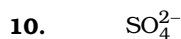


- 1.(A)** The order of acidic strength of conjugate acids is $\text{HOCl} < \text{HClO}_2 < \text{HClO}_3 < \text{HClO}_4$
Reverse is the order of basic strength of their conjugate base, i.e. ClO^- is the strongest base.
- 2.(B)** $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-6} \times 10^{-6} = 10^{-12}$
- 3.(D)** No matter, what is the concentration of HCl , its pH will always be less than 7 at 25°C . In the present case, the solution is very dilute, pH will be between 6 and 7.
- 4.(A)** Acidic buffer is prepared by mixing weak acid with salt of its conjugate base. Therefore, acetic and sodium acetate can be used to prepare acidic buffer.
- 5** For acidic buffer, the Henderson's equation is
- $$\text{pH} = \text{pK}_a + \log \frac{(\text{mole of salt})}{(\text{mole of acid})}$$
- $$4.75 = -\log(1.34 \times 10^{-5}) + \log \frac{x}{0.02}$$
- $\Rightarrow x = 0.015$ mole of sodium propionate.
- Addition of 0.01 mole HCl increase moles of propionic acid by 0.01 and moles of sodium propionate will decrease by same amount.
- New moles of acid = $0.02 + 0.01 = 0.03$
- $$\text{pH} = -\log(1.34 \times 10^{-5}) + \log \left(\frac{0.005}{0.030} \right) = 4.09$$
- pH of 0.001 $\text{HCl} = 2$, just half of the pH of final buffer solution
- 6.** Sodium acetate (CH_3COONa) is a basic salt (salt of strong base and weak acid) therefore, its aqueous solution has $\text{pH} > 7$.
- 7.(B)** For precipitation reaction, $Q_{\text{IP}} > K_{\text{sp}}$
- $$Q_{\text{IP}} = [\text{Ca}^{2+}][\text{F}^-]^2 = \left(\frac{10^{-2}}{2} \right) \times \left(\frac{10^{-3}}{2} \right)^2 = 1.25 \times 10^{-9} > K_{\text{sp}}, \text{ precipitate will be formed.}$$
- 8.** mmol of $\text{NaOH} = 20 \times 0.2 = 4$
mmol of acetic acid = $50 \times 0.2 = 10$
After neutralization, buffer solution is formed which contain 6 mmol CH_3COOH and 4 mmol CH_3COONa .
- $$\text{pH} = \text{pK}_a + \log \frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]} = -\log(1.8 \times 10^{-5}) + \log \frac{4}{6} = 4.56$$
- Now, let x mmol of NaOH is further added so that pH of the resulting buffer solution is 4.74.
Now, the buffer solution contains $(4 + x)$ mmol CH_3COONa
- $$4.74 = -\log(1.8 \times 10^{-5}) + \log \frac{4+x}{6-x}$$
- $$\Rightarrow \frac{4+x}{6-x} \Rightarrow x = 1.0 \text{ mmol} = 0.2 \times V \Rightarrow V = 5 \text{ mL of NaOH.}$$
- 9.(F)** AlCl_3 is Lewis acid because it can accept lone pair of electrons



11. Molarity (C) = 0.10

$$[\text{H}^+]\sqrt{K_a \cdot C} = 7 \times 10^{-5} \text{ M} \quad (\alpha \text{ is negligible})$$

$$\text{pH} = 4.15$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{10^{-14}}{7 \times 10^{-5}} = 1.43 \times 10^{-10} \text{ M}$$

12. $K_{sp}(\text{AgI}) = 8.5 \times 10^{-17} = [\text{Ag}^+][\text{I}^-]$

$$[\text{I}^-] \text{ required to start precipitation of AgI} = \frac{8.5 \times 10^{-17}}{0.10} = 8.5 \times 10^{-16} \text{ M}$$

$$K_{sp}(\text{HgI}_2) = 2.5 \times 10^{-26} = [\text{Hg}^{2+}][\text{I}^-]^2$$

$$[\text{I}^-] \text{ required to start precipitation of HgI}_2 = \sqrt{\frac{2.5 \times 10^{-26}}{0.10}} = 5 \times 10^{-13} \text{ M}$$

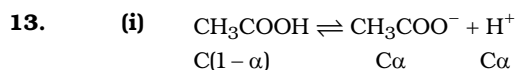
The above calculation indicates that lower $[\text{I}^-]$ is required for precipitation of AgI. When $[\text{I}^-]$ reaches to 5×10^{-13} , AgI gets precipitated almost completely.

When HgI₂ starts precipitating.

$$[\text{Ag}^+] = \frac{8.5 \times 10^{-17}}{5 \times 10^{-13}} = 1.70 \times 10^{-4} \text{ M}$$

$$\% \text{ Ag}^+ \text{ remaining} = \frac{1.70 \times 10^{-4} \times 100}{0.10} = 0.17$$

$$\% \text{ Ag}^+ \text{ precipitated} = 100 - 0.17 = 99.83$$



$$\text{If no HCl is present, } [\text{HCl}] = \frac{0.2}{2} = 0.10 \text{ M}$$

$$[\text{CH}_3\text{COOH}] = 0.10 \text{ M}$$

The major contributor of H^+ in solution is HCl.

$$K_a = \frac{\text{C}\alpha(0.1)}{\text{C}(1-\alpha)} = 1.75 \times 10^{-5}$$

$$\alpha = 1.75 \times 10^{-4} \quad (\because 1-\alpha \approx 1)$$

(ii) $\text{mmol of NaOH added} = \frac{6}{40} \times 1000 = 150$

$$\text{Mmol of HCl} = 500 \times 0.2 = 100$$

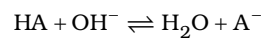
$$\text{mmol of HCl} = 500 \times 0.2 = 100$$

$$\text{After neutralisation, mmol of CH}_3\text{COOH} = 50$$

$$\text{Mmol of CH}_3\text{COONa} = 50$$

$$\text{pH} = \text{p}K_a = 4.75$$

14.(C) The reaction of HA with strong base is



$$K = \frac{[\text{A}^-]}{[\text{HA}][\text{OH}^-]} \times \frac{[\text{H}^+]}{[\text{H}^+]} = \frac{K_a}{K_w} = \frac{10^{-4}}{10^{-14}} = 10^{10}$$

15.(A) $K_a(\text{HX}) = \frac{K_w}{K_b} = 10^{-4}$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{X}^-]}{[\text{HX}]}$$

$$\Rightarrow \text{p}K_a = 4 \quad \left[\because [\text{X}^-] = [\text{HX}] \right]$$