3500 mL of water. Calculate the freezing point of this solution. (Hint: you have a K_{sp} value which indicates that the salt is NOT strong and that this salt does not dissociate to 100% - calculate the number of moles that do dissociate.)

$$K_{sp} = [Ba^{+2}][OH^{-1}]^2$$

	$Ba(OH)_2$	$\rightleftharpoons$	Ba^{+2}	+	$2\ OH^{-1}$
I	-		0		0
C	-		$+x$		$+2x$
E	-		x		$2x$

$$5.0 \times 10^{-3} = [Ba^{+2}][OH^{-1}]^2$$

$$5.0 \times 10^{-3} = [x][2x]^2 = 4x^3$$

$$\frac{5.00\times10^{-3}}{4} = 0.00125 \;;\; x = \sqrt[3]{0.00125} = 0.10772 = \left[Ba^{+2}\right]$$

$$\frac{3500\ mL_{H_2O}}{} \times \frac{1\ g_{H_2O}}{1\ mL_{H_2O}} \times \frac{1\ kg_{H_2O}}{1000\ g_{H_2O}} \times \frac{0.10772\ mol\ Ba^{+2}}{1\ kg_{H_2O}} \times \frac{1\ mol\ Ba(OH)_2}{1\ mol\ Ba^{+2}} = 0.377\ mol\ Ba(OH)_2$$

$$\Delta T_f = i\,k_f\ molality \;;\; \Delta T_f = (3)\left(\frac{1.86\ kg}{mol}\right)\left(\frac{0.3770\ mol}{3.50\ kg}\right) = 0.601^0 C$$

F.P. for water = 0 ^{0}C – 0.601 ^{0}C = - 0.601 ^{0}C

You do NOTHING with the 144 grams… This is to "distract you" – focus on WHAT the question is asking… This IS a past AP Exam question…

4. A solution containing 365.3 grams of a electrolyte (that dissociates into four particles) dissolved in 1200 grams of water boils at 107.44^{0}C. Calculate the approximate molecular weight of the solute.

$$\Delta T_b = 107.44^0C - 100^0C = 7.44^0C$$

$$1200 \text{ grams} = 1.20 \text{ kilograms}$$

$$\Delta T_b = i\, k_b \; molality$$

$$7.44\ ^0C = \frac{(4)\ 0.51^0 C \cdot kg}{(mol)} \left(\frac{?\ mol}{1.2\ kg} \right)$$

$$?\ mol = \frac{(7.44\ ^0C)(1.2\ kg_{H_2O})(mol)}{(4)\ (0.51^0 C \cdot kg)} = 4.376\ mol$$

$$molecular\ weight = \frac{365.3\ g\ solute}{4.376\ mol\ solute} = 83.5\ \frac{g}{mol}$$

5. What mass of aluminum perchlorate must be dissolved in 3.53 kg of water to give a solution with a freezing point of –13.33 ^{0}C?

$$\Delta T_f = (4)(1.86\ ^0C \cdot kg \cdot mol^{-1})(\text{molality}) = 13.33\ ^0C$$

$$13.33\ ^0C = \frac{(4)\ 1.86^0 C \cdot kg}{(mol)} \left(\frac{?\ mol}{3.53\ kg} \right)$$

$$moles = \frac{(13.33\ ^0C)(3.53\ kg_{H_2O})(mol)}{(4)\ (1.86^0 C \cdot kg)} = 6.32458\ mol$$

$$\frac{6.32458\ mol}{} \times \frac{325.33\ g}{mol} = 2.06 \times 10^3\ g\ Al(ClO_4)_3$$