

# B.Sc. Semester – I

## Subject: - CHE - 101: Ionic Equilibrium (Numerical)

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### ➤ Formula:

$$\checkmark [\text{H}_3\text{O}^+] = \sqrt{K_a C_0}$$

$$\checkmark [\text{OH}^-] = \sqrt{K_b C_0}$$

$$\checkmark K_a \times K_b = K_w$$

$$\checkmark \text{Degree of Dissociation } (\alpha) = \frac{\text{Number of molecules present as ions}}{\text{Total number of molecules dissolved}}$$

$$\checkmark \alpha = \sqrt{\frac{K_a}{c}}$$

$$\checkmark \text{pH} = \frac{1}{2} [\text{p}K_a - \log_{10} c]$$

$$\checkmark \alpha = \sqrt{\frac{K_b}{c}}$$

$$\checkmark \text{pOH} = \frac{1}{2} [\text{p}K_b - \log_{10} c]$$

$$\checkmark C = \frac{1}{V}$$

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

$$\checkmark \text{pOH} = -\log [\text{OH}^-]$$

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

$$\checkmark \text{pH} + \text{pOH} = \text{p}K_w \text{ or } 14$$

$$\checkmark \text{Degree of Hydrolysis } (h) = \frac{\text{Number of moles of salt undergoing hydrolysis}}{\text{Total number of moles of the salt dissolved}}$$

$$\checkmark K_h = \frac{K_w}{K_a}$$

$$\checkmark h = \sqrt{\frac{K_w}{K_a \cdot c}} \text{ or } h = \sqrt{\frac{K_h}{c}}$$

$$\checkmark \text{pH} = \frac{1}{2} [\text{p}K_w + \text{p}K_a + \log_{10} c]$$

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

$$\checkmark h = \sqrt{\frac{K_w}{K_b \cdot c}} \text{ or } h = \sqrt{\frac{K_h}{c}}$$

$$\checkmark \text{pH} = \frac{1}{2} [\text{p}K_w - \text{p}K_b - \log_{10} c]$$

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

$$\checkmark h = \sqrt{\frac{K_w}{K_a \cdot K_b}} \text{ or } h = \sqrt{K_h}$$

$$\checkmark \text{pH} = \frac{1}{2} [\text{p}K_w + \text{p}K_a - \text{p}K_b]$$

$$\checkmark \text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$\checkmark \text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$\checkmark \text{pH} = \text{p}K_w - \text{p}K_b - \log_{10} \frac{[\text{Salt}]}{[\text{Base}]}$$

$$\checkmark \beta = \frac{d\text{B}}{d(\text{pH})}$$

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**1. The concentration of hydrogen ion in a sample of soft drink is  $3.8 \times 10^{-2}$  M. What is its pH?**

$$[\text{H}^+] = 3.8 \times 10^{-2}, \quad \text{pH} = ?$$

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

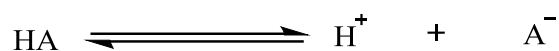
$$= -\log [3.8 \times 10^{-2}]$$

$$= -\log [3.8] - \log [10^{-2}]$$

$$= -0.58 + 2.0$$

$$\text{pH} = 1.42$$

2. The pH of 0.1M monobasic acid is 4.50. Calculate the concentration of species H<sup>+</sup>, A<sup>-</sup> and HA at equilibrium. Also, determine the value of K<sub>a</sub> and pK<sub>a</sub> of the monobasic acid.



$$\text{pH} = 4.5, [\text{H}^+] = ?, [\text{A}^-] = ?, [\text{HA}] = ?, K_a = ?, \text{p}K_a = ?$$

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

Therefore,

$$\begin{aligned} [\text{H}^+] &= \text{antilog} [-\text{pH}] \\ &= \text{antilog} [-4.50] \\ &= \text{antilog} [\bar{5}.50] \end{aligned}$$

$$[\text{H}^+] = 3.16 \times 10^{-5} \quad \text{at equilibrium } [\text{H}^+] = [\text{A}^-], \text{ So } [\text{A}^-] = 3.16 \times 10^{-5}$$

$$\text{at equilibrium} \quad [\text{HA}] = 0.1 - (3.16 \times 10^{-5})$$

$$[\text{HA}] \cong 0.1$$

$$\text{Thus } K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

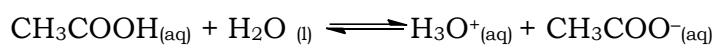
$$K_a = \frac{(3.16 \times 10^{-5})^2}{0.1}$$

$$K_a = 9.98 \times 10^{-9}$$

$$\begin{aligned} \text{p}K_a &= -\log [K_a] \\ &= -\log [9.98 \times 10^{-9}] \\ &= -\log [9.98] - \log [10^{-9}] \\ &= -0.99 + 9.0 \end{aligned}$$

$$\text{p}K_a = 8.01$$

3. Calculate the pH of 0.08 M solution of Acetic acid. The ionization constant (K<sub>a</sub>) of the acid is 2.5 × 10<sup>-5</sup>.



$$C_0 = 0.08\text{M}, K_a = 2.5 \times 10^{-5}, \text{pH} = ?$$

$$\begin{aligned} \text{p}K_a &= -\log [K_a] \\ &= -\log [2.5 \times 10^{-5}] \\ &= -\log [2.5] - \log [10^{-5}] \\ &= -0.39 + 5.0 \end{aligned}$$

$$\text{pH} = \frac{1}{2} [\text{p}K_a - \log_{10} c]$$

$$\text{pH} = \frac{1}{2} [4.61 - \log_{10} 0.08]$$

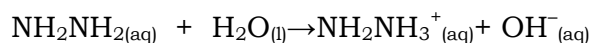
$$\text{pH} = \frac{1}{2} [4.61 - 1.09]$$

$$\text{pH} = \frac{1}{2} [3.52]$$

$$\text{pH} = 1.76$$

$$\text{p}K_a = 4.61$$

4. The pH of 0.004 M hydrazine solution is 9.7. Calculate its ionization constant K<sub>b</sub> and pK<sub>b</sub>.



$$C_0 = 0.004\text{M}, \text{pH} = 9.7, K_b = ?, \text{p}K_b = ?$$

$$\text{pH} + \text{pOH} = 14 \quad \therefore \text{pOH} = 14 - 9.7 \quad \therefore \text{pOH} = 4.3$$

$$\text{pOH} = -\log [\text{OH}^-]$$

Therefore,

$$\begin{aligned} [\text{OH}^-] &= \text{antilog} [-\text{pOH}] \\ &= \text{antilog} [-4.3] \\ &= \text{antilog} [\bar{5}.70] \end{aligned}$$

$$[\text{OH}^-] = 5.01 \times 10^{-5}$$

The dissociation is very small, so the concentration of the undissociated hydrazine to be taken equal to 0.004M and  $[\text{OH}^-] = [\text{NH}_2\text{NH}_3^+]$ .

$$\text{Now, } [\text{OH}^-] = \sqrt{K_b C_0}$$

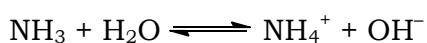
$$\therefore K_b = \frac{[\text{OH}^-]^2}{C_0} \quad \therefore K_b = \frac{[5.01 \times 10^{-5}]^2}{4.0 \times 10^{-3}}$$

$$\therefore K_b = 6.275 \times 10^{-7}$$

$$\begin{aligned} \text{p}K_b &= -\log [K_b] \\ &= -\log [6.275 \times 10^{-7}] \\ &= -\log [6.275] - \log [10^{-7}] \\ &= -0.797 + 7.0 \end{aligned}$$

$$\text{p}K_a = 6.203$$

**5. Determine the degree of ionization and pH of a 0.05M of ammonia solution. The ionization constant of ammonia is  $1.77 \times 10^{-5}$ . Also, calculate the ionization constant of the conjugate acid of ammonia.**



$$C = 0.05\text{M}, K_b = 1.77 \times 10^{-5}, \text{pH} = ?, \alpha = ?$$

$$\alpha = \sqrt{\frac{K_b}{c}} \quad \alpha = \sqrt{\frac{1.77 \times 10^{-5}}{5.0 \times 10^{-2}}} \quad \alpha = \sqrt{3.54 \times 10^{-4}} \quad \alpha = 1.88 \times 10^{-2}$$

$$\text{pOH} = \frac{1}{2} [\text{p}K_b - \log_{10} c]$$

$$\text{pOH} = \frac{1}{2} [4.76 - \log 0.05]$$

$$\text{pOH} = \frac{1}{2} [4.76 - (-1.30)]$$

$$\text{pOH} = 3.03$$

$$\text{Now pH} + \text{pOH} = 14 \therefore \text{pH} = 14 - 3.03$$

$$\therefore \text{pH} = 10.97$$

$$\text{p}K_b = -\log [K_b]$$

$$= -\log [1.77 \times 10^{-5}]$$

$$= -\log [1.77] - \log [10^{-5}]$$

$$= -0.24 + 5.0$$

$$\text{p}K_b = 4.76$$

Now, using the relation for conjugate acid-base pair,

$$K_a \times K_b = K_w \quad K_a = K_w / K_b$$

$$= 1.0 \times 10^{-14} / 1.77 \times 10^{-5}$$

$$K_a = 5.64 \times 10^{-10}$$

**6. The  $\text{p}K_a$  of acetic acid and  $\text{p}K_b$  of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.**

$$\text{p}K_a = 4.76, \text{p}K_b = 4.75, \text{pH} = ?$$

$$\text{pH} = \frac{1}{2} [\text{p}K_w + \text{p}K_a - \text{p}K_b]$$

$$= \frac{1}{2} [14 + 4.76 - 4.75]$$

$$= \frac{1}{2} [14.01]$$

$$= 7.005$$

**7. How many grams of  $\text{NH}_4\text{Cl}$  added to prepare 1.2 liter 0.25 M  $\text{NH}_3$  buffer solution having pH 9.25.  $\text{p}K_b = 4.77$ .**

**Buffer solution volume** = 1.2 lit, **Base con.** = 0.25 M buffer pH = 9.25,  **$\text{NH}_4\text{Cl}$  grams** = ?

$$\text{pOH} = 14 - \text{pH} = 14 - 9.25 = 4.75$$

$$\therefore \text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

$$\therefore 4.75 = 4.77 + \log \frac{[\text{salt}]}{0.25}$$

$$\therefore -0.02 = \log \frac{[\text{salt}]}{0.25}$$

$$\therefore -0.02 = \log [\text{salt}] - \log 0.25$$

$$\therefore -0.02 = \log [\text{salt}] - (-0.6020)$$

$$\therefore -0.622 = \log [\text{salt}]$$

$$\therefore [\text{salt}] = \text{antilog}[-0.622]$$

$$\therefore [\text{salt}] = \text{antilog}[\bar{1}.378]$$

$$\therefore [\text{salt}] = \mathbf{0.238 \text{ M}}$$

$$\text{Weight of NH}_4\text{Cl} = \text{Molarity} \times \text{M.W.} = 0.239 \times 53.5 = 12.79 \text{ gm/liter}$$

$$\therefore 12.79 \times 1.2 = \mathbf{15.34 \text{ gm in 1.2 liter}}$$

**8. How many grams of Sodium acetate added to prepare 2.5 liter 0.1 M Acetic acid buffer solution having pH 4.75.  $\text{pK}_a = 4.74$ .**

$$\therefore \text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$\therefore 4.75 = 4.74 + \log \frac{[\text{salt}]}{[0.1]}$$

$$\therefore 0.01 = \log \frac{[\text{salt}]}{[0.1]}$$

$$\therefore 0.01 = \log [\text{salt}] - \log 0.1$$

$$\therefore 0.01 = \log [\text{salt}] - (-1.0)$$

$$\therefore -0.99 = \log [\text{salt}]$$

$$\therefore [\text{salt}] = \text{antilog}[-0.991]$$

$$\therefore [\text{salt}] = \text{antilog}[\bar{1}.01]$$

$$\therefore [\text{salt}] = \mathbf{0.102 \text{ M}}$$

$$\text{Weight of Sodium acetate (CH}_3\text{COONa)} = \text{Molarity} \times \text{M.W.} = 0.102 \times 82$$

$$= 8.364 \text{ gm/liter} \quad \therefore 8.364 \times 2.5 = \mathbf{20.91 \text{ gm in 2.5 liter}}$$

**9. Calculate the pH for 0.15 M CH<sub>3</sub>COOH and 0.2 M CH<sub>3</sub>COONa for 1 lit. buffer solution.  $\text{K}_a = 1.74 \times 10^{-5}$ .**

$$[\text{acid}] = 0.15 \text{ M}, [\text{salt}] = 0.2 \text{ M}, \text{K}_a = 1.74 \times 10^{-5}, \text{pH} = ?$$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$= 4.76 + \log \frac{0.2}{0.15}$$

$$= 4.76 + \log 1.33$$

$$= 4.76 + 0.124$$

$$\mathbf{\text{pH} = 4.884}$$

$$\text{pK}_a = -\log [\text{K}_a]$$

$$= -\log [1.74 \times 10^{-5}]$$

$$= -\log [1.74] - \log [10^{-5}]$$

$$= -0.24 + 5.0$$

$$\mathbf{\text{pK}_a = 4.76}$$

**10. Calculate the pH for 0.25 M CH<sub>3</sub>COOH solution.  $\text{K}_a = 1.74 \times 10^{-5}$ .**

$$C = 0.25 \text{ M}, \text{K}_a = 1.74 \times 10^{-5}$$

$$[\text{H}_3\text{O}^+] = \sqrt{\text{K}_a \cdot c} = (1.74 \times 10^{-5} \times 0.25)^{1/2} = (0.435 \times 10^{-5})^{1/2} = (4.35 \times 10^{-6})^{1/2}$$

$$[\text{H}_3\text{O}^+] = 2.087 \times 10^{-3}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log [2.087 \times 10^{-3}] \\ &= -\log [2.087] - \log [10^{-3}] \\ &= -0.44 + 3.0 \end{aligned}$$

$$\text{pH} = 2.56$$

11. Find out the Hydrolysis constant and pH for 0.5 M  $\text{CH}_3\text{COONa}$  solution.  $K_a = 1.74 \times 10^{-5}$ .

$\text{CH}_3\text{COONa}$  is a salt of weak acid ( $\text{CH}_3\text{COOH}$ ) and strong base ( $\text{NaOH}$ ).

$$C = 0.5 \text{ M}, K_a = 1.74 \times 10^{-5}, \text{pH} = ?, K_h = ?$$

$$\begin{aligned} K_h &= \frac{K_w}{K_a}, \quad \therefore K_h = \frac{1 \times 10^{-14}}{1.74 \times 10^{-5}} & \therefore K_h &= 5.74 \times 10^{-10} \\ \text{pH} &= \frac{1}{2} [\text{p}K_w + \text{p}K_a + \log_{10} c] & \text{p}K_a &= -\log [K_a] \\ & & &= -\log [1.74 \times 10^{-5}] \\ & & &= -\log [1.74] - \log [10^{-5}] \\ & & &= -0.24 + 5.0 \\ \text{pH} &= \frac{1}{2} [14 + 4.76 + \log 0.5] \\ \text{pH} &= \frac{1}{2} [14 + 4.76 - 0.30] & \text{p}K_a &= 4.76 \\ \text{pH} &= \frac{1}{2} [19.06] \\ \text{pH} &= 9.53 \end{aligned}$$

12. Find out the Degree of Hydrolysis and pH for 0.25 M  $\text{NH}_4\text{Cl}$  solution.  $K_b = 1.77 \times 10^{-5}$ .

$\text{NH}_4\text{Cl}$  is a salt of weak base ( $\text{NH}_4\text{OH}$ ) and strong acid ( $\text{HCl}$ ).

$$C = 0.25 \text{ M}, K_b = 1.77 \times 10^{-5}, \text{pH} = ?, K_h = ?, h = ?$$

$$\begin{aligned} K_h &= \frac{K_w}{K_b}, \quad \therefore K_h = \frac{1 \times 10^{-14}}{1.77 \times 10^{-5}} & \therefore K_h &= 5.64 \times 10^{-10} \\ h &= \sqrt{\frac{K_h}{c}}, \quad \therefore h = \sqrt{\frac{5.64 \times 10^{-10}}{2.5 \times 10^{-1}}} & \therefore h &= \sqrt{22.56 \times 10^{-10}} & \therefore h &= 4.47 \times 10^{-5} \\ \text{pH} &= \frac{1}{2} [\text{p}K_w - \text{p}K_b - \log_{10} c] & \text{p}K_b &= -\log [K_b] \\ & & &= -\log [1.77 \times 10^{-5}] \\ & & &= -\log [1.77] - \log [10^{-5}] \\ & & &= -0.25 + 5.0 \\ \text{pH} &= \frac{1}{2} [14 - 4.75 - \log 0.25] \\ \text{pH} &= \frac{1}{2} [14 - 4.76 + 0.60] & \text{p}K_b &= 4.75 \\ \text{pH} &= \frac{1}{2} [9.84] \\ \text{pH} &= 4.92 \end{aligned}$$

13. Calculate the Degree of Hydrolysis, Hydrolysis constant and pH for  $\text{CH}_3\text{COONH}_4$  solution.  $K_a = 1.74 \times 10^{-5}$  and  $K_b = 1.77 \times 10^{-5}$ .

$\text{CH}_3\text{COONH}_4$  is a salt of weak base ( $\text{NH}_4\text{OH}$ ) and weak acid ( $\text{CH}_3\text{COOH}$ ).

$$K_a = 1.74 \times 10^{-5}, K_b = 1.77 \times 10^{-5}, \text{pH} = ?, K_h = ?, h = ?$$

$$\begin{aligned} K_h &= \frac{K_w}{K_a + K_b} & h &= \sqrt{K_h} \\ \therefore K_h &= \frac{1 \times 10^{-14}}{1.74 \times 10^{-5} + 1.77 \times 10^{-5}} & \therefore h &= \sqrt{3.24 \times 10^{-4}} \\ \therefore K_h &= 3.24 \times 10^{-4} & \therefore h &= 1.8 \times 10^{-2} \end{aligned}$$

$$\begin{aligned} \text{pK}_a &= -\log [\text{K}_a] \\ &= -\log [1.74 \times 10^{-5}] \\ &= -\log [1.74] - \log [10^{-5}] \\ &= -0.24 + 5.0 \\ \mathbf{\text{pK}_a} &= \mathbf{4.76} \end{aligned}$$

$$\begin{aligned} \text{pK}_b &= -\log [\text{K}_b] \\ &= -\log [1.77 \times 10^{-5}] \\ &= -\log [1.77] - \log [10^{-5}] \\ &= -0.25 + 5.0 \\ \mathbf{\text{pK}_b} &= \mathbf{4.75} \end{aligned}$$

$$\begin{aligned} \text{pH} &= \frac{1}{2}[\text{pK}_w + \text{pK}_a - \text{pK}_b] \\ &= \frac{1}{2}[14 + 4.76 - 4.75] \\ &= \frac{1}{2}[14.01] \\ &= 7.005 \end{aligned}$$

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**Strong Concepts + Smart Practice = Success**  
**Best of Luck**