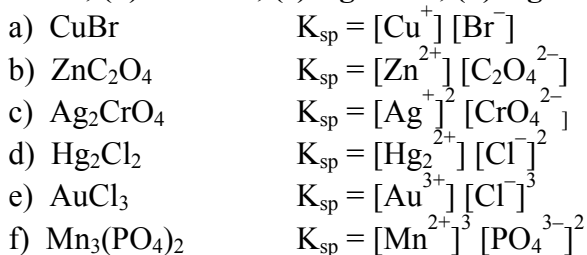


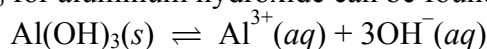
- Ch. 6: Solubility Equilibria &  $K_{sp}$
- **Homework #16-6:** Problems pg. 723 #41 (in d,  $\text{Hg}_2^{2+}$  is the cation), 45b, 46b, 50, 53, 54

41 Write balanced equations and solubility product expressions for the following compounds: (a)  $\text{CuBr}$ , (b)  $\text{ZnC}_2\text{O}_4$ , (c)  $\text{Ag}_2\text{CrO}_4$ , (d)  $\text{Hg}_2\text{Cl}_2$ , (e)  $\text{AuCl}_3$ , (f)  $\text{Mn}_3(\text{PO}_4)_2$ .



45 Calculate the concentrations of ions in the following saturated solution: (b)  $[\text{Al}^{3+}]$  in  $\text{Al}(\text{OH})_3$  solution with  $[\text{OH}^-] = 2.9 \times 10^{-9} \text{ M}$  (note: this is the *equilibrium* concentration of  $\text{OH}^-$ ).

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



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

$$[\text{Al}^{3+}] = \frac{K_{sp}}{[\text{OH}^-]^3} = \frac{1.8 \times 10^{-33}}{(2.9 \times 10^{-9})^3} = 7.4 \times 10^{-8} \text{ M}$$

46 From the solubility data given, calculate the solubility product for the following compound:

(b)  $\text{Ag}_3\text{PO}_4$ ,  $6.7 \times 10^{-3} \text{ g/L}$ .

$$\frac{6.7 \times 10^{-3} \text{ g Ag}_3\text{PO}_4}{1 \text{ L soln}} \times \frac{1 \text{ mol Ag}_3\text{PO}_4}{418.7 \text{ g Ag}_3\text{PO}_4} = 1.6 \times 10^{-5} \text{ mol/L} = s$$

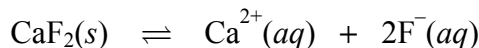
Substitute  $s$  into the equilibrium constant expression to solve for  $K_{sp}$ .



Initial (M):	0	0	
Change (M):	-s	+3s	+s
Equilibrium (M):	3s	s	

$$K_{sp} = [\text{Ag}^+]^3[\text{PO}_4^{3-}] = (3s)^3(s) = 27s^4 = 27(1.6 \times 10^{-5})^4 = 1.8 \times 10^{-18}$$

50 Using data from Table 2, calculate the molar solubility of  $\text{CaF}_2$ .



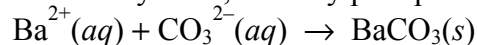
Initial (M):	0	0	
Change (M):	-s	+s	+2s
Equilibrium (M):	s	2s	

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = 4.0 \times 10^{-11} = (s)(2s)^2 = 4s^3$$

$$s = \text{molar solubility} = 2.2 \times 10^{-4} \text{ mol/L}$$

**53** If 20.0 mL of 0.10 M Ba(NO<sub>3</sub>)<sub>2</sub> are added to 50.0 mL of 0.10 M Na<sub>2</sub>CO<sub>3</sub>, will BaCO<sub>3</sub> precipitate?

According to the solubility rules, the only precipitate that might form is BaCO<sub>3</sub>:



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

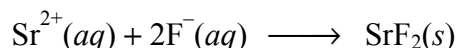
$$[\text{CO}_3^{2-}] = \frac{(50.0 \text{ mL})(0.10 \text{ M})}{20.0 \text{ mL} + 50.0 \text{ mL}} = 7.1 \times 10^{-2} \text{ M}$$

$$Q = [\text{Ba}^{2+}]_0[\text{CO}_3^{2-}]_0 = (2.9 \times 10^{-2})(7.1 \times 10^{-2}) = 2.1 \times 10^{-3}$$

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

**54** A volume of 75 mL of 0.060 M NaF is mixed with 25 mL of 0.15 M Sr(NO<sub>3</sub>)<sub>2</sub>. Calculate the concentrations in the final solution of NO<sub>3</sub><sup>-</sup>, Na<sup>+</sup>, Sr<sup>2+</sup>, and F<sup>-</sup>. ( $K_{sp}$  for SrF<sub>2</sub> =  $2.0 \times 10^{-10}$ .)

The net ionic equation is:



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<sup>-</sup> are required to react with Sr<sup>2+</sup>. This would require 0.0076 mol of F<sup>-</sup>, but we only have 0.0045 mol. Thus, F<sup>-</sup> is the limiting reagent. Take the above reaction to completion, then equilibrate with  $K_{sp}$ .

	$\text{Sr}^{2+}(aq)$	$+ 2 \text{F}^{-}(aq)$	$\longrightarrow$	$\text{SrF}_2(s)$
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.

	$\text{SrF}_2(s)$	$\rightleftharpoons$	$\text{Sr}^{2+}(aq)$	$+ 2\text{F}^{-}(aq)$
Initial (M):	0.0225		0.0155	0
Change (M):	-s		+s	+2s
Equilibrium (M):	$0.0225 - s$		$0.0155 + s$	$2s$

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

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

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

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

$$[\text{NO}_3^{-}] = \frac{2(0.0038) \text{ mol}}{0.10 \text{ L}} = 0.076 \text{ M}; \quad [\text{Na}^{+}] = \frac{0.0045 \text{ mol}}{0.10 \text{ L}} = 0.045 \text{ M}$$