Name _____ AP Chemistry

Chapter 15 HW #3: Due 12/7/15 Complete both free response questions. One will be graded. Show all work. Box and clearly label all final answers

#1. Answer the following questions that relate to solubility of salts of lead and barium.

(a) A saturated solution is prepared by adding excess $PbCl_2(s)$ to distilled water to form 1.0 L of solution at 25°C. The solubility of $PbCl_2$ is found to be 0.4415 grams/100. ml H_2O. The chemical equation for the dissolution of $PbCl_2(s)$ in water is shown below.

$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$

(i) Write the equilibrium-constant expression for the equation.

(ii) Calculate the molar concentration of $Cl^-(aq)$ in the solution.

(iii) Calculate the value of the equilibrium constant, K_{sp} .

(b) A saturated solution is prepared by adding $PbCl_2(s)$ to distilled water to form 4.0 L of solution at 25°C. What are the molar concentrations of $Pb^{2+}(aq)$ and $Cl^-(aq)$ in the solution? Justify your answer.

(c) Solid NaCl is added to a saturated solution of $PbCl_2$ at 25°C. Assuming that the volume of the solution does not change, does the molar concentration of $Pb^{2+}(aq)$ in the solution increase, decrease, or remain the same? Justify your answer.

(d) The value of K_{sp} for the salt $BaSO_4$ is 1.5×10^{-9}.

(i) When a 117.0 mL sample of 2.16×10^{-2} *M* $Ba(NO_3)_2$ is added to 229.0 mL of 5.19×10^{-2} *M* Na_2SO_4 does a precipitate form (you must justify with calculations).

(ii) Calculate the concentration of the Ba^{2+} at equilibrium.

a. $K_{sp} = [Pb^{2+}][Cl^-]^2$

ai $0.4415\,g \times \frac{1\,mol}{278.2\,g} \times \frac{1000\,mL}{100\,mL} = 0.0159\,M$

$0.0159\,M \times 2 = 0.0317\,M\,Cl^-$

aii $K = [0.0159][0.03174]^2$

$K = 1.60 \times 10^{-5}$

b. $[Pb^{2+}] = 0.0159M$

$[Cl^-] = 0.0317M$

These are the same because solution is still saturated.

c. $[Pb^{2+}]$ decreases. Since $K_{sp} = [Pb^{2+}][Cl^-]^2$ and K_{sp} is constant if $[Cl^-]\uparrow$ then $[Pb^{2+}]$ has to decrease

d. Ba^{2+}: $(2.16 \times 10^{-2})(117) = (M)(346)$ $M = 0.00730$

SO_4^{2-}: $(5.19 \times 10^{-2})(229) = (M)(346)$ $M = 0.0344$

$Q = [0.00703][0.0344]$ $Q = 2.51 \times 10^{-4}$

$Q > K$ so a precipitate will form

dii preferred

Ba^{2+}	+	SO_4^{2-}	→	$BaSO_4$
2.53 mmol		11.9 mmol		

$11.9 - 2.53 = 9.36$ mmol excess SO_4^{2-}

$\frac{9.36}{346} = 0.027$

$K_{sp} = [Ba^{2+}][SO_4^{2-}]$

$1.5 \times 10^{-9} = [Ba^{2+}][0.027]$

$[Ba^{2+}] = 5.55 \times 10^{-8}$

dii ALT. b/c 1:1 ratio $0.0344 - 0.0073 = 0.027\,M$ excess

$1.5 \times 10^{-9} = [Ba^{2+}][0.027]$

$[Ba^{2+}] = 5.55 \times 10^{-8}$

#2. Answer the following questions about the solubility of the salts Li_3PO_4 and $PbCl_2$. Assume that hydrolysis effects are negligible.
The equation for the dissolution of $Li_3PO_4(s)$ is shown below.
$Li_3PO_4(s) \rightleftharpoons 3\ Li^+(aq) + PO_4^{3-}(aq)\ K_{sp} = 3.2 \times 10^{-9}$ at 25°C
(a) Write the equilibrium-constant expression for the dissolution of $Li_3PO_4(s)$.
(b) Assuming that volume changes are negligible, calculate the maximum number of moles of $Li_3PO_4(s)$ that can dissolve in
(i) 0.50 L of water at 25°C
(ii) 0.50 L of 0.20 *M* $LiNO_3$ at 25°C
The equation for the dissolution of $PbCl_2$ is shown below.
$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\ Cl^-(aq)\ K_{sp} = 1.6 \times 10^{-5}$ at 25°C
(c) Calculate the concentration of $Cl^-(aq)$ in a saturated solution of $PbCl_2$ at 25°C .
(d) An open container holds 1.000 L of 0.00400 *M* $PbCl_2$, which is unsaturated at 25°C. Calculate the minimum volume of water, in mL, that must evaporate from the container before solid $PbCl_2$ can precipitate.

a. $K_{sp} = [Li^+]^3[PO_4^{3-}]$

bi. $3.2\times10^{-9} = [3x]^3[x]$

$x = \frac{0.00330\ mol}{1\ L} \times 0.50\ L = 0.0016\ mol$

bii $3.2\times10^{-9} = [0.20]^3[x]$

$x = \frac{4\times10^{-7}\ mol}{1\ L} \times 0.50\ L = 2.0\times10^{-7}\ mol$

c. $1.6\times10^{-5} = [x][2x]^2$

$1.6\times10^{-5} = 4x^3$

$x = 0.016\ M$ solubility $\times 2 = 0.032\ M\ Cl^-$

d. $M_1V_1 = M_2V_2$

$(0.00400)(1.0) = (0.016)(V_2)$

$V_2 = 0.250\ L$

$1.0\ L - 0.250 = 0.750\ L$ must evaporate