PHYSICAL CHEMISTRY

Plus two

Unit-8

IONIC EQUILIBRIUM

Bronsted-Lowry: Implications

they are complementaries

Theory

acid = H⁺ donor

base = H⁺ acceptor

they can play both roles

Amphoteric

In a reaction we and conjugated species

Acids & Bases can not exist without the other

In a reaction they form conjugated species

FACTORS AFFECTING EQUILIBRIUM

- CONCENTRATION CHANGES
- PRESSURE CHANGES
- TEMPERATURE CHANGES
- ADDITION OF A CATALYST

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IONIC EQUILIBRIUM

1. **What are Lewis acids and bases? Give two example for each.**

   i) Lewis acid is a species that accepts an electron pair and it is a positive ion (or) an electron deficient molecule

   example :Electron deficient molecules BF_3 ,$AlCl_3$.etc

   All metal ions (or) atoms : Fe^{2+} , Cr^{3+}, Cu^{2+} and Carbonium ion

   ii) Lewis base is a species that donates an electron pair and it is a anion (or) neutral molecule with at least one lone pair of electrons.

   example :more lone pairs of electrons: NH_3, H_2O, R-O-H, R-O-R, R-NH_2

   All anions : F^- , Cl^-, CN^- and Carbanion ..

2. **Discuss the Lowry – Bronsted concept of acids and bases.**

   According to their concept,

   An Acid is defined as a substance that has a tendency to donate a proton to another substance

   Base is a substance that has a tendency to accept a proton from other substance.

   In other words, an acid is a proton donor and a base is a proton acceptor.

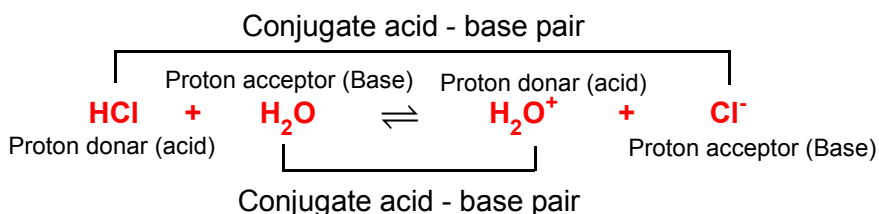   In general, Lowry – Bronsted (acid – base) reaction is represented as

   $$Acid_1 + Base_2 \rightleftharpoons Acid_2 + Base_1$$
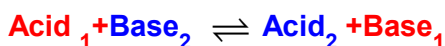
   Acid₁ $\longrightarrow$ conjugate acid   - (Proton donar)

   Base₂ $\longrightarrow$ conjugate base   - (Proton acceptor)

   Acid₂ $\longrightarrow$ conjugate acid   - (Proton donar)

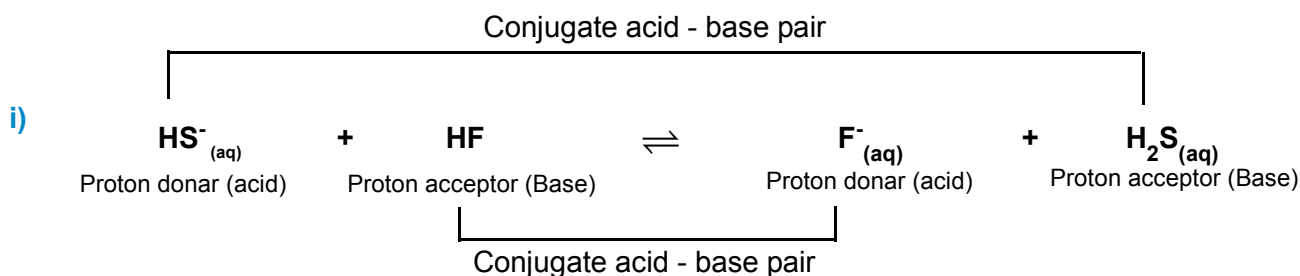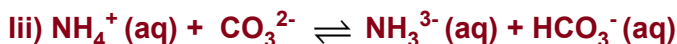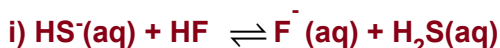   **Base₁** $\longrightarrow$ conjugate base   - (Proton acceptor)

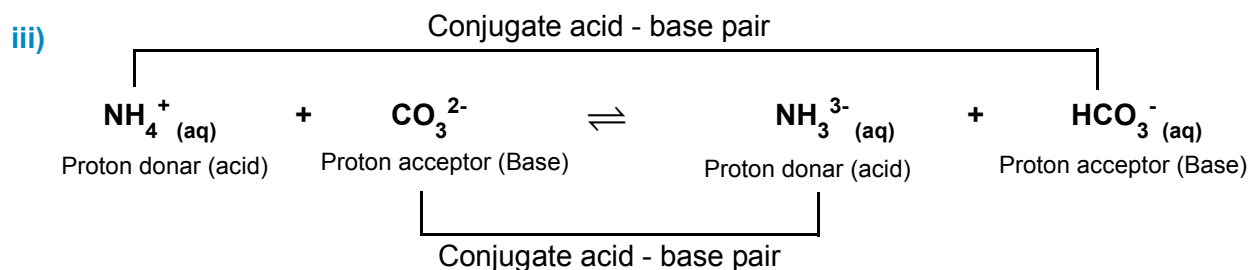   Conjugate acid - base pair

   Proton acceptor (Base) Proton donar (acid)
   $$HCl + H_2O \rightleftharpoons H_2O^+ + Cl^-$$
   Proton donar (acid)   Proton acceptor (Base)

   Conjugate acid - base pair

3. **Indentify the conjugate acid base pair for the following reaction in aqueous solution**

   i) $HS^-(aq) + HF \rightleftharpoons F^-(aq) + H_2S(aq)$

   ii) $HPO_4^-(aq) + SO_3^{2-} \rightleftharpoons PO_4^{3-}(aq) + HSO_3^-(aq)$

   Iii) $NH_4^+(aq) + CO_3^{2-} \rightleftharpoons NH_3^{3-}(aq) + HCO_3^-(aq)$

   Conjugate acid - base pair

   i)

   $HS^-_{(aq)}$ + HF $\rightleftharpoons$ $F^-_{(aq)}$ + $H_2S_{(aq)}$
   Proton donar (acid)   Proton acceptor (Base)   Proton donar (acid)   Proton acceptor (Base)

   Conjugate acid - base pair

ii)

Conjugate acid - base pair

$$HPO_4^-{}_{(aq)} + SO_3^{2-} \rightleftharpoons PO_4^{3-}{}_{(aq)} + HSO_3^-{}_{(aq)}$$

| Proton donar (acid) | Proton acceptor (Base) | Proton donar (acid) | Proton acceptor (Base) |

Conjugate acid - base pair

iii)

Conjugate acid - base pair

$$NH_4^+{}_{(aq)} + CO_3^{2-} \rightleftharpoons NH_3^{3-}{}_{(aq)} + HCO_3^-{}_{(aq)}$$

| Proton donar (acid) | Proton acceptor (Base) | Proton donar (acid) | Proton acceptor (Base) |

Conjugate acid - base pair

4. **Account for the acidic nature of $HClO_4$. In terms of Bronsted – Lowry theory, identify its conjugate base.**

$$HClO_4 + H_2O \rightleftharpoons H_3O^+ + ClO_4^-$$
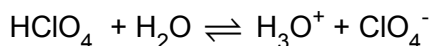
When an acid dissociates in water into H_3O^+ ions and the anion of the acid, the anion (ClO_4^-) is termed the conjugate base of that acid.

In this case, the acid is perchloric acid and its conjugate base is the perchlorate ion (ClO_4^-)

5. **When aqueous ammonia is added to $CuSO_4$ solution, the solution turns deep blue due to the formation of tetramminecopper (II) complex, $[Cu(H_2O)_4]^{2+} + 4NH_3$ (aq) $\rightleftharpoons [Cu(NH_3)_4]^{2+}$, among H_2O and NH_3 which is stronger Lewis base**

H_2O has two nonbonding electron pairs and is, therefore, a stronger Lewis base than NH_3 which only has one.

H_2O is a stronger Lewis base because the extra hydrogen in NH_3 shields the lone pair of the central nitrogen atom.

H_2O should be a stronger Lewis base because oxygen is more electronegative than nitrogen.

6. **The concentration of hydroxide ion in a water sample is found to be 2.5×10^{-6}M. Identify the nature of the solution**

$$[OH^-] = 2.5 \times 10^{-6}M$$

$$pOH = -\log[OH^-]$$

$$pOH = -\log[2.5 \times 10^{-6}]$$

$$pOH = -[\log 2.5 - 6 \log 10]$$

$$pOH = -0.3979 - (-6 \times 1)$$

$$pOH = -0.3979 + 6 \times 1$$

$$pOH = 5.6021$$

$$pOH + pH = 14 \qquad\qquad pH = 14 - 5.6021 = 8.3979$$

pH = 8.3979 pH > 7 , The nature of the solution is basic

7. **A lab assistant prepared a solution by adding a calculated quantity of HCl gas 25°C to get a solution with [H$_3$O$^+$] = 4x10^{-5}M . Is the solution neutral (or) acidic (or) basic.**

$$[H_3O^+] = 4 \times 10^{-5}M$$

$$pH = - \log[H_3O^+]$$

$$pH = - \log[4 \times 10^{-5}]$$

$$pH = - [\log 4 - 5 \log 10]$$

$$pH = - \log 4 + 5 \log 10$$

$$pH = - 0.6020 + 5 \times 1 = 4.398$$

pH = 4.398 , pH < 7 , The solution is acidic nature

8. **Calculate the pH of 0.04 M HNO$_3$ Solution.**

$$pH = -\log [H_3O^+]$$

$$pH = -\log [0.04]$$

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

$$pH = - [\log 4 - 2 \log 10]$$

$$pH = - \log 4 + 2 \log 10$$

$$pH = - 0.6020 + 2 \times 1$$

$$pH = 1.398 \simeq 1.40$$

9. **Define solubility product**

   The solubility product of a compound is defined as the product of the molar concentration of the constituent ions, each raised to the power of its stoichiometric co – efficient in a balanced equilibrium equation.

10. **Define ionic product of water. Give its value at room temperature.**

    The product of concentrations of H$^+$ and OH$^-$ ions in water at a particular temperature is known as ionic product of water.

    Therefore, the ionic product of water at 25°C is
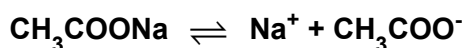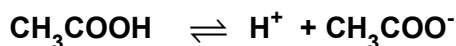
    $$K_w = [H_3O] + [OH^-]$$

    $$K_w = (1 \times 10^{-7})(1 \times 10^{-7})$$

    $$K_w = 1 \times 10^{-14} \ mol^2dm^{-6} \ at \ 298 \ K$$

11. **Explain common ion effect with an example**

    The dissociation of a weak acid is suppressed in the presence of a salt containing an ion common to the weak electrolyte. It is called the common ion effect.

    For example,

    $$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

 $$CH_3COONa \rightleftharpoons Na^+ + CH_3COO^-$$

    the addition of sodium acetate to acetic acid solution leads to the suppression in the dissociation of acetic acid which is already weakly dissociated. In this case, CH$_3$COOH and CH$_3$COONa have the **common ion,CH$_3$COO$^-$.**

12. Derive an expression for Ostwald's dilution law

Ostwald's dilution law relates the dissociation constant of the weak acid (K_a) with its degree of dissociation (α) and the concentration (c).

$$\alpha = \frac{\text{Number of moles dissociated}}{\text{total number of moles}}$$

The dissociation of acetic acid can be represented as
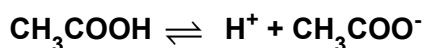
$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

	CH_3COOH	H^+	CH_3COO^-
Initial number of moles	1	-	-
Degree of dissociation	α		
Number of moles at equilibrium	1 - α	α	α
Equilibrium concentration	C (1 - α)	C α	C α

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$K_a = \frac{C\alpha \times C\alpha}{C(1-\alpha)}$$

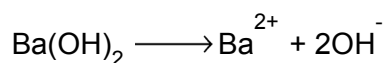$$K_a = \frac{C\alpha^2}{1-\alpha} \quad , (1-\alpha \simeq 1)$$

$$K_a = C\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}}$$

$$[H+] = C\alpha \qquad [H+] = C\sqrt{\frac{K_a}{C}} \qquad \text{for weak acid is } [H+] = \sqrt{K_a.C}$$

13. Define pH

pH of a solution is defined as the negative logarithm of base 10 of the molar concentration of the hydronium ions present in the solution.

$$pH = -\log [H_3O^+]$$

14. Calculate the pH of 1.5×10^{-3}M solution of Ba (OH)$_2$

$$Ba(OH)_2 \longrightarrow Ba^{2+} + 2OH^-$$

pOH = - log [OH⁻]

$pOH = -\log [2 \times 1.5 \times 10^{-3}]$

$pOH = -\log [3 \times 10^{-3}]$

$pOH = -[\log 3 - 3 \log 10]$

$pOH = -\log 3 + 3 \log 10$

$pOH = -\log 3 + 3 \times 1$

$pOH = -0.4771 + 3 = 2.53 \qquad \qquad \therefore$ **pH = 14 - 2.53 = 11.47**

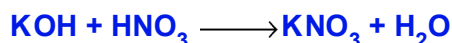**15**. **50ml of 0.05M HNO$_3$ is added to 50ml of 0.025 M KOH.Calculate the pH of the resultant solution.**

Mole = Concentration x Volume

mole of KOH = 50 ml x 0.025 = 1.25

mole of HNO$_3$ = 50 ml x 0.05M = 2.5

Total volume of solution (50+50 ml = 100 ml)

1.25 mole of KOH exactly neutralized by 1.25 mole of HNO$_3$

$$KOH + HNO_3 \longrightarrow KNO_3 + H_2O$$

that is 2.5 -1.25 = 1.25 HNO$_3$mole

Conc of H$^+$ = $\dfrac{\text{no of mole of H}^+}{\text{total volume of solution (50 + 50 ml = 100 ml)}}$

Conc of [H$^+$] = $\dfrac{1.25}{100}$ =0.0125

pH = - log [H$^+$]

pH = - log [0.0125]

pH = - log [1.25 x 10^{-2}]

pH = - [log 1.25 - 2 log10]

pH = - log 1.25 + 2 log10

pH = - 0.0969 + 2 x1

pH = 1.9031

16. **The K$_a$ value for HCN is 10^{-9} .What is the pH of 0.4M HCN solution?**

K$_a$ = 10^{-9} C= 0.4 M

[H$^+$] = $\sqrt{K_a.C}$

[H$^+$] = $\sqrt{10^{-9} \text{ x } 0.4}$

[H$^+$] = 2 x10^{-5}

pH = - log [H$^+$]

pH = - log [2 x 10^{-5}]

pH = - [log 2 - 5 log10]

pH = - log 2 + 5 log10

pH = - 0.3010 + 5 x1

pH = 4.699

17. Calculate the extent of hydrolysis and the pH of 0.1M ammonium acetate
Given that $K_a = K_b = 1.8 \times 10^{-5}$

For hydrolysis of salt of weak acid and weak base,

$$K_h = \frac{K_w}{K_a \times K_b}$$

$$K_h = \frac{\alpha^2}{(1-\alpha)^2}$$

$$K_h = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5} \times 1.8 \times 10^{-5}}$$

$$\frac{\alpha^2}{(1-\alpha)^2} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5} \times 1.8 \times 10^{-5}}$$

$$\frac{\alpha}{(1-\alpha)} = \sqrt{\frac{1 \times 10^{-14}}{1.8 \times 10^{-5} \times 1.8 \times 10^{-5}}} = \sqrt{\frac{1 \times 10^{-4}}{3.24}} = \frac{1 \times 10^{-2}}{1.8} = 0.55 \times 10^{-2}$$

hydrolysis = 0.55×10^{-2}

$$pH = \frac{1}{2}pK_w + \frac{1}{2}pK_a - \frac{1}{2}pK_b$$
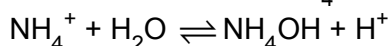
$$\frac{1}{2}pK_w = -\frac{1}{2}\log(1 \times 10^{-14}) = 7$$

$$\frac{1}{2}pK_a = -\frac{1}{2}\log K_a = -\frac{1}{2}\log(1.8 \times 10^{-5}) = \frac{1}{2}4.7447 = 2.3723$$

$$\frac{1}{2}pK_b = -\frac{1}{2}\log K_b = -\frac{1}{2}\log(1.8 \times 10^{-5}) = \frac{1}{2}4.7447 = 2.3723$$

pH = 7 + 2.3723 – 2.3723 = 7 + 0
pH = 7

18. Derive an expression for the hydrolysis constant and degree of hydrolysis of salt of strong acid and weak base.

Weak base - NH_4OH and strong acid - HCl

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

Hydrolysis constant : Applying the Law of Mass Action

$$K_h = \frac{[NH_4OH][H^+]}{[NH_4^+][H_2O]}$$

Since $[H_2O]$ is very large, it is taken to be constant and the hydrolysis constant expression is reduced to

$$K_h = \frac{[NH_4OH][H^+]}{[NH_4^+]} \quad \text{---------------- 1}$$

Relation between K_h, K_w and K_b. We know that the ionic product of water K_w is expressed as :

$$K_w = [H^+][OH^-] \quad \text{------------2}$$

For the dissociation of a weak base, NH_4OH
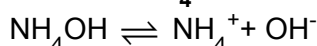
$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

The dissociation constant, K_b, can be expressed as

$$K_b = \frac{[NH_4^+]\,[OH^-]}{[NH_4OH]} \quad\text{-----------------------------} \quad 3$$

Dividing (2) by (3)

$$\frac{K_w}{K_b} = [H^+][OH^-] \times \frac{[NH_4OH]}{[NH_4^+]\,[OH^-]} = \frac{[NH_4OH][H^+]}{[NH_4^+]} = K_h$$

$$\frac{K_w}{K_b} = K_h$$

As discussed in the salt hydrolysis of strong base and weak acid. In this case also, we can establish a relationship between the K_h and K_b as

$$K_h . K_b = K_w$$

Let us calculate the K_h value in terms of degree of hydrolysis (h) and the concentration of salt

$$K_h = h^2C \quad \text{and}$$

$$[H^+] = \sqrt{K_h.C}$$

$$[H^+] = \sqrt{\frac{K_w}{K_b}.C}$$

$$pH = -\log[H^+]$$

$$pH = -\log\sqrt{\frac{K_w}{K_b}.C} \quad \text{(or)} \quad pH = -\log\left(\frac{K_w}{K_b}.C\right)^{\frac{1}{2}}$$
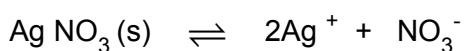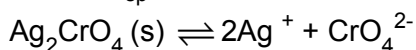
$$pH = -\frac{1}{2}\log K_w - \frac{1}{2}\log C + \frac{1}{2}\log K_b$$

$$pH = 7 + \frac{1}{2}pK_b - \frac{1}{2}\log C \quad \text{(or)} \quad pH = 7 + \frac{1}{2}pK_b = \frac{1}{2}\log C$$

In this case it is evident that pH will always be less than 7. Thus, **the solution of a salt of weak base and strong acid will always be acidic.**

19.   **Solubility product of Ag_2CrO_4 is 1×10^{-12}. What is the solubility of Ag_2CrO_4 in 0.01M $AgNO_3$ solution?**

$$K_{sp} = 1 \times 10^{-12}$$

$$Ag_2CrO_4\,(s) \rightleftharpoons 2Ag^+ + CrO_4^{2-}$$

$$Ag\,NO_3\,(s) \rightleftharpoons 2Ag^+ + NO_3^-$$

0.01 M   0.01 M   0.01 M

$$[Ag^+] = 2s + 0.01$$

$$\because 0.01 \gg 2s$$

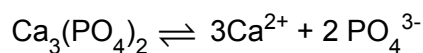$$\therefore [Ag^+] = 0.01 \text{ M}$$

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

$$K_{sp} = (s)^2(s)$$

$$1 \times 10^{-12} = (0.01)^2 (s)$$

$$[CrO_4^{2-}] = (s) = \frac{1 \times 10^{-12}}{(10^{-2})^2} = 1 \times 10^{-8}$$

20. Write the expression for the solubility product of $Ca_3(PO_4)_2$

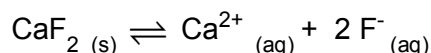$$Ca_3(PO_4)_2 \rightleftharpoons 3Ca^{2+} + 2PO_4^{3-}$$

$$K_{sp} = [Ca^{2+}]^3 . [PO_4^{3-}]^2$$

$$K_{sp} = (3s)^3.(2s)^2$$

$$K_{sp} = (27s^3).(4s^2)$$

$$K_{sp} = 108s^5 .$$

21. A saturated solution, prepared by dissolving $CaF_{2(s)}$ in water, has $[Ca^{2+}] = 3.3 \times 10^{-4}$ What is the K_{sp} of CaF_2?

$$CaF_{2\ (s)} \rightleftharpoons Ca^{2+}_{\ (aq)} + 2F^-_{\ (aq)}$$

(breaks into 2 F^- ions not F_2)

$[F^-]$ is double the $[Ca^{2+}]$

$$[2F^-] = 2[Ca^{2+}] = 2 \times 3.3 \times 10^{-4}$$

$$[2F^-] = 6.6 \times 10^{-4} M$$

$$K_{sp} = [Ca^{2+}] . [F^-]^2$$

$$K_{sp} = 3.3 \times 10^{-4} . (6.6 \times 10^{-4})^2$$

$$K_{sp} = 1.44 \times 10^{-10} .$$

22. K_{sp} of AgCl is 1.8×10^{-10}. Calculate molar solubility in 1 M $AgNO_3$.

$$AgCl_{(s)} \rightleftharpoons Ag^+_{\ (aq)} + Cl^-_{\ (aq)}$$
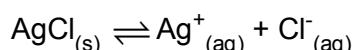
x = solubility of AgCl in 1M $AgNO_3$.

$$AgNO_{3\ (aq)} \rightleftharpoons Ag^+_{\ (aq)} + NO^-_{3\ (aq)}$$
$$\quad\quad\quad\quad\quad\quad 1M \quad\quad\quad 1M$$
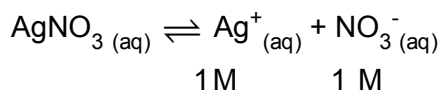
$[Ag^+] = x + 1\ M \simeq 1M \quad (\because x << 1)$

$[Cl^-] = x$

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

$1.8 \times 10^{-10} = (1)(x)$

$x = 1.8 \times 10^{-10} M$

23. A particular saturated solution of silver chromate Ag_2CrO_4 has $[Ag^+] = 5 \times 10^{-5}$ and $[CrO_4^{2-}] = 4.4 \times 10^{-4}$ What is the value of K_{sp} for Ag_2CrO_4?

$$Ag_2CrO_4\ (s) \rightleftharpoons 2Ag^+ + CrO_4^{2-}$$
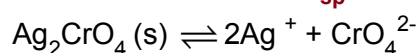
$$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$$

$$K_{sp} = (5 \times 10^{-5})^2 (4.4 \times 10^{-4})$$
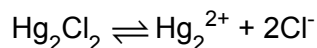
$$K_{sp} = (25 \times 10^{-10})(4.4 \times 10^{-4})$$

$$K_{sp} = (25 \times 4.4) \times 10^{-10} \times 10^{-4}$$

$$K_{sp} = 110 \times 10^{-14}$$

$$\mathbf{K_{sp} = 1.10 \times 10^{-12}}$$

24. **Write the expression for the solubility product of Hg_2Cl_2.**

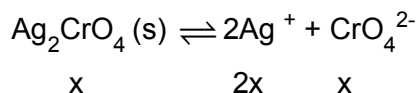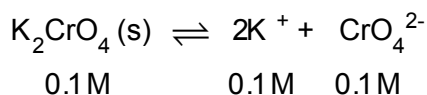$$Hg_2Cl_2 \rightleftharpoons Hg_2^{2+} + 2Cl^-$$

$$K_{sp} = [Hg_2^{2+}][Cl^-]^2$$

$$K_{sp} = [s].[2s]^2$$

$$K_{sp} = 4s^3$$

25. **K_{sp} of Ag_2CrO_4 is 1.1×10^{-12}. What is solubility of Ag_2CrO_4 in $0.1M\ K_2CrO_4$.**

$$Ag_2CrO_4\ (s) \rightleftharpoons 2Ag^+ + CrO_4^{2-}$$
$$\quad\quad x \quad\quad\quad 2x \quad\quad x$$

x is the solubility of Ag_2CrO_4 in $0.1M\ K_2CrO_4$.

$$K_2CrO_4\ (s) \rightleftharpoons 2K^+ + CrO_4^{2-}$$
$$\quad 0.1\,M \quad\quad 0.1\,M \quad 0.1\,M$$

$$[Ag^+] = 2x$$

$$[CrO_4^{2-}] = (x + 0.1) \simeq 0.1 \quad (\because x << 0.1)$$

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

$$1.1 \times 10^{-12} = (2x)^2 (0.1)$$

$$1.1 \times 10^{-12} = 0.4x^2$$

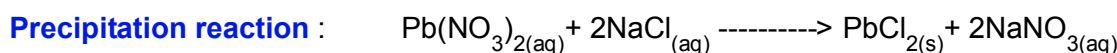$$x^2 = \frac{1.1 \times 10^{-12}}{0.4} \qquad\qquad x = \sqrt{\frac{1.1 \times 10^{-12}}{0.4}}$$

$$x = \sqrt{2.75 \times 10^{-12}} \qquad\qquad \mathbf{x = 1.658 \times 10^{-6} M}$$

26. **Will a precipitate be formed when 0.150 L of 0.1M $Pb(NO_3)_2$ and 0.100L of 0.2 M NaCl are mixed? $K_{sp}(PbCl_2) = 1.2 \times 10^{-5}$.**

Precipitation reaction : $\quad Pb(NO_3)_{2(aq)} + 2NaCl_{(aq)} \text{----------}> PbCl_{2(s)} + 2NaNO_{3(aq)}$

Solubility reaction : $\quad PbCl_2(s) \text{-----}> Pb^{2+} + 2Cl^-$

K_{sp} is given and in order to find out if a precipitate form we need to find out Q **Q = $[Pb^{2+}][Cl^-]^2$**

1:1 mole ratios so initial $[Pb^{2+}] = [Pb(NO_3)_2] = 0.1\ M$; initial $[Cl^-] = [NaCl] = 0.2\ M$

After we combine the solution the total volume becomes .250 L ,

so the concentration of Pb^{2+} and Cl^- ions is :

Number of moles Pb^{2+}= molarity x Volume of the solution in lit= 0.1 × 0.150 L

$$[Pb^{2+}] = \frac{0.1\ M\ Pb^{2+}\ x\ 0.150L}{0.250\ L} = 0.06\ M\ Pb^{2+}$$

$$[Cl^-] = \frac{0.2\ M\ Cl^-\ x\ 0.100L}{0.250\ L} = 0.08\ M\ Cl^-$$

$Q = [Pb^{2+}].[Cl^-]^2 \qquad Q = (0.060)(0.080)^2 = 3.8 \times 10^{-4}$

$Q > K_{sp}$, [$3.8 \times 10^{-4} < 1.2 \times 10^{-5}$] so $PbCl_2$ does precipitate.Eq shifts to the left .
   S.SHANMUGAM ,St.John's M.H.S.S porur Chennai -116 Mob: 9841945665

27. K_{sp} of $Al(OH)_3$ is 1×10^{-15}M. At what pH does 1.0×10^{-3}M Al^{3+} precipitate on the addition of buffer of NH_4Cl and NH_4OH solution?

1) Write the dissociation equation: $Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-$

2) Write the Ksp expression: $K_{sp} = [Al^{3+}][OH^-]^3$

3. Plug into the K_{sp} expression: 1×10^{-15}M $= (1.0 \times 10^{-3}$M$) [OH^-]^3$

4) Solve for s, which is the $[OH^-]$: $[OH^-]^3 = \dfrac{K_{sp}}{[Al^{3+}]}$

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

$$[OH^-] = 3\sqrt{\dfrac{1 \times 10^{-15}}{1 \times 10^{-3}}}$$

$$[OH^-] = 3\sqrt{1 \times 10^{-12}}$$

$$[OH^-] = \left(1 \times 10^{-12}\right)^{\frac{1}{3}}$$

$[OH^-] = 1 \times 10^{-4}$

5) Calculate the pH: $pOH = -\log[OH^+]$

$pOH = -\log(1 \times 10^{-4}) = 4$

pH = 14 - pOH = 14 - 4 = 10

27. K_{sp} of $Al(OH)_3$ is 1×10^{-15}M. At what pH does 1.0×10^{-3}M Al^{3+} precipitate on the addition of buffer of NH_4Cl and NH_4OH solution?

1) Write the dissociation equation: $Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-$

2) Write the Ksp expression: $K_{sp} = [Al^{3+}][OH^-]^3$

3. Plug into the K_{sp} expression: 1×10^{-15}M $= (1.0 \times 10^{-3}$M$) [OH^-]^3$