

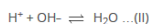
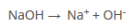
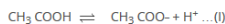
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Solved Examples on Equilibrium

Question 1:

Calculate the pH of the solution when 0.1 M CH_3COOH (50 ml) and 0.1 M NaOH (50 ml) are mixed, [$K_a(\text{CH}_3\text{COOH})=10^{-5}$]

Solution:



(I) + (II)



$$0.05 - x \quad 0.05 - x \quad x$$

Keq of eq. (III) = K_a/K_w

conc. of H_2O remain constant

$$10^9 = x/(0.05-x)^2$$

because value of eq. Const. is very high

here for $x \gg 0.05$

let $0.05-x=a$

let $0.05-x=a$

$$10^9 = 0.05/a^2$$

$$a = 7.07 \times 10^{-6}$$

$$\text{pOH} = 6 - \log 7.07$$

$$\text{pOH} = 6 - 0.85$$

$$\text{pH} = 14 - 6 + 0.85 = 8.85$$

Question 2:

Calculate the pH at the equivalence point of the titration between 0.1M CH_3COOH (25 ml) with 0.05 M NaOH , $K_a(\text{CH}_3\text{COOH}) = 1.8 \times 10^{-5}$.

Solution:

We have already seen that even though when CH_3COOH is titrated with NaOH the reaction does not go to completion but instead reaches equilibrium. We can assume that the reaction is complete and then salt gets hydrolysed because, this assumption will help us to do the problem easily and it does not effect our answer.

$$[\text{H}^+] = \sqrt{\frac{K_w K_a}{C}}$$

First of all we would calculate the concentration of the salt, CH_3COONa . For reaching equivalence point,

$$N_1 V_1 = N_2 V_2$$

$$0.1 \cdot 25 = 0.05 \cdot V_2$$

$$0.1 \cdot 25 = 0.05 \cdot V_2$$

$$\Rightarrow V_2 = 50 \text{ ml}$$

$$\text{Therefore } [\text{CH}_3\text{COONa}] = (0.1 \times 25)/75 = 0.1/3$$

$$[\text{H}^+] = \sqrt{\frac{10^{-14} \times 1.8 \times 10^{-5}}{0.1/3}}$$

$$\Rightarrow [\text{H}^+] = 2.32 \cdot 10^{-5}$$

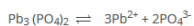
$$\Rightarrow \text{pH} = -\log 2.32 \cdot 10^{-5} = 8.63$$

Question 3:

Given the solubility product of $\text{Pb}_3(\text{PO}_4)_2$ is 1.5×10^{-32} . Determine the solubility in gms/litre.

Solution:

Solubility product of $\text{Pb}_3(\text{PO}_4)_2 = 1.5 \times 10^{-32}$



If x is the solubility of $\text{Pb}_3(\text{PO}_4)_2$

$$\text{Then } K_{sp} = (3x)^3 (2x)^2 = 108 x^5$$

$$x = 5 \sqrt[5]{\frac{K_{sp}}{108}} = \left(\frac{1.5 \times 10^{-32}}{108} \right)^{\frac{1}{5}}$$

Molecular mass of $\text{Pb}_3(\text{PO}_4)_2 = 811$

$$x = 1.692 \cdot 10^{-7} \cdot 811 \text{ g/lit} = 1.37 \times 10^{-4} \text{ g/lit}$$

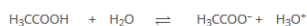
Solubility product is

$$K_{sp}(\text{SrC}_2\text{O}_4) = [\text{Sr}^{2+}][\text{C}_2\text{O}_4^{2-}] = (5.4 \times 10^{-4})^2 = 2.92 \times 10^{-7}$$

Question 4:

What is pH of 1M CH_3COOH solution? To what volume must one litre of this solution be diluted so that the pH of resulting solution will be twice the original value. Given : $K_a = 1.8 \times 10^{-5}$

Solution:



$$t = 0 \quad 1\text{M} \quad 0 \quad 0$$

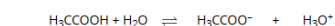
$$-x\text{M} \quad x\text{M} \quad x\text{M}$$

$$t = t_{\text{eq}} \quad (1-x)\text{M} \quad x \quad x$$

$$K_a = \frac{x^2}{1-x} \approx \frac{x^2}{1}$$

$$\therefore x = \sqrt{K_a} = 4.2 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

Now, let 1L of 1M ACOH solution be diluted to VL to double the pH and the conc. of diluted solution be C.



$$t = 0 \quad C \quad 0 \quad 0$$

$$-1.8 \times 10^{-5} \quad 1.8 \times 10^{-5} \quad 1.8 \times 10^{-5}$$

$$t = t_{\text{eq}} \quad C - 1.8 \times 10^{-5} \quad 1.8 \times 10^{-5} \quad 1.8 \times 10^{-5}$$

$$\text{New pH} = 2 \times \text{old pH} = 2 \times 2.37 = 4.74$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = 4.74$$

$$[\text{H}_3\text{O}^+] = 1.8 \times 10^{-5}$$

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.8 \times 10^{-5} = \frac{1.8 \times 10^{-5} \times 1.8 \times 10^{-5}}{C - 1.8 \times 10^{-5}}$$

$$C = 3.6 \times 10^{-5} \text{ L}$$

on dilution

$$M_1V_1 = M_2V_2$$

$$1\text{M} \times 1\text{L} = 3.6 \times 10^{-5} \text{ L} \times V_2$$

$$V_2 = 2.78 \times 10^4 \text{ L}$$

Question 5:

Find the concentration of H^+ , HCO_3^- and CO_3^{2-} , in a 0.01M solution of carbonic acid if the pH of this is 4.18.

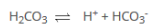
$$K_{a1}(\text{H}_2\text{CO}_3) = 4.45 \times 10^{-7} \text{ and } K_{a2} = 4.69 \times 10^{-11}$$

Solution:

$$\text{pH} = -\log[\text{H}^+]$$

$$4.18 = -\log[\text{H}^+]$$

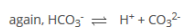
$$[\text{H}^+] = 6.61 \times 10^{-5}$$



$$K_a = \frac{[\text{H}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$\text{or, } 4.45 \times 10^{-7} = \frac{6.61 \times 10^{-5} \times [\text{HCO}_3^-]}{0.01}$$

$$\therefore [\text{HCO}_3^-] = 6.73 \times 10^{-6}$$



$$\therefore K_{a2} = \frac{[\text{H}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]} = 4.69 \times 10^{-11} = \frac{6.61 \times 10^{-5} [\text{CO}_3^{2-}]}{6.73 \times 10^{-6}}$$

$$[\text{CO}_3^{2-}] = 4.8 \times 10^{-11}$$

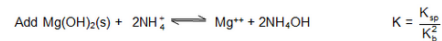
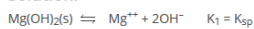
Question 6:

Calculate the molar solubility of $\text{Mg}(\text{OH})_2$ in $1\text{MNH}_4\text{Cl}$

$$K_{sp}\text{Mg}(\text{OH})_2 = 1.8 \times 10^{-11}$$

$$K_b(\text{NH}_3) = 1.8 \times 10^{-5}$$

Solution:



Let molar solubility of $\text{Mg}(\text{OH})_2$ in $1\text{MNH}_4\text{Cl}$ be x

$$\frac{1.8 \times 10^{-11}}{1.8 \times 1.8 \times 10^{-10}} = \frac{x \times (2x)^2}{1^2} = 4x^3 = \frac{1}{1.8} = \frac{1}{18}$$

$$\therefore x = .24 \text{ M}$$

Question 7:

An aqueous solution of metal bromide MBr_2 (0.05M) is saturated with H_2S . What is the minimum pH at which MS will ppt.?

$$K_{sp}(\text{MS}) = 6 \times 10^{-21}$$

Concentration of standard $\text{H}_2\text{S} = 0.1$

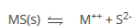
$$K_{a1}(\text{H}_2\text{S}) = 1 \times 10^{-7}$$

$$K_{a1}(\text{H}_2\text{S}) = 1 \times 10^{-7}$$

$$K_{a2}(\text{HS}^-) = 1.3 \times 10^{-13}$$

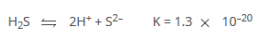
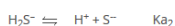
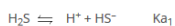
Solution:

In saturated solution of MS



$$\therefore [\text{S}^{2-}] = \frac{K_{sp}}{[\text{M}^{2+}]} = \frac{6 \times 10^{-21}}{0.05} = 1.2 \times 10^{-19}$$

The precipitate of MS will form only if $[\text{S}^{2-}]$ exceeds the concentration of 1.2×10^{-19}



$$K = \frac{[\text{H}^+]^2 [\text{S}^{2-}]}{[\text{H}_2\text{S}]}$$

$$1.3 \times 10^{-20} = \frac{[\text{H}^+]^2 \times 1.2 \times 10^{-19}}{0.1}$$

$$[\text{H}^+] = 0.109$$

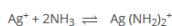
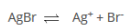
$$\text{pH} = 0.96$$

Question 8 :

How much AgBr could dissolve in 1.0 L of 0.4 M NH_3 ? Assume that $[\text{Ag}(\text{NH}_3)_2]^+$ is the only complex formed given.

$$K_f[\text{Ag}(\text{NH}_3)_2^+] = 1.0 \times 10^8, K_{sp}(\text{AgBr}) = 5.0 \times 10^{-13}$$

Solution:



Let $x =$ solubility.

$$\text{Then } x = [\text{Br}^-] = [\text{Ag}^+] + [\text{Ag}(\text{NH}_3)_2^+]$$

$$\therefore \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2} = 1 \times 10^8$$

since, $[\text{NH}_3] \gg [\text{Br}^-]$,

$$\frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+](0.4)^2} = 1.0 \times 10^8 \Rightarrow (\text{Ag}^+) = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{1.6 \times 10^7}$$

Since, the majority of silver is in the form of the complex,

$$x = [\text{Br}^-] = [\text{Ag}(\text{NH}_3)_2^+]$$

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

$$= \frac{[\text{Ag}(\text{NH}_3)_2^+][\text{Br}^-]}{1.6 \times 10^7} = \frac{x^2}{1.6 \times 10^7} = 5.0 \times 10^{-13}$$

$$x^2 = 8.0 \times 10^{-6}$$

$$\Rightarrow x = 2.8 \times 10^{-3} \text{ M}$$

Question 9:

Equal volumes of 0.02 M AgNO₃ and 0.02 M HCN were mixed. Calculate [Ag⁺] at equilibrium given, K_{sp}(AgCN) = 2.2 × 10⁻¹⁶
 K_a(HCN) = 6.2 × 10⁻¹⁰

Solution.

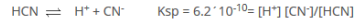
Initially, assume complete precipitation

Ag⁺ + HCN → AgCN + H⁺, since the solution

were diluted to double volume

concentration of [H⁺] = 0.02/2 = 0.01 M

Now consider the equilibrium



Since every dissolved CN⁻ is also hydrolyzes into HCN up to certain extent.

$$\text{so, } [\text{Ag}^+] = [\text{CN}^-] + [\text{HCN}]$$

$$\text{or, } \frac{2.2 \times 10^{-16}}{[\text{CN}^-]} = [\text{CN}^-] + \frac{[\text{CN}^-](0.01)}{6.2 \times 10^{-10}}$$

$$\Rightarrow [\text{CN}^-]^2 = \frac{(2.2 \times 10^{-16})(6.2 \times 10^{-10})}{0.01}$$

$$\Rightarrow [\text{CN}^-] = 3.7 \times 10^{-12}$$

$$\text{so, } [\text{Ag}^+] = \frac{K_{sp}}{[\text{CN}^-]} = \frac{2.2 \times 10^{-16}}{3.7 \times 10^{-12}} = 5.9 \times 10^{-5} \text{ M}$$

Question 10 :

What is solubility of PbS (a) ignoring the hydrolysis of ions (b) including the hydrolysis of ions (assume pH of solution = 7).

Given that:

$$K_{sp}(\text{Pb(OH)}^+) = 1.5 \times 10^{-8}$$

$$K_{a1}(\text{H}_2\text{S}) = 1 \times 10^{-7}$$

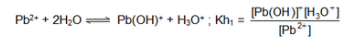
$$K_{a2}(\text{HS}^-) = 1 \times 10^{-14}, K_{a3}(\text{H}_2\text{S}) = 10^{-14}$$

Solution:

$$K_{sp} = [\text{Pb}^{2+}][\text{S}^{2-}] = S \cdot S = S^2 = 7 \cdot 10^{-29}$$

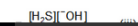
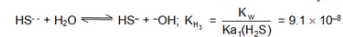
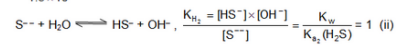
$$S = 8.4 \times 10^{-15}$$

b) Including hydrolysis: The equilibria of interest are



$$= \frac{K_w}{K_{b2}(\text{Pb(OH)}^+)} \quad \text{--- (i)}$$

$$= \frac{10^{-14}}{1.5 \times 10^{-8}} = 6.7 \times 10^{-7}$$



$$= \frac{[\text{H}_2\text{S}][\text{OH}^-]}{[\text{HS}^-]} \quad \text{--- (iii)}$$

Mass balance expression are:

$$[\text{Pb}^{2+}]_0 = [\text{Pb}^{2+}] + [\text{Pb(OH)}^+] \quad \text{--- (a)}$$

$$[\text{S}^{2-}]_0 = [\text{S}^{2-}] + [\text{HS}^-] + [\text{H}_2\text{S}] \quad \text{--- (b)}$$

Substituting the value of [Pb(OH)⁺] from equation (i) into equation (a)

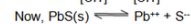
$$[\text{Pb}^{2+}]_0 = [\text{Pb}^{2+}] + \frac{K_{h1} \times [\text{Pb}^{2+}]}{[\text{H}_3\text{O}^+]}$$

$$= [\text{Pb}^{2+}] \left\{ 1 + \frac{K_{h1}}{[\text{H}_3\text{O}^+]} \right\}$$

$$\therefore [\text{Pb}^{2+}] = \frac{[\text{Pb}^{2+}]_0}{1 + \frac{K_{h1}}{[\text{H}_3\text{O}^+]}} \quad \text{--- (c)}$$

Similarly using equation (ii) and (iii) in equation (b) we have

$$[\text{S}^{2-}] = \frac{[\text{S}^{2-}]_0}{1 + \frac{K_{h2} + K_{h2} \times K_{h3}}{[\text{OH}^-]} + \frac{K_{h2} \times K_{h3}}{[\text{OH}^-]^2}} \quad \text{--- (d)}$$



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

Substituting the values of [Pb²⁺] and [S²⁻] from equations (c) and (d), we get

$$K_{sp}(\text{PbS}) = \frac{[\text{Pb}^{2+}]_0 [\text{S}^{2-}]_0}{\left(1 + \frac{K_{h1}}{[\text{H}_3\text{O}^+]}\right) \left(1 + \frac{K_{h2} + K_{h2} \times K_{h3}}{[\text{OH}^-]} + \frac{K_{h2} \times K_{h3}}{[\text{OH}^-]^2}\right)}$$

$$K_{sp}(\text{Pbs}) = \frac{[\text{Pb}^{2+}]_b [\text{S}^{2-}]_b}{\left(1 + \frac{K_{h1}}{[\text{H}_3\text{O}^+]}\right) \left(1 + \frac{K_{h2}}{[\text{OH}^-]} + \frac{K_{h2} \times K_{h3}}{[\text{OH}]^2}\right)}$$

If y be the amount of Pbs dissolve then
 $[\text{Pb}^{2+}]_b = [\text{S}^{2-}]_b = [\text{pH} = 7, \therefore [\text{H}_3\text{O}^+] = 10^{-7} = [\text{-OH}]]$

$$\therefore K_{sp}(\text{Pbs}) = \frac{y^2}{\left(1 + \frac{6.7 \times 10^{-7}}{10^{-7}}\right) \left(1 + \frac{1}{10^{-7}} + \frac{9.1 \times 10^{-3}}{10^{-7}}\right)} = 7 \times 10^{-29}$$

On solving, $y = 1.0146 \times 10^{-10}$

Question 11:

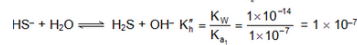
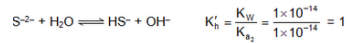
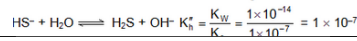
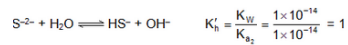
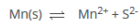
Calculate the solubility of MnS in pure water. Assume hydrolysis of S^{2-} ions.

$$K_{sp}(\text{MnS}) = 2.5 \times 10^{-10}$$

K_{a1} and K_{a2} of H_2S are 1×10^{-7} and 1×10^{-14} respectively

Solution:

Let molar solubility of MnS be XM



As $K_h \gg K'_h$, first step hydrolysis is almost complete.

$$x = [\text{Mn}^{2+}] = [\text{HS}^-] = [\text{OH}^-]$$

Consider first step hydrolysis

$$K'_h = \frac{[\text{HS}^-] \times [\text{OH}^-]}{[\text{S}^{2-}]}$$

$$\text{or, } [\text{S}^{2-}] = \frac{[\text{HS}^-] \times [\text{OH}^-]}{K'_h} = \frac{x^2}{1}$$

At equilibrium, $[\text{Mn}^{2+}] [\text{S}^{2-}] = K_{sp} = 2.5 \times 10^{-10}$

$$\text{Or } n = (2.5 \times 10^{-10})/x^2$$

$$x = 6.3 \times 10^{-4} \text{ M}$$

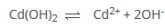
Question 12:

How much solid $\text{Na}_2\text{S}_2\text{O}_3$ should be added to 1.0 L of water so that 0.0005 mole $\text{Cd}(\text{OH})_2$ could just barely dissolve ?

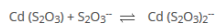
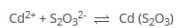
K_1 and K_2 for $\text{S}_2\text{O}_3^{2-}$ complexation with Cd^{2+} are 8.3×10^3 and 2.5×10^2 , respectively. $K_{sp}(\text{Cd}(\text{OH})_2) = 4.5 \times 10^{-15}$

Solution:

Solution:



$$K_{sp} = [\text{Cd}^{2+}] [\text{OH}^-]^2 = 4.5 \times 10^{-15}$$



$$K_1 = \frac{[\text{Cd}(\text{S}_2\text{O}_3)]}{[\text{Cd}^{2+}] [\text{S}_2\text{O}_3^{2-}]} = 8.3 \times 10^3$$

$$K_2 = \frac{[\text{Cd}(\text{S}_2\text{O}_3)_2^{2-}]}{[\text{Cd}(\text{S}_2\text{O}_3)] [\text{S}_2\text{O}_3^{2-}]}$$

$$K_2 = 2.5 \times 10^2$$

Assume that $\text{S}_2\text{O}_3^{2-}$ dose not hydrolyze

$$[\text{Cd}^{2+}] + [\text{Cd}(\text{S}_2\text{O}_3)] + [\text{Cd}(\text{S}_2\text{O}_3)_2^{2-}] = 0.00050$$

$$[\text{Cd}^{2+}] + K_1 [\text{Cd}^{2+}] \text{S}_2\text{O}_3^{2-} + K_1 K_2 [\text{Cd}^{2+}] [\text{S}_2\text{O}_3^{2-}]^2 = 0.00050$$

$$[\text{Cd}^{2+}] = K_{sp} / [\text{OH}^-]^2 = 4.5 \times 10^{-9} \text{ M}$$

$$\text{let } [\text{S}_2\text{O}_3^{2-}] = x$$

$$\text{then, } 1 + K_1 + K_1 K_2 x^2 = 1.1 \times 10^5$$

$$x = 0.2009$$

$$\text{Wt/Mwt} = 0.2009$$

$$Wt/Mwt = 0.2009$$

$$wt = 0.2009 \times 158 = 31.74$$

Question 13:

A solution contains a mixture of Ag^+ (0.1M) and Hg_2^{2+} (0.1M) which are to be separated by selective precipitation. Calculate the maximum concentration of iodide ion at which one of them gets precipitated almost completely. What percentage of that metal ion is precipitated?

$$K_{sp}(\text{AgI}) = 8.5 \times 10^{-17}$$

$$K_{sp}(\text{Hg}_2\text{I}_2) = 2.5 \times 10^{-26}$$

Solution:

Let us first calculate $[\text{I}^-]$ to precipitate AgI and Hg_2I_2

$$K_{sp}(\text{AgI}) = [\text{Ag}^+][\text{I}^-]$$

$$8.5 \times 10^{-17} = (0.1)[\text{I}^-]$$

$$[\text{I}^-] \text{ to precipitate as AgI} = (8.5 \times 10^{-17})$$

$$K_{sp}(\text{Hg}_2\text{I}_2) = [\text{Hg}_2^{2+}][\text{I}^-]^2 = 8.5 \times 10^{-16} \text{ M}$$

$$2.5 \times 10^{-26} = 0.1 [\text{I}^-]^2$$

$$[\text{I}^-] \text{ to precipitate Hg}_2\text{I}_2 = 5.0 \times 10^{-13} \text{ M}$$

$[\text{I}^-]$ to precipitate AgI is smaller. Therefore, AgI will start precipitating first. On further addition of I^- more AgI will precipitate and when $[\text{I}^-] = 5.0 \times 10^{-13} \text{ M}$, Hg_2I_2 will start precipitating. The maximum concentration of Ag^+ at this stage will thus be

$[\text{I}^-]$ to precipitate AgI is smaller. Therefore, AgI will start precipitating first. On further addition of I^- more AgI will precipitate and when $[\text{I}^-] = 5.0 \times 10^{-13} \text{ M}$, Hg_2I_2 will start precipitating. The maximum concentration of Ag^+ at this stage will thus be calculated as:

$$K_{sp}(\text{AgI}) = [\text{Ag}^+][\text{I}^-]$$

$$8.5 \times 10^{-17} = [\text{Ag}^+](5.0 \times 10^{-13})$$

$$\text{or, } [\text{Ag}^+] = 1.7 \times 10^{-4} \text{ M}$$

$$\text{Percentage of Ag}^+ \text{ remained precipitated} = [(1.7 \times 10^{-4} \text{ M})/0.1] \times 100 = 0.17\%$$

$$\text{Thus percentage of Ag}^+ \text{ precipitated} = 99.83\%$$

Question 14:

What $[\text{H}^+]$ must be maintained in a saturated H_2S (0.1M) to precipitate CdS but not ZnS , if $[\text{Cd}^{2+}] = [\text{Zn}^{2+}] = 0.1 \text{ (M)}$ initially?

Solution:

In order to prevent precipitating of ZnS ,

$$[\text{Zn}^{2+}][\text{S}^{2-}] < K_{sp}(\text{ZnS}) = 1 \times 10^{-21}$$

Ionic product

$$\text{or, } 0.1 [\text{S}^{2-}] < 1 \times 10^{-21}$$

$$\text{or, } [\text{S}^{2-}] < 1 \times 10^{-20}$$

This is the maximum value of $[\text{S}^{2-}]$ before ZnS will precipitate

Let is the maximum value of $[\text{S}^{2-}]$ be x.

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Thus for $\text{H}_2\text{S} \rightleftharpoons 2\text{H}^+ + \text{S}^{2-}$

$$K_a = \frac{[\text{H}^+][\text{S}^{2-}]}{[\text{H}_2\text{S}]} = \frac{x^2(1 \times 10^{-20})}{0.1} = 1.1 \times 10^{-21}$$

$$\text{or, } x = [\text{H}^+] = 0.1 \text{ (M)}$$

$$\text{or, } x = [\text{H}^+] = 0.1 \text{ (M)}$$

No ZnS will precipitate at and concentration of H^+ greater than 0.1M.

Question 15:

Assuming the complete dissociation of HCl and the lead salt, calculate how much HCl is added to 0.001M lead salt solution to just percent precipitation when saturated with H_2S . The concentration of H_2S in its saturated solution is 0.1M

$$K_a(\text{H}_2\text{S}) = 1.1 \times 10^{-23}$$

$$K_{sp}(\text{PbS}) = 3.4 \times 10^{-28}$$

Solution:

We know, $K_{sp}(\text{PbS}) = [\text{Pb}^{2+}][\text{S}^{2-}]$

Since lead salt is completely dissociated, $[\text{Pb}^{2+}]$ is equal to the concentration of lead salt, i.e. $[\text{Pb}^{2+}] = 0.001\text{M}$. If $[\text{S}^{2-}]$ is the concentration of S^{2-} required to just start precipitation of PbS .

$$[\text{S}^{2-}] = 3.4 \times 10^{-25}$$

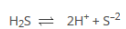
Now the addition of HCl will suppress the dissociation of H_2S to that extent that $[\text{S}^{2-}] = 3.4 \times 10^{-25} \text{ (M)}$

$$[S^{2-}] = 3.4 \times 10^{-25}$$

Now the addition of HCl will suppress the dissociation of H_2S to that extent that $[S^{2-}] = 3.4 \times 10^{-25}$ (M)

HCl is completely ionised, $[H^+] = [HCl]$

Let $[HCl]$ be x . Therefore $[H^+] = x$



At equilibrium

$$[H_2S] = 0.1 - 3.4 \times 10^{-25} \approx 0.1$$

$$[H^+] = 2 \times 3.4 \times 10^{-25} + x \approx x$$

$$[S^{2-}] = 3.4 \times 10^{-25}$$

Related Resources:-

- Click here for the Detailed [Syllabus of IIT JEE Chemistry](#)
- Look into the [Sample Papers of Previous Years](#)
- You can get the knowledge of [Useful Books of Chemistry](#)