- Ch. 6: Solubility Equilibria & K_{sp}
- **Homework #16-6:** Problems pg. 723 #41 (in d, Hg_2^{2+} is the cation), 45b, 46b, 50, 53, 54

41 Write balanced equations and solubility product expressions for the following compounds: (a) CuBr, (b) ZnC_2O_4, (c) Ag_2CrO_4, (d) Hg_2Cl_2, (e) $AuCl_3$, (f) $Mn_3(PO_4)_2$.

a) CuBr	$K_{sp} = [Cu^+]\,[Br^-]$
b) ZnC_2O_4	$K_{sp} = [Zn^{2+}]\,[C_2O_4^{2-}]$
c) Ag_2CrO_4	$K_{sp} = [Ag^+]^2\,[CrO_4^{2-}]$
d) Hg_2Cl_2	$K_{sp} = [Hg_2^{2+}]\,[Cl^-]^2$
e) $AuCl_3$	$K_{sp} = [Au^{3+}]\,[Cl^-]^3$
f) $Mn_3(PO_4)_2$	$K_{sp} = [Mn^{2+}]^3\,[PO_4^{3-}]^2$

45 Calculate the concentrations of ions in the following saturated solution: (b) $[Al^{3+}]$ in $Al(OH)_3$ solution with $[OH^-]=2.9\times10^{-9}$ *M* (note: this is the *equilibrium* concentration of OH^-).

(b) The value of K_{sp} for aluminum hydroxide can be found in Table 2 of the text.

$$Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$$

$$K_{sp} = [Al^{3+}][OH^-]^3$$

$$\mathbf{[Al^{3+}]} = \frac{\boldsymbol{K}_{sp}}{[OH^-]^3} = \frac{1.8\times10^{-33}}{(2.9\times10^{-9})^3} = \mathbf{7.4\times10^{-8}}\ \boldsymbol{M}$$

46 From the solubility data given, calculate the solubility product for the following compound: (b) Ag_3PO_4, 6.7×10^{-3} g/L.

$$\frac{6.7\times10^{-3}\text{ g }Ag_3PO_4}{1\text{ L soln}}\times\frac{1\text{ mol }Ag_3PO_4}{418.7\text{ g }Ag_3PO_4} = 1.6\times10^{-5}\text{ mol/L} = \boldsymbol{s}$$

Substitute *s* into the equilibrium constant expression to solve for K_{sp}.

	$Ag_3PO_4(s)$	$\rightleftharpoons$	$3Ag^+(aq)$	+	$PO_4^{3-}(aq)$
Initial (*M*):			0		0
Change (*M*):	$-s$		$+3s$		$+s$
Equilibrium (*M*):			$3s$		s

$$\boldsymbol{K_{sp}} = [Ag^+]^3[PO_4^{3-}] = (3s)^3(s) = 27s^4 = 27(1.6\times10^{-5})^4 = \mathbf{1.8\times10^{-18}}$$

50 Using data from Table 2, calculate the molar solubility of CaF_2.

	$CaF_2(s)$	$\rightleftharpoons$	$Ca^{2+}(aq)$	+	$2F^-(aq)$
Initial (*M*):			0		0
Change (*M*):	$-s$		$+s$		$+2s$
Equilibrium (*M*):			s		$2s$

$$K_{sp} = [Ca^{2+}][F^-]^2 = 4.0\times10^{-11} = (s)(2s)^2 = 4s^3$$

$$\boldsymbol{s} = \text{molar solubility} = \mathbf{2.2\times10^{-4}\ mol/L}$$

53 If 20.0 mL of 0.10 *M* $Ba(NO_3)_2$ are added to 50.0 mL of 0.10 *M* Na_2CO_3, will $BaCO_3$ precipitate?

According to the solubility rules, the only precipitate that might form is $BaCO_3$:

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$$

$$[Ba^{2+}] = \frac{(20.0\text{ mL})(0.10\text{M})}{20.0\text{ mL} + 50.0\text{ mL}} = 2.9 \times 10^{-2}\ \boldsymbol{M}$$

$$[CO_3^{2-}] = \frac{(50.0\text{ mL})(0.10\text{ M})}{20.0\text{ mL} + 50.0\text{ mL}} = 7.1 \times 10^{-2}\ \boldsymbol{M}$$

$$Q = [Ba^{2+}]_0[CO_3^{2-}]_0 = (2.9 \times 10^{-2})(7.1 \times 10^{-2}) = 2.1 \times 10^{-3}$$

Since $(2.1 \times 10^{-3}) > (8.1 \times 10^{-9})$, $Q > K_{sp}$. Therefore, $BaCO_3$ will precipitate.

54 A volume of 75 mL of 0.060 *M* NaF is mixed with 25 mL of 0.15 *M* $Sr(NO_3)_2$. Calculate the concentrations in the final solution of NO_3^-, Na^+, Sr^{2+}, and F^-. (K_{sp} for $SrF_2 = 2.0\times10^{-10}$.)

The net ionic equation is:

$$Sr^{2+}(aq) + 2F^-(aq) \longrightarrow SrF_2(s)$$

First find the limiting reagent in the precipitation reaction.

$$\text{Moles } F^- = 75\text{ mL} \times \frac{0.060\text{ mol}}{1000\text{ mL soln}} = 0.0045\text{ mol}$$

$$\text{Moles } Sr^{2+} = 25\text{ mL} \times \frac{0.15\text{ mol}}{1000\text{ mL soln}} = 0.0038\text{ mol}$$

From the stoichiometry of the balanced equation, twice as many moles of F^- are required to react with Sr^{2+}. This would require 0.0076 mol of F^-, but we only have 0.0045 mol. Thus, F^- is the limiting reagent. Take the above reaction to completion, then equilibrate with K_{sp}.

	$Sr^{2+}(aq)$	+ $2F^-(aq)$	$\longrightarrow SrF_2(s)$
Initial (mol):	0.0038	0.0045	0
Change (mol):	–0.00225	–0.0045	+0.00225
Final (mol):	0.00155	0	0.00225

Now, let's establish the equilibrium reaction. The total volume of the solution is 100 mL = 0.100 L. Divide the above moles by 0.100 L to convert to molar concentration.

	$SrF_2(s)$	$\rightleftharpoons Sr^{2+}(aq)$	+ $2F^-(aq)$
Initial (*M*):	0.0225	0.0155	0
Change (*M*):	$-s$	$+s$	$+2s$
Equilibrium (*M*):	$0.0225 - s$	$0.0155 + s$	$2s$

$$K_{sp} = [Sr^{2+}][F^-]^2 = 2.0 \times 10^{-10} = (0.0155 + s)(2s)^2 \approx (0.0155)(2s)^2$$

$$s = 5.7 \times 10^{-5}\ M$$

$$\mathbf{[F^-]} = 2s = \mathbf{1.1 \times 10^{-4}}\ \boldsymbol{M}; \qquad \mathbf{[Sr^{2+}]} = 0.0155 + s = \mathbf{0.016}\ \boldsymbol{M}$$

Both sodium ions and nitrate ions are spectator ions and do not enter into the precipitation reaction.

$$\mathbf{[NO_3^-]} = \frac{2(0.0038)\,mol}{0.10\,L} = \mathbf{0.076}\ \boldsymbol{M}\ ;\ \mathbf{[Na^+]} = \frac{0.0045\text{ mol}}{0.10\text{ L}} = \mathbf{0.045}\ \boldsymbol{M}$$