

Daily Tutorial Sheet-3

JEE Advanced (Archive)

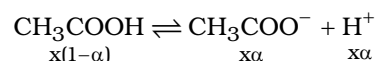


When concentration of CH_3COOH is 1.0 M, ' α ' is negligible.

$$[\text{H}^+] = \sqrt{K_a C} = 4.24 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log(4.24 \times 10^{-3}) = 2.37$$

Now, let us assume that solution is diluted to a volume where concentration of CH_3COOH (without considering ionisation) is x .



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

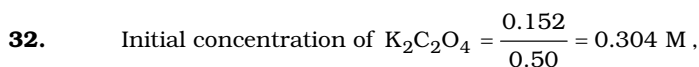
Also, desired $\text{pH} = 2 \times 2.37 = 4.74$

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

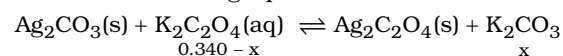
$$K_a = 1.8 \times 10^{-5} = \frac{1.8 \times 10^{-5} \alpha}{1-\alpha}$$

$$\alpha = 0.5 \text{ and } x = 3.6 \times 10^{-5} \text{ M}$$

$$\text{Volume (final)} = 1 / 3.6 \times 10^{-5} = 27.78 \times 10^3 \text{ L}$$



Also for the following equilibrium:

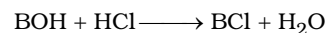
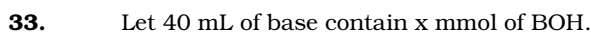


$$K = \frac{[\text{CO}_3^{2-}]}{[\text{C}_2\text{O}_4^{2-}]} \times \frac{[\text{Ag}^+]^2}{[\text{Ag}^+]^2} = \frac{K_{\text{sp}}(\text{Ag}_2\text{CO}_3)}{K_{\text{sp}}(\text{Ag}_2\text{C}_2\text{O}_4)}$$

Given, $0.304 - x = 0.0358$

$$\Rightarrow x = 0.2682 \Rightarrow K = \frac{0.2682}{0.358} = 7.5$$

$$K_{\text{sp}}(\text{Ag}_2\text{CO}_3) = K \times K_{\text{sp}}(\text{Ag}_2\text{C}_2\text{O}_4) = 7.5 \times 1.29 \times 10^{-11} = 9.675 \times 10^{-11}$$



$x - 0.5$	0.5	when 5 mL acid is added
$x - 2$	2.0	when 20 mL of acid is added

When pH is 10.04, $\text{pOH} = 3.96$ and when pH is 9.14, pOH is 4.86. Therefore,

$$3.96 = \text{p}K_b + \log \frac{0.50}{x - 0.5} \quad \dots\dots(i)$$

$$3.96 = \text{p}K_b + \log \frac{2.0}{x - 2} \quad \dots\dots(ii)$$

Subtracting Eq. (i) from Eq. (ii) gives

$$0.90 = \log \left(\frac{2}{x-2} \times \frac{x-0.5}{0.5} \right) \Rightarrow 28 = \frac{4(x-0.5)}{x-2}$$

$\Rightarrow x = 3.5$, substitution in equation (i) gives

$$3.96 = \text{p}K_b + \log \frac{0.5}{3} \Rightarrow K_b = 1.8 \times 10^{-5}$$

34. $K_{sp} = 4S^3 = 4.42 \times 10^{-5}$

$$S = 0.022\text{M}$$

mmol of Ca(OH)_2 in 500 mL saturated solution = 11

mmol of NaOH in 500 mL 0.40 M solution = 200

Total mmol of $\text{OH}^- = 200 + 2 \times 11 = 222$

$$[\text{OH}^-] = 0.222\text{ M}$$

$$\text{Solubility in presence of NaOH} = \frac{K_{sp}}{[\text{OH}^-]^2}$$

mmol of Ca^{2+} remaining in solution = 0.9

mmol of Ca(OH)_2 precipitated = 10.1

mg of Ca(OH)_2 precipitated = $10.1 \times 7.4 = 747.4\text{ mg}$

35.(D) $75\text{mL} \frac{\text{m}}{5} \text{HCl} = 15\text{mmol HCl}$

$$25\text{mL} \frac{\text{M}}{5} \text{NaOH} = 5\text{mmol NaOH}$$

After neutralization, 10mmol HCl will be remaining in 100 mL of solution.

$$\text{Molarity of HCl in the final solution} = \frac{10}{100} = 0.10$$

$$\text{pH} = -\log[\text{H}^+] = -\log(0.10) = 1$$

36. For H_2S , $\text{H}_2\text{S} \rightleftharpoons 2\text{H}^+ + \text{S}^{2-}$

$$K = K_1 \times K_2 = 1.3 \times 10^{-20}$$

Minimum $[\text{S}^{2-}]$ required to begin precipitation of

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

$$K = 1.3 \times 10^{-20} = \frac{[\text{H}^+]^2 [\text{S}^{2-}]}{[\text{H}_2\text{S}]} = [\text{H}^+]^2 \frac{(1.2 \times 10^{-19})}{0.10}$$

$$[\text{H}^+] = 0.10\text{ M} \Rightarrow \text{pH} = 1$$

37. Mixing H_2CO_3 with NaHCO_3 results in buffer solution.

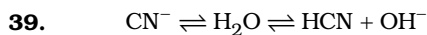
$$\text{pH} = \text{p}K_a + \log \frac{[\text{NaHCO}_3]}{[\text{H}_2\text{CO}_3]} = \text{p}K_a + \log \frac{n(\text{NaHCO}_3)}{n(\text{H}_2\text{CO}_3)}$$

$$\Rightarrow 7.4 = -\log(7.8 \times 10^{-7}) + \log \frac{x}{20} \Rightarrow x = 400\text{ mmol}$$

$$\Rightarrow \text{NaHCO}_3 = 5 \times V \Rightarrow V = 80\text{ mL}$$

38. For salts of weak acid and weak base.

$$\text{pH} = 7 + \frac{1}{2}(\text{p}K_a - \text{p}K_b) = 7 + \frac{1}{2}(3.8 - 4.8) = 6.50$$



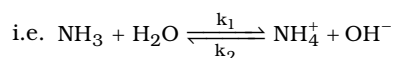
$$K_h = 2 \times 10^{-5}$$

$$[\text{OH}^-] = \sqrt{K_h C} = \sqrt{2 \times 10^{-5} \times 0.5} = \sqrt{10^{-5}}$$

$$\text{pOH} = 2.5 \text{ and } \text{pH} = 11.5$$



$$K_b(\text{NH}_3) = K_w / K_a = \frac{10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$

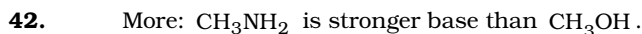


$$K = \frac{k_1}{k_2} = 1.8 \times 10^{-5} \Rightarrow k_1 = K k_2 = 1.8 \times 10^{-5} \times 3.4 \times 10^{10} = 6.12 \times 10^5$$



P_2O_5 : Acidic

Na_2O_2 and MgO : Basic



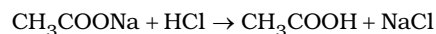
*43.(BC) pH of 10^{-8} M solution will be between 6 and 7 never 8. The conjugate base of an acid is formed by removing a proton (H^+) from acid. Therefore, HPO_4^{2-} is conjugate base of H_2PO_4^- .



Increasing temperature will increase equilibrium constant of the above endothermic reaction. At the mid-point of titration $\text{pH} = \text{p}K_a$

44.(B) NaCN is basic salt, has highest pH while HCl has lowest pH. NaCl is neutral salt has $\text{pH} = 7$ while NH_4Cl is acidic salt, has pH less than 7.

*45.(ABC) $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa} = \text{Buffer solution}$



If HCl is taken in limited quantity, final solution will have both CH_3COOH and CH_3COONa needed for buffer solution.

Ammonia and ammonium chloride forms basic buffer.