

Calculating K_{sp} from Solubility

The molar solubility of CaF_2 at 35°C is $1.24 \times 10^{-3} \text{ M}$.

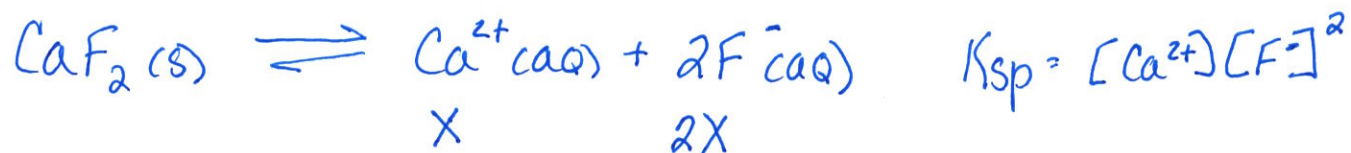
(a) What is the solubility of CaF_2 in g/L? A: 0.0968g/L

$$1.24 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{78.0748 \text{ g}}{\text{mol}} = \boxed{0.0968 \text{ g/L}}$$

Molar Mass Ca $40.078 \text{ g/mol} \times 1 = 40.078 \text{ g/mol}$

F $18.9984 \text{ g/mol} \times 2 = 37.9968 \text{ g/mol}$
 78.0748 g/mol

(b) What is K_{sp} at this temperature? A: 7.63×10^{-9}



$$[\text{Ca}^{2+}] = 1.24 \times 10^{-3} \text{ M}$$

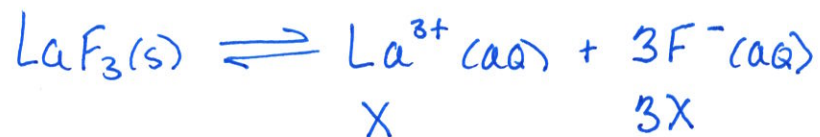
$$[\text{F}^{-}] = 2(1.24 \times 10^{-3} \text{ M}) = 0.00248 \text{ M}$$

$$K_{sp} = [1.24 \times 10^{-3}][0.00248]^2 = \boxed{7.63 \times 10^{-9}}$$

Calculating Solubility from K_{sp}

The K_{sp} for LaF_3 is 2.0×10^{-19} .

(a) What is the molar solubility of LaF_3 in water? A: $9.3 \times 10^{-6} \text{M}$



$$K_{sp} = [\text{La}^{3+}][\text{F}^{-}]^3$$

$$= [\text{X}][3\text{X}]^3 = 2.0 \times 10^{-19}$$

$$\frac{27\text{X}^4}{27} = \frac{2.0 \times 10^{-19}}{27}$$

$$\text{X}^4 = 7.4074 \times 10^{-21}$$

$$\text{X} = (7.4074 \times 10^{-21})^{1/4} = \boxed{9.3 \times 10^{-6} \text{M}}$$

(b) What is the solubility in g/L? A: $1.8 \times 10^{-3} \text{g/L}$

$$\frac{9.3 \times 10^{-6} \text{ mol}}{\text{L}} \times \frac{195.9012 \text{ g}}{\text{mol}} = 1.8 \times 10^{-3} \text{ g/L}$$

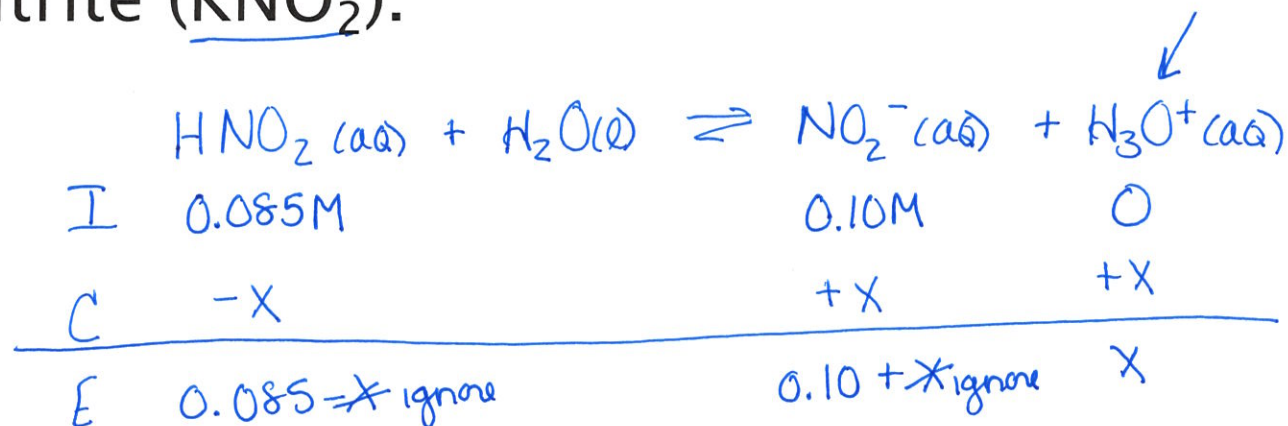
$$\text{La: } 138.906 \text{ g/mol}$$

$$\text{F: } 18.9984 \text{ g/mol} \times 3 = 56.9952 \text{ g/mol}$$

$$\frac{56.9952 \text{ g/mol}}{195.9012 \text{ g/mol}}$$

Common-Ion Calculations

1. Calculate the pH of a solution containing 0.085M nitrous acid (HNO₂; $K_a = 4.5 \times 10^{-4}$) and 0.10M potassium nitrite (KNO₂).



$$K_a = \frac{[0.10][x]}{[0.085]} = 4.5 \times 10^{-4}$$

$$\frac{[0.10][x]}{0.10} = \frac{3.825 \times 10^{-5}}{0.10}$$

$$x = 3.825 \times 10^{-4} = [\text{H}_3\text{O}^+]$$

$$\begin{aligned}
 \text{pH} &= -\log(3.825 \times 10^{-4}) \\
 &= \boxed{3.42}
 \end{aligned}$$

2. The K_{sp} of $Mn(OH)_2$ is 1.6×10^{-13} . Calculate the molar solubility of $Mn(OH)_2$ in:

a.) water A: $3.4 \times 10^{-5} M$

b.) A solution that contains 0.020M NaOH A: $4.0 \times 10^{-10} M$

c.) Compare the solubility of $Mn(OH)_2$ in these solutions

A: 85,000 times more soluble in water



$$(a) K_{sp} = [X][2X]^2 = 1.6 \times 10^{-13}$$

X

2X

$$\frac{4X^3}{4} = \frac{1.6 \times 10^{-13}}{4}$$

$$X^3 = 4.0 \times 10^{-14}$$

$$X = (4.0 \times 10^{-14})^{1/3} = \boxed{3.4 \times 10^{-5} M}$$

$$(b) K_{sp} = [X][0.02]^2 = 1.6 \times 10^{-13}$$

X

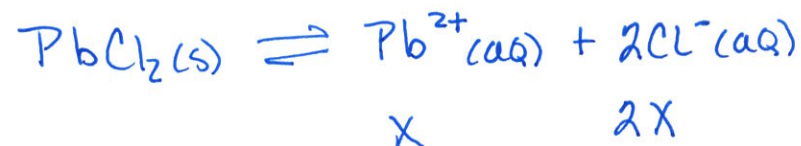
ignore
0.02M + 2X

$$\frac{[X][4.0 \times 10^{-4}]}{4.0 \times 10^{-4}} = \frac{1.6 \times 10^{-13}}{4.0 \times 10^{-4}}$$

$$\boxed{X = 4.0 \times 10^{-10} M}$$

$$(c) \frac{3.4 \times 10^{-5} M}{4.0 \times 10^{-10} M} = 85,000 \text{ times more soluble in water than } 0.020 M \text{ NaOH.}$$

3.) How much is the solubility of lead (II) chloride changed in the presence of 0.85M NaCl? $K_{sp} = 1.6 \times 10^{-5}$



Water: $K_{sp} = [x][2x]^2 = 1.6 \times 10^{-5}$

$$\frac{4x^3}{4} = \frac{1.6 \times 10^{-5}}{4}$$

$$x^3 = 4.0 \times 10^{-6}$$

$$x = (4.0 \times 10^{-6})^{1/3} = 1.59 \times 10^{-2} \text{ M}$$

In 0.85M NaCl

$$K_{sp} = [x][0.85]^2 = 1.6 \times 10^{-5}$$

$$\frac{[0.7225][x]}{[0.7225]} = \frac{1.6 \times 10^{-5}}{0.7225}$$

$$x = 2.2145 \times 10^{-5} \text{ M}$$

Compare: $\frac{1.59 \times 10^{-2} \text{ M}}{2.2145 \times 10^{-5} \text{ M}} = 718$

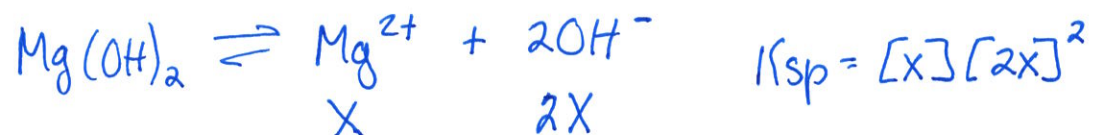
more than 700 times more soluble
in water than in 0.85M NaCl

More than 700X less
soluble than in water

Effect of pH Calculations

Calculate the solubility of $\text{Mg}(\text{OH})_2$ (one of the ingredients in the antacid Maalox) in grams per liter when buffered at pH: (a) 12.50; (b) 7.00 A: a.) $7.01 \times 10^{-7} \text{ g/L}$; b.) $7.00 \times 10^4 \text{ g/L}$

K_{sp} of $\text{Mg}(\text{OH})_2 = 1.2 \times 10^{-11}$; MM $\text{Mg}(\text{OH})_2 = 58.32 \text{ g/mol}$



(a) pH = 12.50 pOH = 14 - 12.50 = 1.5

$$[\text{OH}^-] = 10^{-1.5} = 3.16 \times 10^{-2} \text{ M}$$

$$K_{\text{sp}} = [\text{X}][3.16 \times 10^{-2} + \overset{\text{ignore}}{2\text{X}}]^2$$

$$\frac{1.2 \times 10^{-11}}{9.9856 \times 10^{-4}} = \frac{[\text{X}][9.9856 \times 10^{-4}]}{9.9856 \times 10^{-4}}$$

$$\text{X} = 1.202 \times 10^{-8} \frac{\text{mol}}{\text{L}} \times \frac{58.32 \text{ g}}{\text{mol}}$$

$$\text{X} = 7.01 \times 10^{-7} \text{ g/L}$$

(b) pH = 7 pOH = 14 - 7 = 7.00

$$[\text{OH}^-] = 10^{-7} = 1.00 \times 10^{-7} \text{ M}$$

$$K_{\text{sp}} = 1.2 \times 10^{-11} = [\text{X}][1.00 \times 10^{-7} + \overset{\text{ignore}}{2\text{X}}]^2$$

$$1.2 \times 10^{-11} = [\text{X}][1.00 \times 10^{-7}]^2$$

$$\frac{1.2 \times 10^{-11}}{1.00 \times 10^{-14}} = \frac{[\text{X}][1.00 \times 10^{-14}]}{1.00 \times 10^{-14}}$$

$$\text{X} = 1.2 \times 10^3 \frac{\text{mol}}{\text{L}} \times \frac{58.32 \text{ g}}{\text{mol}}$$

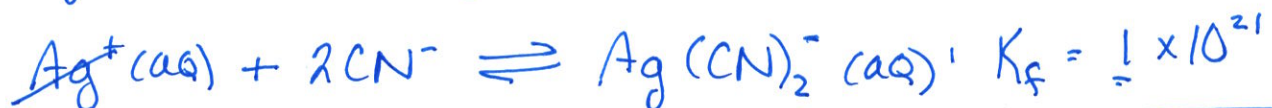
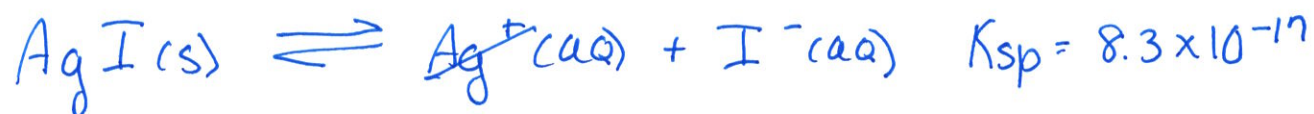
$$= \boxed{7.00 \times 10^4 \text{ g/L}}$$

Complex Ion Formation & Solubility

1. Determine the equilibrium constant for the reaction:



$$K_{\text{sp}} \text{ AgI} = 8.3 \times 10^{-17}; K_{\text{f}} \text{ Ag(CN)}_2^- = 1 \times 10^{21}$$



$$K_{\text{c}} = K_{\text{sp}} \times K_{\text{f}}$$

$$= (8.3 \times 10^{-17})(1 \times 10^{21}) = 8.3 \times 10^4$$

$$\hookrightarrow \boxed{8 \times 10^4}$$

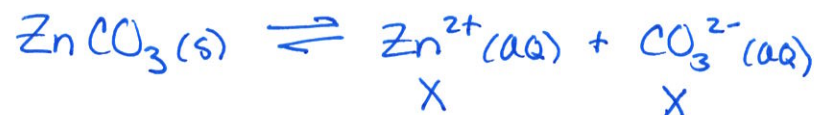
Complex Ion Formation & Solubility

2. Calculate the molar solubility of zinc carbonate at 25°C in (a) pure water and (b) 1.0M NH_3

$K_{sp} = 1.4 \times 10^{-11}$; $K_f = 4.1 \times 10^8$ for $\text{Zn}(\text{NH}_3)_4^{2+}$

A: (a) $3.7 \times 10^{-6} \text{M}$

(b) $7.6 \times 10^{-2} \text{M}$



$$(a) = K_{sp} = 1.4 \times 10^{-11} = [X][X]$$

$$X^2 = 1.4 \times 10^{-11}$$

$$X = (1.4 \times 10^{-11})^{1/2} = \boxed{3.7 \times 10^{-6} \text{M}}$$

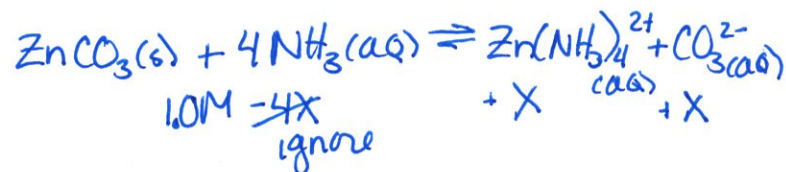
$$(b) \text{ In } 1.0 \text{M } \text{NH}_3 \quad K_f = 4.1 \times 10^8 \quad \text{Zn}(\text{NH}_3)_4^{2+}$$

$$K_c = 1.4 \times 10^{-11} \times 4.1 \times 10^8 = 5.74 \times 10^{-3}$$

$$5.74 \times 10^{-3} = \frac{[\text{Zn}(\text{NH}_3)_4^{2+}][\text{CO}_3^{2-}]}{[\text{NH}_3]^4}$$

$$5.74 \times 10^{-3} = \frac{[X][X]}{[1.0]^4} = \frac{X^2}{1} = 5.74 \times 10^{-3} \quad \boxed{X = 7.6 \times 10^{-2} \text{M}}$$

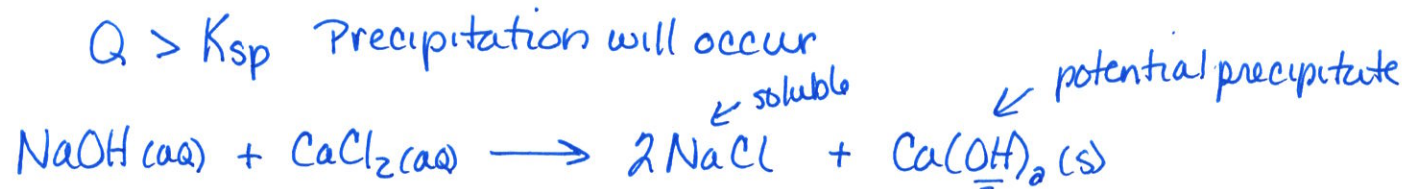
$X^2 = 5.74 \times 10^{-3}$



Prediction of Precipitation

1.) If 2.00mL of 0.200M NaOH are added to 1.00L of 0.100M CaCl₂, will a precipitate form? **A: no precipitate will form**

$Q > K_{sp}$ Precipitation will occur



$$K_{sp} = [\text{Ca}^{2+}][\text{OH}^-]^2 \quad K_{sp} = 8.0 \times 10^{-6}$$

$$[\text{Ca}^{2+}]: 0.100 \text{ M CaCl}_2 \xrightarrow{\text{0.100 mol CaCl}_2} \left(\frac{1 \text{ Ca}^{2+}}{1 \text{ CaCl}_2} \right) (1.00 \text{ L}) = \frac{0.100 \text{ mol Ca}^{2+}}{1.002 \text{ L}}$$

Total volume: 1.00L + 0.002L = 1.002L

$$= [\text{Ca}^{2+}] = 0.0998 \text{ M}$$

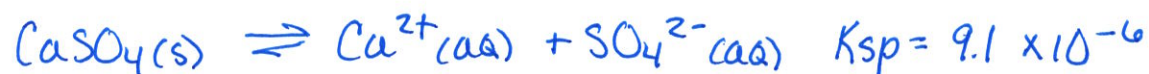
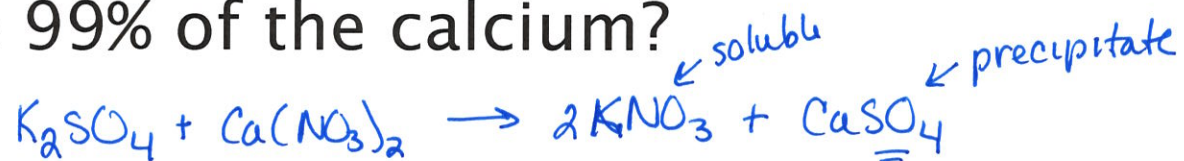
$$[\text{OH}^-]: \left(\frac{0.200 \text{ mol NaOH}}{\text{1 mol NaOH}} \right) \left(\frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} \right) (0.002 \text{ L}) = \frac{4.00 \times 10^{-4} \text{ mol OH}^-}{1.002 \text{ L}} = [\text{OH}^-] = 3.992 \times 10^{-4} \text{ M}$$

$$Q = [0.0998][3.992 \times 10^{-4}]^2 = 1.59 \times 10^{-8} = Q < K_{sp} = 8.0 \times 10^{-6}$$

no precipitation

Prediction of Precipitation

2.) How many grams of solid K_2SO_4 (174.3g/mol) would need to be added to 50.0mL of a 0.0010M $Ca(NO_3)_2$ solution in order to (a) start precipitation and (b) precipitate 99% of the calcium?



(a) Start precipitation $Q = K_{sp}$

$$K_{sp} = [Ca^{2+}][SO_4^{2-}] \quad [Ca^{2+}] = 0.0010M \quad [SO_4^{2-}] = ?$$

$$9.1 \times 10^{-6} = [0.0010M][x]$$

$$[x] = [SO_4^{2-}] = 9.1 \times 10^{-3}M$$

$$\frac{9.1 \times 10^{-3} \text{ mol } SO_4^{2-}}{L} = \frac{9.1 \times 10^{-3} \text{ mol } K_2SO_4}{L} \times 0.0500L = 4.55 \times 10^{-4} \text{ mol}$$

$$4.55 \times 10^{-4} \text{ mol} \times \frac{174.3 \text{ g}}{\text{mol}} = \boxed{7.93 \times 10^{-2} \text{ g}}$$

A: 0.079g

Prediction of Precipitation

2.) How many grams of solid K_2SO_4 (174.3g/mol) would need to be added to 50.0mL of a 0.0010M $Ca(NO_3)_2$ solution in order to (a) start precipitation and (b) precipitate 99% of the calcium?



$$0.0010M \times 0.99 = 9.9 \times 10^{-4}M \text{ removed}$$

$$\text{Remaining} = 0.0010M - 9.9 \times 10^{-4}M = 1.0 \times 10^{-5}M \text{ } Ca^{2+} \text{ left}$$

$$K_{sp} = 9.1 \times 10^{-6} = [Ca^{2+}][SO_4^{2-}]$$

$$9.1 \times 10^{-6} = [1.0 \times 10^{-5}][x]$$

$$x = [SO_4^{2-}] = 0.91M \text{ } SO_4^{2-} = 0.91M \text{ } K_2SO_4$$

$$\frac{0.91 \text{ mol}}{L} \times 0.0500L = 4.55 \times 10^{-2} \text{ mol} \left(\frac{174.3g}{\text{mol}} \right) = \boxed{7.93g}$$

Selective Precipitation

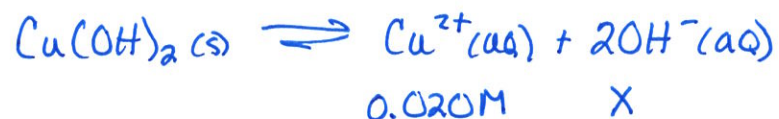
A solution contains 0.050M Mg^{2+} and 0.020M Cu^{2+} . Solid NaOH is added to the solution. $K_{\text{sp}} \text{Mg}(\text{OH})_2 = 1.8 \times 10^{-11}$; $K_{\text{sp}} \text{Cu}(\text{OH})_2 = 4.8 \times 10^{-20}$

(a) Which ion will precipitate first? **A: Cu^{2+}**

$$\begin{array}{cc} \text{Cu}(\text{OH})_2 & \text{Mg}(\text{OH})_2 \\ 4.8 \times 10^{-20} & << 1.8 \times 10^{-11} \end{array}$$

(b) What concentration of OH^- is necessary to begin precipitation of each cation. **A: Cu^{2+} needs $1.5 \times 10^{-9}\text{M}$**

Mg^{2+} needs $1.9 \times 10^{-5}\text{M}$

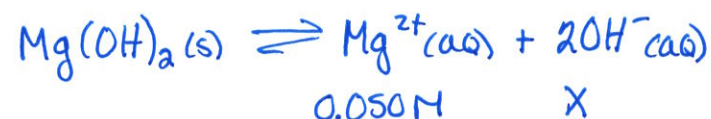


$$K_{\text{sp}} = [\text{Cu}^{2+}][\text{OH}^{-}]^2$$

$$4.8 \times 10^{-20} = [0.020][\text{X}]^2$$

$$2.4 \times 10^{-18} = \text{X}^2$$

$$\text{X} = 1.549 \times 10^{-9} \text{ M} = [\text{OH}^{-}]$$



$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^{-}]^2$$

$$1.8 \times 10^{-11} = [0.050][\text{X}]^2$$

$$3.6 \times 10^{-10} = \text{X}^2$$

$$\text{X} = 1.897 \times 10^{-5} \text{ M} = [\text{OH}^{-}]$$