

Ionic S-1

$$\textcircled{1} \quad [H^+] = 10^{-pH} = 10^{-13} \text{ mol/L}$$

$$\begin{aligned} \text{no. of } H^+ &= \frac{10^{-13}}{1000} \times N_A = \frac{10^{-13}}{1000} \times 6.023 \times 10^{23} \\ &= 6.023 \times 10^7 \end{aligned}$$

$$\textcircled{2} \quad (i) \quad [H^+] = \sqrt{K_w} = 3 \times 10^{-7}$$
$$pH \text{ at } 60^\circ C = 7 - \log 3 = 6.53$$

$$(ii) \quad \text{Neutral pH at } 60^\circ C = 6.53 \text{ (from (i) part)}$$

$$(a) \quad pH(6.7) > \text{Neutral pH (6.53)} \Rightarrow \text{Basic}$$

$$(b) \quad pH(6.35) < \text{Neutral pH (6.53)} \Rightarrow \text{Acidic}$$

$$\textcircled{3} \quad [H^+] = \sqrt{K_w} = 1.6 \times 10^{-7} = 16 \times 10^{-8}$$
$$pH = 8 - \log 16 = 6.8$$

$$\textcircled{4} \quad (a) \quad [H^+] = 10^{-1} M \Rightarrow pH = 1$$

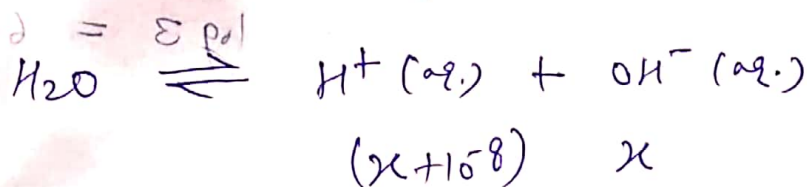
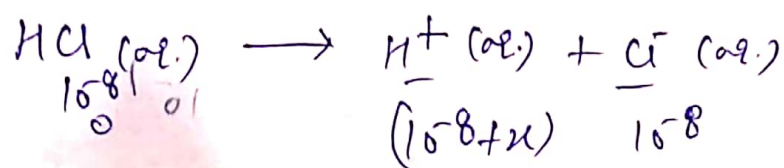
$$\begin{aligned} (b) \quad \frac{K_a}{C} < 10^{-3} &\Rightarrow [H^+] = \sqrt{K_a \cdot C} = \sqrt{1.8 \times 10^{-5} \times 0.1} \\ &= \sqrt{1.8} \times 10^{-3} \\ pH &= 3 - \log \sqrt{1.8} = 2.87 \end{aligned}$$

$$(c) \quad \frac{K_b}{C} < 10^{-3} \Rightarrow [\text{OH}^-] = \sqrt{K_b \cdot C} = \sqrt{1.8 \times 10^{-5} \times 0.1} \\ = \sqrt{1.8} \times 10^{-3}$$

$$p^{\text{OH}} = 3 - \log \sqrt{1.8} = 2.87$$

$$pH = 14 - p^{\text{OH}} = 14 - 2.87 = 11.83$$

(d)



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

$$10^{-14} = (x + 10^{-8}) x$$

$$x^2 + 10^{-8}x - 10^{-14} = 0$$

$$x = \frac{-10^{-8} \pm \sqrt{10^{-16} + 4 \times 10^{-14}}}{2}$$

$$[\text{H}^+] = x + 10^{-8} = 1.051 \times 10^{-7}$$

$$pH = 7 - \log 1.051 \Rightarrow 6.97$$

$$(e) \quad [\text{OH}^-] = 10^{-7} + 10^{-7} \approx 10^{-7}$$

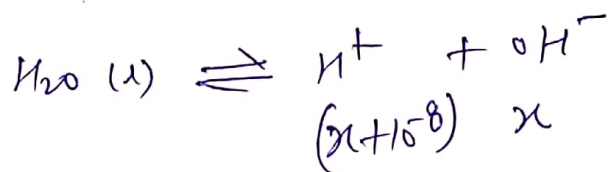
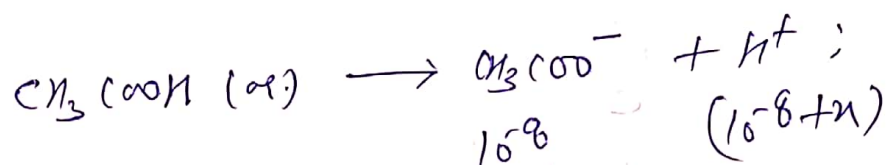
$$p^{\text{OH}} \approx 7 \Rightarrow pH = 14 - 7 = 7$$

$$(f) \quad \frac{K_a}{C} > 10 \Rightarrow \text{Behave like strong acid}$$

$$[H^+] = 10^{-6} \Rightarrow pH = 6$$

$$(g) \quad \frac{K_a}{C} = \frac{1.8 \times 10^{-5}}{10^{-8}} > 10 \Rightarrow \text{Behave like strong acid}$$

So $10^{-8} \text{ M CH}_3\text{COOH}$ is like 10^{-8} M HCl (as part (d))



$$K_w = 10^{-14} = (x+10^{-8})x$$

Solving quadratic x can be calculated

$$[H^+] = 10^{-8} + x = 1.051 \times 10^{-7}$$

$$pH = 7 - \log 1.051 = 6.97$$

$$(h) \quad [OH^-] = 2 \times M_{\text{final}} \\ = 2 \times 10^{-3} \text{ M}$$

$$M_{\text{final}} = \frac{M_{\text{initial}}}{100} \\ = \frac{10}{100} = 10^{-3}$$

$$pOH = 3 - \log 2 = 2.7 \Rightarrow pH = 14 - 2.7 = \underline{11.30}$$

$$(i) \quad [OH^-] = \frac{10^{-3}}{100} = 10^{-5} \text{ M} \Rightarrow pOH = 5 \\ \Rightarrow pH = 14 - 5 = 9$$

$$(j) \quad M_1 = 10^{-4} M \quad : \quad M_2 = 19 \times 10^{-4} M$$

Let V litre of both the two are mixed

$$[H^+]_{\text{final}} = \frac{M_1 V_1 + M_2 V_2}{V_1 + V_2} = 10^{-3} M$$

$$pH = 3$$

$$(5) (a) \quad K_a = \frac{x^2}{c} \quad ; \quad x = [H^+] = 10^{-4.5}$$

$$= \frac{(10^{-4.5})^2}{0.1} = 10^{-8}$$

$$(b) \quad K_b = \frac{x^2}{c} \quad ; \quad x = [OH^-] = 10^{-pOH} = 10^{-3.5}$$

$$= \frac{(10^{-3.5})^2}{0.1} = 10^{-6}$$

$$(6) \quad \alpha = \sqrt{\frac{K_a}{c}} \Rightarrow \frac{\alpha_2}{\alpha_1} = \sqrt{\frac{c_1}{c_2}} = \sqrt{\frac{c_1}{(c_1/100)}} = 10$$

$$(7) \quad \alpha = \sqrt{\frac{K_a}{c}} \Rightarrow \frac{\alpha_2}{\alpha_1} = \sqrt{\frac{K_{a1}}{K_{a2}}} = \sqrt{\frac{1.8 \times 10^{-5}}{6 \times 10^{-10}}}$$

$$= \frac{1.732 \times 100}{1} = 173.2 : 1$$

$$(8) \quad H^+ \text{ removed} = \text{moles of } H^+ \text{ initially} - \text{finally}$$

$$= M_1 V_1 - M_2 V_2$$

$$= 10^{-2} \times 1 - 10^{-3} \times 1$$

$$= 9 \times 10^{-3}$$

$$\textcircled{9} \quad K_b = \frac{x^2}{c-x} \quad \left| \quad x = [\text{OH}^-] = 10^{-4} \right.$$

$$10^{-5} = \frac{(10^{-4})^2}{c - 10^{-4}} \Rightarrow c = 1.1 \times 10^{-3} \text{ M}$$

$$\textcircled{10} \quad K_a = \frac{x^2}{c-x} = \frac{(10^{-3})^2}{(10^{-2} - 10^{-3})} = 1.1 \times 10^{-5}$$

$$\textcircled{11} \quad \frac{K_a}{c} < 10^{-3} \Rightarrow [\text{H}^+] = \sqrt{K_a \cdot c} = \sqrt{8 \times 10^{-10} \times 0.5}$$

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

$$\text{pH} = 5 - \log 2 = 4.7$$

$\textcircled{12}$ When equal volume are mixed

$$[\text{H}^+]_{\text{final}} = \frac{10^{-5} \times V + 10^{-3} \times V}{V + V} = 50.5 \times 10^{-5} \text{ M}$$

$$\text{pH} = 5 - \log(50.5) = \underline{\underline{3.3}}$$

$$\textcircled{13} \quad \text{(a)} \quad [\text{H}^+]_{\text{final}} = \frac{0.1 \times 50 \times 2 + 0.4 \times 50 \times 1}{(50 + 50)} = 0.3$$

$$= 3 \times 10^{-1}$$

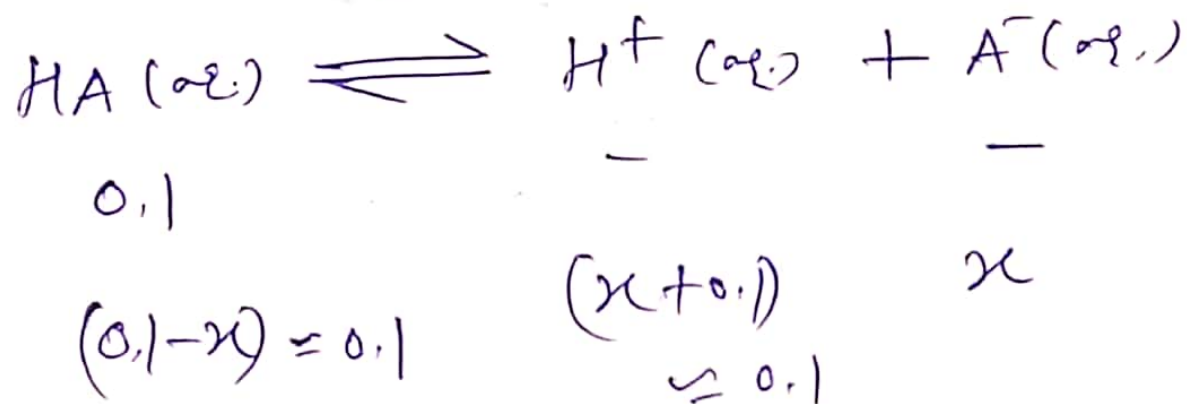
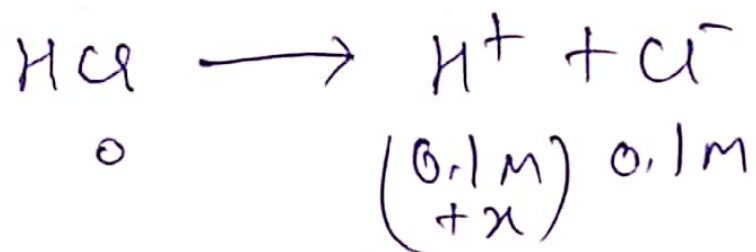
$$\text{pH} = 1 - \log 3 = 0.522$$

$$\text{(b)} \quad [\text{H}^+]_{\text{final}} = \sqrt{K_{a1}c_1 + K_{a2}c_2} \quad (\text{Mixing of weak acids})$$

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

$$\text{pH} = 3 - \log 3 = 2.522$$

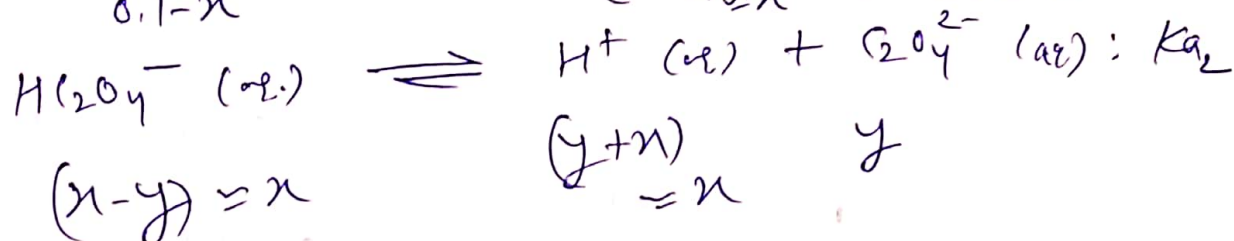
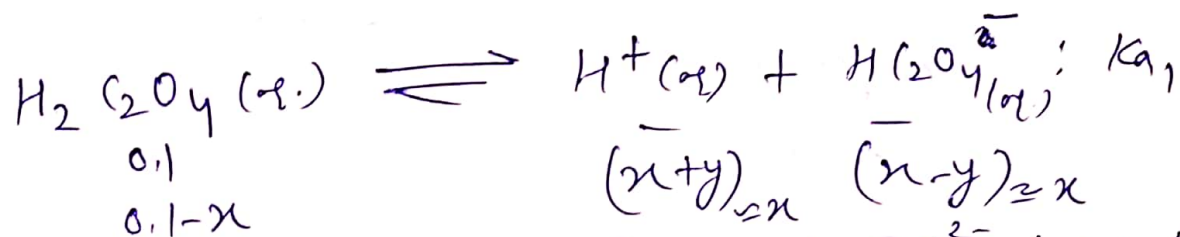
①4



$$\approx \frac{K_a}{C} < 10^{-3} \Rightarrow x \text{ can be neglected}$$

$$[\text{H}^+] = 0.1\text{M} \Rightarrow \underline{\text{pH} = 1}$$

(15)



(as $K_{a1} \gg K_{a2} \Rightarrow x \gg y$)

$$10^{-2} = K_{a1} = \frac{x^2}{0.1-x} \Rightarrow x^2 + 10^{-2}x - 10^{-3} = 0$$

$$\Rightarrow x = 2.7 \times 10^{-2} \text{ M (solving quadratic)}$$

$$K_{a2} = 10^{-5} = \frac{x \cdot y}{x} = y$$

$$[\text{H}^+] = x = 0.027 \text{ M} = [\text{HCO}_3^-]$$

$$[\text{CO}_3^{2-}] = y = 10^{-5} \text{ M} ; [\text{H}_2\text{CO}_3] = 0.1 - x = 0.073 \text{ M}$$

(16.)

Diacidic weak base gives OH^- mainly due to I step

$$K_{b1} = \frac{x^2}{c} \Rightarrow 2 \times 10^{-5} = \frac{x^2}{0.2}$$

$$x = [\text{OH}^-] = \sqrt{2 \times 10^{-5} \times 0.2} = 2 \times 10^{-3} \text{ M}$$

$$\text{pOH} = 3 - \log 2 \Rightarrow \text{pH} = 14 - \text{pOH} = 11 + \log 2 = 11.3$$

(17)

Salt of S.B & W.A

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{1.8 \times 10^{-5}} = \frac{x^2}{C}$$

$$x = [OH^-] = \sqrt{\frac{10^{-14}}{1.8 \times 10^{-5}}} \times 0.18 = 10^{-5} M$$

(18)

$$pH = 7 - \frac{1}{2} pK_b - \frac{1}{2} \log C \quad (\text{Salt of S.A \& W.B.})$$

$$= 7 - \frac{1}{2} \times (5 - \log 2) - \frac{1}{2} \log 2 = 4.5$$

(19)

Salt of S.A & W.B.

$$K_h = \frac{10^{-14}}{K_b} = \frac{x^2}{C} \quad \left| \begin{array}{l} x = [H^+] = 10^{-pH} \\ = 2 \times 10^{-3} \\ C = 0.25 \end{array} \right.$$

$$K_b = 6.25 \times 10^{-10}$$

(20)

$$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b \quad (\text{Salt of W.A \& W.B.})$$

$$= 7 \quad (\text{as } K_a = K_b)$$

$$\% h = \sqrt{\frac{10^{-14}}{K_a \times K_b}} \times 100 = 0.56 \%$$

(21)

$$\% h = \sqrt{\frac{K_h}{C} \times 100} = \sqrt{\frac{10^{-14}}{K_a \times C}} \times 100 = 1.667 \%$$

Salt of W.A & S.B.

(22)

Amphiprotic anion

$$pH = \frac{pK_{a1} + pK_{a2}}{2} = \frac{7 - \log 5 + 11 - \log 5}{2} = 8.3$$

(23)

$$a) K_h = \frac{x^2}{c} \Rightarrow x = [H^+] = \sqrt{K_h \cdot c} \\ = \sqrt{10^{-9} \times 0.001} = 10^{-6} M$$

$$\Rightarrow pH = 6$$

$$b) K_h = \frac{10^{-14}}{K_{b2}} \Rightarrow K_{b2} = \frac{10^{-14}}{K_h} = \frac{10^{-14}}{10^{-9}} = 10^{-5}$$

(24)

$$pH = pK_a + \log\left(\frac{S}{a}\right) = 9 + \log\left(\frac{0.1}{0.1}\right) = 9$$

(25)

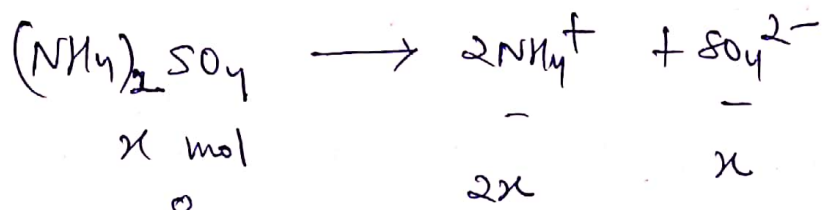
$$pOH = pK_b + \log\left(\frac{S}{b}\right) = 5 - \log 1.8 + \log\left(\frac{0.1}{0.2}\right) \\ pH = 14 - pOH = 9.56$$

(26)

$$pH = pK_a + \log\left(\frac{S}{A}\right) = 5 - \log 1.8 + \log\left(\frac{0.4}{0.4 \times \frac{500}{1000}}\right)$$

$$= 5.04$$

(27)



$$pOH = pK_b + \log\left(\frac{S}{b}\right)$$

$$(14 - 9.26) = 4.74 + \log\left(\frac{2x}{0.1}\right)$$

$$x = \frac{0.1}{2} = 0.05$$

(28)

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} = \frac{[0.1][OH^-]}{0.05}$$

1.8×10^{-5}

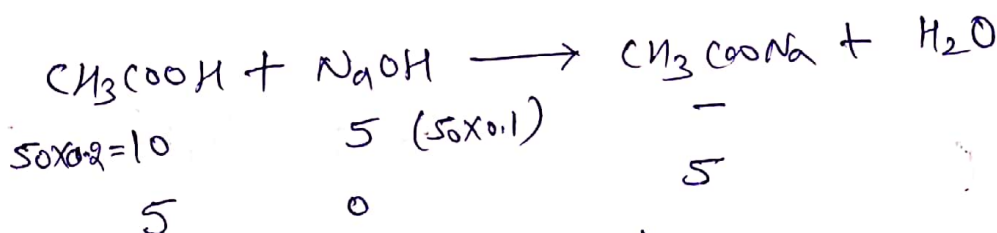
$$[OH^-] = 9 \times 10^{-6} M$$

(29)

$$pH = pK_a(HCO_3^-) + \log\left(\frac{[CO_3^{2-}]}{[HCO_3^-]}\right) = 11 - \log 4 + \log\left(\frac{0.1}{0.2}\right)$$

$$= 10.1$$

(30)



for acidic buffer formed

$$pH = pK_a + \log\left(\frac{S}{A}\right) = 4.74 + \log\left(\frac{5}{5}\right) = 4.74$$

(31)



$$75 \times 0.1 = 7.5$$

$$50 \times 0.1 = 5$$

2.5

X

5

5



Basic buffer

$$\text{pOH} = \text{pK}_b + \log\left(\frac{S}{b}\right) = (14 - 9.26) + \log\left(\frac{2.5}{5}\right)$$

$$= 14 - 9.26 - \log 2$$

$$\text{pH} = 14 - \text{pOH} = 9.26 + \log 2 = 9.56$$

(32)



20

15

-

15

5

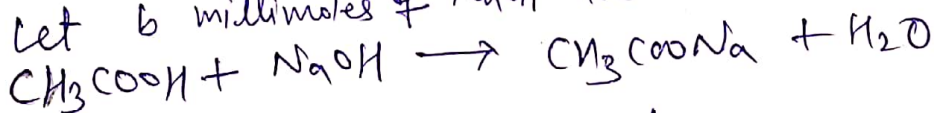
X

Basic Buffer

$$\text{pOH} = \text{pK}_b + \log\left(\frac{S}{b}\right) = 4.74 + \log\left(\frac{15}{5}\right) = 4.74 + \log 3$$

$$\text{pH} = 14 - \text{pOH} = 8.7782$$

(33)

Let b millimoles of NaOH are added

initially

10

 b

10

After

reaction

 $10 - b$

X

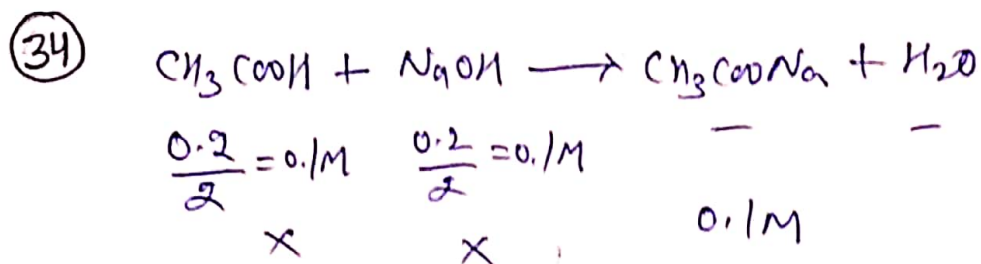
 $10 + b$

$$\text{pH}_1 = \text{pK}_a + \log\left(\frac{10}{10}\right) = \text{pK}_a$$

$$\text{pH}_2 = \text{pK}_a + \log\left(\frac{10+b}{10-b}\right)$$

$$\text{pH}_2 - \text{pH}_1 = \log\left(\frac{10+b}{10-b}\right) = 0.3 = \log 2 \Rightarrow b = \frac{10}{3} = 3.33$$

on solving



(at equivalence point equal volumes are added so concentration becomes half)

for salt of s.b. & w.a.

$$K_h = \frac{10^{-14}}{K_a} = \frac{x^2}{c} \approx 0.1$$

$$x = [\text{OH}^-] = \sqrt{\frac{10^{-14}}{K_a} \times c} = 10^{-5} \text{ M}$$

35) at equivalence point CH_3COONa forms, $c = \frac{0.1}{2} \text{ M}$

$$K_h = \frac{10^{-14}}{K_a} = \frac{x^2}{c} \quad \left| \quad x = [\text{OH}^-] \right.$$

$$[\text{OH}^-] = x = \sqrt{\frac{10^{-14}}{2 \times 10^{-5}} \times \frac{0.1}{2}} = \frac{10^{-5}}{2}$$

$$[\text{H}^+] = \frac{10^{-14}}{[\text{OH}^-]} = \frac{10^{-14}}{(10^{-5}/2)} = 2 \times 10^{-9}$$

$$\text{pH} = 9 - \log 2 = 8.7$$

36) at equivalence point NH_4Cl forms, $c = \frac{0.4}{2} = 0.2 \text{ M}$

$$K_h = \frac{10^{-14}}{K_b} = \frac{x^2}{c} \Rightarrow x = [\text{H}^+] = 10^{-5} \text{ M}$$

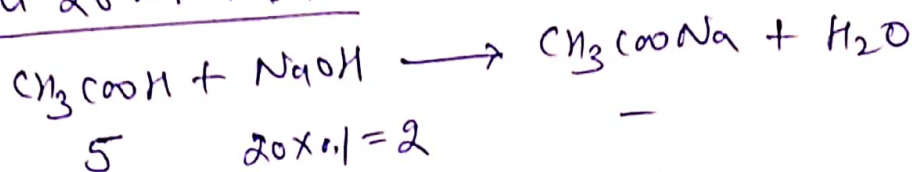
$$\text{pH} = 5$$

(37)

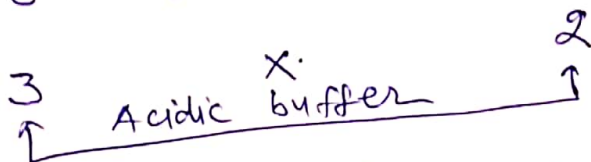
a) at 0 ml NaOH
for CH_3COOH , $(\text{H}^+) = \sqrt{K_a \cdot C} = \sqrt{2} \times 10^{-3}$

$$\text{pH} = 3 - \frac{1}{2} \log 2 = 2.85$$

b) at 20 ml NaOH



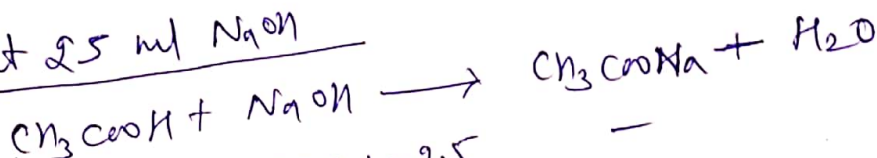
$$5 \quad 20 \times 0.1 = 2$$



Acidic buffer

$$\text{pH} = \text{p}K_a + \log\left(\frac{S}{A}\right) = (5 - \log 2) + \log\left(\frac{2}{3}\right)$$

c) at 25 ml NaOH



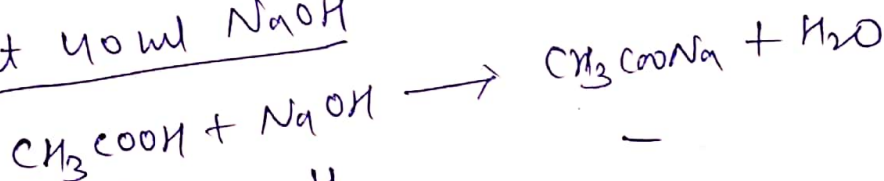
$$5 \quad 25 \times 0.1 = 2.5$$



$$\text{pH} = \text{p}K_a + \log\left(\frac{S}{A}\right) = 5 - \log 2 + \log\left(\frac{2.5}{2.5}\right)$$

$$= 4.699$$

d) at 40 ml NaOH



$$5 \quad 4$$



$$\text{pH} = \text{p}K_a + \log\left(\frac{S}{A}\right) = 5 - \log 2 + \log\left(\frac{4}{1}\right)$$

$$= 5.0310$$

(e) at 50 ml NaOH :- CH_3COONa , $C = \frac{0.1}{2} \text{ M}$

$$\text{pH} = 7 + \frac{1}{2} \text{p}K_a + \frac{1}{2} \log C = 8.699$$

(38)

$$pH = pK_{In} + \log \frac{[In^-]}{[HIn]}$$

when $[In^-] = [HIn] = 50\% \Rightarrow pH = pK_{In} = 3$

but $\frac{K_{In}}{10^{-3}} = \frac{[H^+][In^-]}{[HIn]} = 4 \times 10^{-3} \times \frac{(100-x)}{x}$

$\Rightarrow x = \% \text{ of } HIn = 80\%$

(39)

$$K_a = 6 \times 10^{-5} = \frac{[H^+][In^-]}{[HIn]} = 10^{-5} \times \left(\frac{x}{100-x} \right)$$

$\Rightarrow x = \% \text{ of basic form } (In^-) = 85.71\%$

(40)

theoretically indicator changes the colour when

$$pH = pK_{In} = 4 - \log 4 = 3.4$$

i.e. pH range = 2.4 to 4.4

So suitable for titrations involving strong acids only
like option (b) & (c)

(41)

$$pH = pK_{In} + \log \left(\frac{[In^-]}{[HIn]} \right)$$

$pH_1 = pK_{In} + \log (25/75)$		GrK-1	HIn (Red)	In^- (blue)
$pH_2 = pK_{In} + \log (75/25)$			75%	25%
$\Delta pH = pH_2 - pH_1 = 2 \log 3 = 0.954$		GrK-2	25%	75%

42) Mx :- $K_{sp} = s^2 \Rightarrow s_1 = \sqrt{K_{sp}} = 2 \times 10^{-9} \text{ M}$

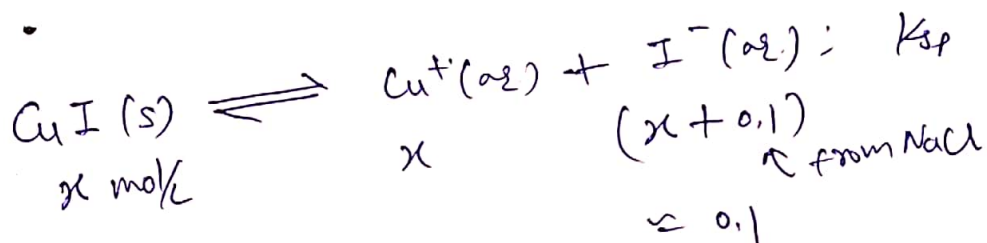
QX₂ :- $K_{sp} = 4s^3 \Rightarrow s_2 = \left(\frac{K_{sp}}{4}\right)^{1/3} = 10^{-6} \text{ M}$

$$s_2 > s_1$$

43) $S = \frac{0.0608}{304} = 2 \times 10^{-4} \text{ M}$

$$K_{sp} = s^2 = (2 \times 10^{-4})^2 = 4 \times 10^{-8}$$

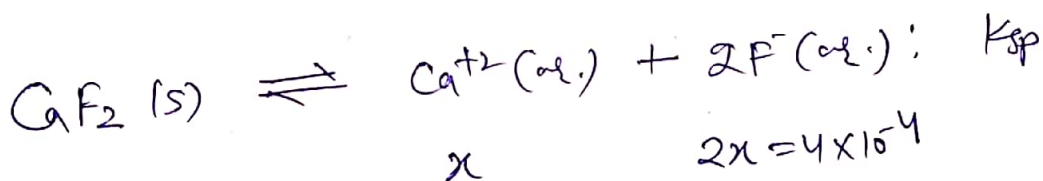
44)



$$K_{sp} = [\text{Cu}^+][\text{I}^-] = x \times 0.1$$

$$x = \frac{K_{sp}}{0.1} = 5 \times 10^{-11} \text{ M}$$

45)



$$K_{sp} = x(2x)^2 = 2 \times 10^{-4} \times (4 \times 10^{-4})^2 = 3.2 \times 10^{-11}$$

46)

$$\begin{aligned} K_{sp} &= 4s^3 \\ &= 4(4 \times 10^{-6})^3 \\ &= 2.56 \times 10^{-16} \end{aligned}$$

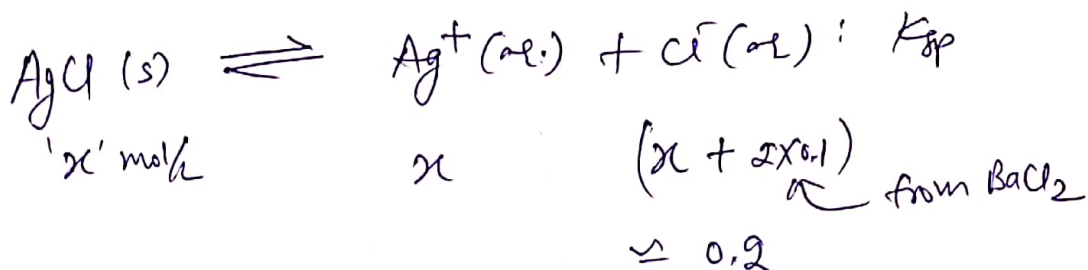
$$\begin{aligned} S &= \frac{2.4 \times 10^{-5} \times 1000}{60 \times 100} \\ &= 4 \times 10^{-6} \text{ mol/L} \end{aligned}$$

(47)

$$K_{sp} = (2)^2 (3)^3 S^5 = 108 S^5$$

$$S = \left(\frac{K_{sp}}{108} \right)^{1/5} = \left(\frac{1.08 \times 10^{-23}}{108} \right)^{1/5} = 10^{-5} M$$

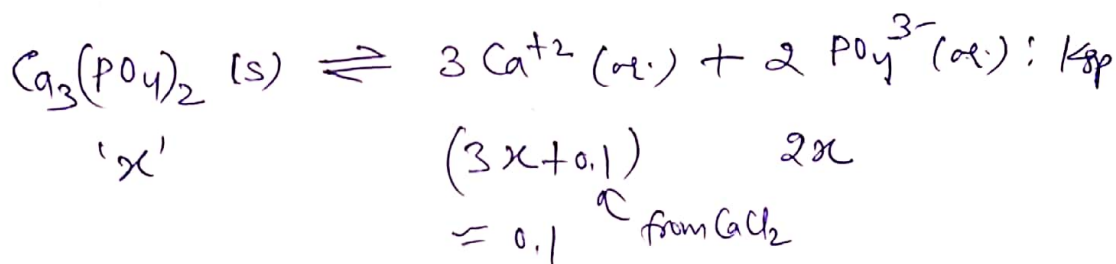
(48)



$$K_{sp} = x \times 0.2$$

$$x = \frac{K_{sp}}{0.2} = 5 \times 10^{-10}$$

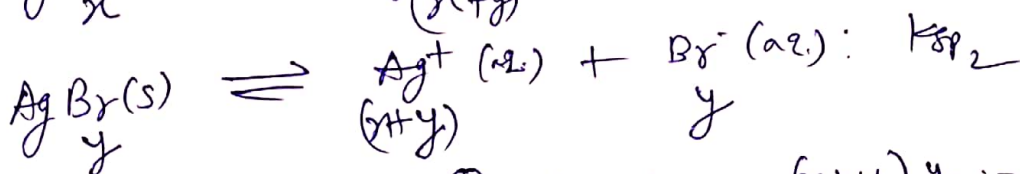
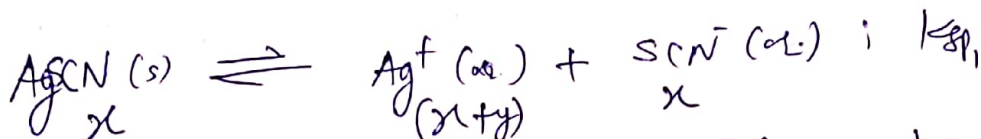
(49)



$$K_{sp} = [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2 = (0.1)^3 (2x)^2$$

$$x = \left(\frac{10^{-15}}{(0.1)^3 \times 4} \right)^{1/2} = 5 \times 10^{-7} M$$

(50)



$$K_{sp1} = (n+y)x \dots \textcircled{1} ; K_{sp2} = (n+y)y \dots \textcircled{2}$$

$$\textcircled{1} + \textcircled{2} \Rightarrow (n+y) = \sqrt{K_{sp1} + K_{sp2}} = 2 \times 10^{-6}$$

$$\Rightarrow x = 1.6 \times 10^{-4} M ; y = 4 \times 10^{-7} M$$

(51)

$$K_{sp} \times K_f = \left(\frac{x}{c-2x} \right)^2 \Rightarrow \frac{x}{c-2x} = \sqrt{K_{sp} \times K_f}$$

here $c = 0.2M$ & x is solubility

$$\frac{x}{0.2-2x} = 4 \times 10^{-4} \Rightarrow x =$$

$$\Rightarrow x = 8 \times 10^{-3} M$$

(52)

$$\frac{K_{sp}}{K_a} = \frac{x^2}{[H^+]} ; x = \text{solubility}$$

$$x = \sqrt{\frac{K_{sp}}{K_a} \times [H^+]} = \sqrt{\frac{8 \times 10^{-10} \times 10^{-3}}{5 \times 10^{-10}}} = 4 \times 10^{-2} M$$

(53)

$$(a) Q_{sp} = [Mg^{2+}] [OH^-]^2 = 10^{-3} \times (10^{-5})^2 = 10^{-13}$$

$$Q_{sp} < K_{sp} \Rightarrow \text{no ppt.}$$

$$(b) Q_{sp} = (10^{-3}) (10^{-3})^2 = 10^{-9}$$

$$Q_{sp} > K_{sp} \Rightarrow \text{ppt.}$$

(54)

$$\text{After mixing of solns.} \\ [Ag^+] = \frac{C_1 V_1}{V_1 + V_2} = \frac{2 \times 10^{-4} \times 200}{600} = \frac{2}{3} \times 10^{-4} M$$

$$[Cl^-] = \frac{C_2 V_2}{V_1 + V_2} = \frac{1.2 \times 10^{-6} \times 400}{600} = 8 \times 10^{-7} M$$

$$Q_{sp} = [Ag^+] [Cl^-] = \frac{16}{3} \times 10^{-11} < K_{sp} \Rightarrow \underline{\text{No ppt}}$$

(55)

$$K_{sp} = 8 \times 10^{-8} = \underbrace{[Ba^{2+}]}_{2 \times 10^{-5}} [SO_4^{2-}]$$

$$[SO_4^{2-}] = 4 \times 10^{-3} \text{ mol/L}$$

$$\text{mass of } Na_2SO_4 \text{ in 500 ml} = \frac{4 \times 10^{-3}}{2} \times 142 \xrightarrow{\text{Mw. of } Na_2SO_4} = 0.284 \text{ gm}$$