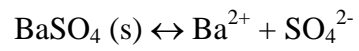


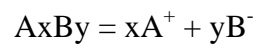
KSP



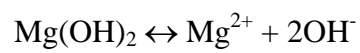
$$K = \frac{[\text{Ba}^{2+}][\text{SO}_4^{2-}]}{[\text{BaSO}_4]_{(\text{s})}}$$

$$K[\text{BaSO}_4]_{(\text{s})} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

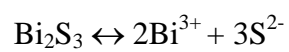
$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$



$$K_{\text{sp}} = [\text{A}^+]^x[\text{B}^-]^y$$

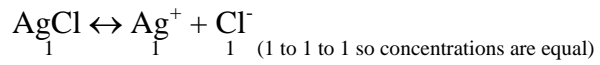


$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^-]^2$$



$$K_{\text{sp}} = [\text{Bi}^{3+}]^2[\text{S}^{2-}]^3$$

Solubility in Water



∴ Concentration of Ag^+ in terms of AgCl

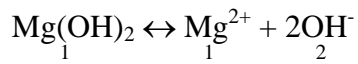
$$\text{Ag}^+ = \text{AgCl}$$

and the concentration of Cl^- in terms of AgCl

$$\text{Cl}^- = \text{AgCl}$$

$$\begin{aligned} K_{\text{sp}_{\text{AgCl}}} &= [\text{Ag}^+][\text{Cl}^-] \\ &= [\text{AgCl}][\text{AgCl}] \end{aligned}$$

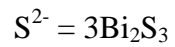
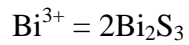
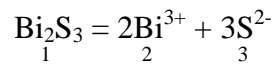
Similarly for



$$\text{Mg}^{2+} = \text{Mg}(\text{OH})_2$$

$$\text{OH}^- = 2\text{Mg}(\text{OH})_2 \quad (\text{Concentration of } \text{OH}^- \text{ 2x concentration of } \text{Mg}(\text{OH})_2)$$

$$\begin{aligned} K_{\text{sp}_{\text{Mg}(\text{OH})_2}} &= [\text{Mg}^{2+}][\text{OH}^-]^2 \\ &= [\text{Mg}(\text{OH})_2][2\text{Mg}(\text{OH})_2]^2 \end{aligned}$$



$$\begin{aligned} K_{\text{spBi}_2\text{S}_3} &= [\text{Bi}^{3+}]^2[\text{S}^{2-}]^3 \\ &= [2\text{Bi}_2\text{S}_3]^2[3\text{Bi}_2\text{S}_3]^3 \end{aligned}$$

Calculate the solubility of AgCl in (a) m/l and (b) g/100 ml given the Ksp of AgCl is

$$1.8 \times 10^{-10}.$$

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = [\text{AgCl}][\text{AgCl}]$$

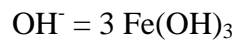
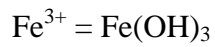
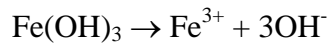
$$1.8 \times 10^{-10} = [\text{AgCl}]^2$$

$$1.3 \times 10^{-5} \text{ m/l} = [\text{AgCl}]$$

$$143 \times 1.3 \times 10^{-5} \text{ g/l}$$

Calculate the solubility of $\text{Fe}(\text{OH})_3$ in (a) m/l and (b) g/100 ml given its K_{sp} is

$$1.1 \times 10^{-36}$$



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

$$1.1 \times 10^{-36} = [\text{Fe}(\text{OH})_3][3\text{Fe}(\text{OH})_3]^3$$

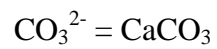
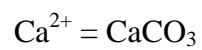
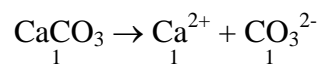
$$11,000 \times 10^{-40} = 27[\text{Fe}(\text{OH})_3]^4$$

$$407 \times 10^{-40} = [\text{Fe}(\text{OH})_3]^4$$

$$\mathbf{4.5 \times 10^{-10} = \text{Fe}(\text{OH})_3 \text{ m/l}}$$

Calculate the Ksp of CaCO₃ given its solubility is 6.9 x 10⁻³ g/l.

$$\frac{6.9 \times 10^{-3} \text{ g/l}}{100 \text{ g/m}} = 6.9 \times 10^{-5} \text{ m/l}$$



$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

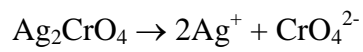
$$= [\text{CaCO}_3][\text{CaCO}_3]$$

$$= [6.9 \times 10^{-5}][6.9 \times 10^{-5}]$$

$$= \mathbf{4.8 \times 10^{-9}}$$

Calculate the K_{sp} of Ag_2CrO_4 given its solubility is 4.3×10^{-2} g/l.

$$\frac{4.3 \times 10^{-2} \text{ g/l}}{332 \text{ g/m}} = 1.3 \times 10^{-4} \text{ m/l}$$



$$Ag^+ = 2Ag_2CrO_4$$

$$CrO_4^{2-} = Ag_2CrO_4$$

$$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$$

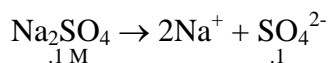
$$= [2Ag_2CrO_4]^2[Ag_2CrO_4]$$

$$= [2 \times 1.3 \times 10^{-4}]^2[1.3 \times 10^{-4}]$$

$$= \mathbf{8.8 \times 10^{-12}}$$

Calculate the maximum concentration of Ba^{2+} that can be added to a solution that is .1 M in

SO_4^{2-} given the K_{sp} of BaSO_4 is 1.1×10^{-10} .



$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

$$1.1 \times 10^{-10} = [\text{Ba}^{2+}][.1]$$

$$\frac{1.1 \times 10^{-10}}{.1} = [\text{Ba}^{2+}] = 1.1 \times 10^{-9} \text{ m/l}$$

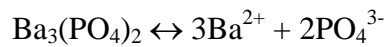
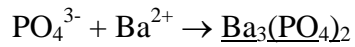
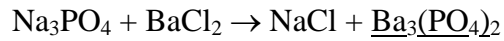
50 ml of .05 M Na_3PO_4 are added to 50 ml of .001 M BaCl_2 . (1) Show by calculation whether

$\text{Ba}_3(\text{PO}_4)_2$ will precipitate, (2) calculate the concentration of Ba^{2+} in the solution and (3)

calculate the % of Ba^{2+} that remains in the solution.

(Note: For a precipitate to occur the product of the molar concentration of each ion raised to the appropriate power must exceed K_{sp} .)

The reaction that is taking place:



Remember mmols = (M)(ml) and $M = \frac{\text{mmol}}{\text{ml}}$.

$$\text{mmols of PO}_4^{3-} = 50 \times .05 = 2.5$$

$$\text{mmols of Ba}^{2+} = 50 \times .001 = .05$$

$$M \text{ PO}_4^{3-} = \frac{2.5}{100} = 2.5 \times 10^{-2}$$

$$M \text{ Ba}^{2+} = \frac{.05}{100} = 5 \times 10^{-4}$$

$$[5 \times 10^{-4}]^3 [2.5 \times 10^{-2}]^2$$

$$[125 \times 10^{-12}] [6.25 \times 10^{-4}]$$

$780 \times 10^{-16} = 7.8 \times 10^{-14} > 10^{-29}$ so $\text{Ba}_3(\text{PO}_4)_2$ will precipitate

(The concentration of SO_4^{2-} is 50 times the concentration of the Ba^{2+} . Therefore we can assume

the concentration of the PO_4^{3-} to be a constant or changed very little when $\text{Ba}_3(\text{PO}_4)_2$

precipitates.)

$$K_{sp} = [\text{Ba}^{2+}]^3[\text{PO}_4^{3-}]^2$$

$$1.3 \times 10^{-29} = [\text{Ba}^{2+}]^3[2.5 \times 10^{-2}]^2$$

$$\frac{1.3 \times 10^{-29}}{6.25 \times 10^{-4}} = [\text{Ba}^{2+}]^3$$

$$2.08 \times 10^{-26} = [\text{Ba}^{2+}]^3$$

$$2.7 \times 10^{-9} \text{ m/l} = [\text{Ba}^{2+}] \quad (\text{concentration of } \text{Ba}^{2+} \text{ in the solution})$$

$$\% \text{ Ba}^{2+} = \frac{\text{conc. of } \text{Ba}^{2+} \text{ in solution}}{\text{initial conc. of } \text{Ba}^{2+}} \times 100$$

$$= \frac{2.7 \times 10^{-9}}{.001} \times 100$$

$$= 2.7 \times 10^{-4} \%$$

A test tube contains:

.1 M Zn^{2+} K_{sp} $Zn(OH)_2$ 4.5×10^{-17} and

.1 M Mg^{2+} K_{sp} $Mg(OH)_2$ 1.5×10^{-11}

(Note: $Mg(OH)_2$ is the most soluble.)

Calculate the maximum concentration of OH^- that can be added to a test tube without affecting the precipitation of $Mg(OH)_2$ and determine the concentration of Zn^{2+} that remains in the solution.

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

$$1.5 \times 10^{-11} = [.1][OH^-]^2$$

$$\frac{1.5 \times 10^{-11}}{.1} = [OH^-]^2$$

$$1.5 \times 10^{-11} = [OH^-]^2$$

$$1.2 \times 10^{-5} \text{ m/l} = OH^-$$

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

$$= [0.1][1.5 \times 10^{-10}] > 10^{-17} \quad Zn \text{ will precipitate while } Mg(OH)_2 \text{ will not.}$$

The concentration of Zn^{2+} in the solution is calculated in the following way:

$$4.5 \times 10^{-17} = [Zn^{2+}][1.2 \times 10^{-5}]^2$$

$$\frac{4.5 \times 10^{-17}}{1.5 \times 10^{-10}} = [Zn^{2+}]$$

$$3 \times 10^{-7} \text{ m/l} = [Zn^{2+}]$$

% of Zn^{2+} in the solution:

$$\% Zn^{2+} = \frac{3 \times 10^{-7}}{.1} \times 100$$

$$= 3 \times 10^{-4} \%$$

Calculate the pH needed to prevent the precipitation of $\text{Mg}(\text{OH})_2$ in the above problem.

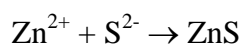
$$\text{H}^+ = \frac{1 \times 10^{-14}}{1.2 \times 10^{-3}}$$

$$\text{H}^+ = 8.2 \times 10^{-10}$$

$$\text{pH} = \mathbf{9.1}$$

(Note: If the pH of the solution is 10, both of the compounds will precipitate because increasing the pH decreases the H^+ concentration which means the OH^- concentration will increase.)

Given the K_{sp} of ZnS is 1×10^{-27} calculate the concentration of the H^+ needed to prevent the precipitation of a solution of ZnS from a solution that contains .1 M Zn^{2+} and the solution is saturated with H_2S .

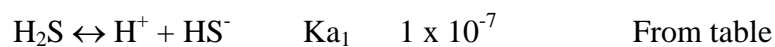


$$1 \times 10^{-27} = [\text{Zn}^{2+}][\text{S}^{2-}]$$

$$1 \times 10^{-21} = [.1][S^{2-}]$$

$$\frac{1 \times 10^{-27}}{1} = [S^{2-}]$$

$1 \times 10^{-26} = [S^{2-}]$ (Max. concentration of S^{2-} that can be added to .1 M Zn^{2+} without affecting the precipitation of ZnS [a saturated solution is formed])



$$K_{a1} = \frac{[H^+][HS^-]}{[H_2S]}$$

$$K_{a2} = \frac{[H^+][S^{2-}]}{[HS^-]}$$

$$K_{a1}K_{a2} = \frac{[H^+][HS^-]}{[H_2S]} \frac{[H^+][S^{2-}]}{[HS^-]}$$

$$K_{a1}K_{a2} = \frac{[H^+]^2[S^{2-}]}{[H_2S]}$$

$$(1 \times 10^{-7})(1.3 \times 10^{-13}) = \frac{[\text{H}^+][\text{S}^{2-}]}{.1}$$

$$1.3 \times 10^{-21} = [\text{H}^+]^2[\text{S}^{2-}] \quad (\text{This is the important relationship.})$$

$$1.3 \times 10^{-21} = [\text{H}^+]^2[1 \times 10^{-26}]$$

$$\frac{1.3 \times 10^{-21}}{1 \times 10^{-26}} = [\text{H}^+]^2$$

$$1.3 \times 10^5 = [\text{H}^+]^2$$

$$3.6 \times 10^{-2} = [\text{H}^+]$$

$$\text{pH} = -2.5$$

Given the K_{sp} of NiS is 2×10^{-27} and MnS is 4.3×10^{-22} (1) calculate the H^+ concentration and pH that will precipitate NiS without precipitating MnS and (2) calculate the % Ni^{2+} that stays in the solution that contains .05 M Mn^{2+} and .05 M Ni^{2+} .



$$4.3 \times 10^{-22} = [5 \times 10^{-2}][S^{2-}]$$

$$\frac{4.3 \times 10^{-22}}{5 \times 10^{-2}} = [S^{2-}]$$

$$8.6 \times 10^{-21} = [S^{2-}] \quad (\text{Max. concentration of } S^{2-} \text{ that can be added without precipitating. MnS}$$

the most soluble compound.)

Concentration of Ni^{2+} in the solution in which the S^{2-} is 8.6×10^{-21} :

$$2 \times 10^{-27} = [Ni^{2+}][8.6 \times 10^{-21}]$$

$$\frac{2 \times 10^{-27}}{8.6 \times 10^{-21}} = [Ni^{2+}]$$

$$2.3 \times 10^{-7} = [Ni^{2+}] \text{ m/l of } Ni^{2+} \text{ in the solution}$$

$$\frac{2.3 \times 10^{-7}}{5 \times 10^{-2}} \times 100 = 4.6 \times 10^{-4} \% \text{ of } Ni^{2+} \text{ in the solution}$$

$$1.3 \times 10^{-21} = [H^+]^2[S^{2-}]$$

$$\frac{1.3 \times 10^{-21}}{8.6 \times 10^{-21}} = [H^+]^2$$

$$.15 = [\text{H}^+]^2$$

$$.39 = [\text{H}^+]$$

$$\text{pH} = 4.1$$