SOLUTIONS

Module - 2 / JEE-2022

IN-CHAPTER EXERCISES	Chemistry	Ionic Equilibrium

EXERCISE-A

1. $[HNO_3] = \dfrac{0.0315/63}{500/1000} = 10^{-3}\,M = [H^+] \quad [\because\ \alpha_{HNO_3} \approx 1]$

$[OH^-] = \dfrac{K_w}{[H^+]} = \dfrac{10^{-14}}{10^{-3}} = 10^{-11}\,M$

2. $K_{a\,(HA)} = \dfrac{c\alpha^2}{1-\alpha} = \dfrac{(0.05)(0.035)^2}{1-0.035} = 6.35 \times 10^{-5}$

3. **(i)** $0.1\,N\ HCl \Rightarrow [H^+] = 0.1\,N \equiv 0.1\,M \Rightarrow pH = 1$

(ii) $\dfrac{N}{50}\,HCl \Rightarrow [H^+] = \dfrac{1}{50}\,N \equiv \dfrac{1}{50}\,M \Rightarrow pH = 1.7$

(iii) $\dfrac{N}{100}\,H_2SO_4 \Rightarrow [H^+] = \dfrac{1}{100}\,N \equiv \dfrac{1}{100}\,M \Rightarrow pH = 2.0$

(iv) $[NaOH] = [OH^-] = \dfrac{1}{10}N \equiv \dfrac{1}{10}M \Rightarrow pOH = 1 \Rightarrow pH = 13$

(v) $pH = 3.69 \equiv 4 - 0.31 = -\log_{10}[H^+] = -\log_{10}10^{-4} - \log_{10}2 \Rightarrow [H^+] = 2 \times 10^{-4}M$

4. **(i)** $[NaOH] = \dfrac{20/40}{1} = \dfrac{1}{2}\,M \equiv [OH^-] \Rightarrow [H^+] = \dfrac{K_w}{[OH^-]} = 2 \times 10^{-14}\,M \Rightarrow pH = 13.7$

(ii) $[Ba(OH)_2] = 0.005\,M$ [Equal vol. are mixed] $\Rightarrow [OH^-] = 2 \times 0.005\,M = 0.01\,M$

$\Rightarrow pOH = 2 \Rightarrow pH = 12$

5. $K_a = 1.8 \times 10^{-5}$

(i) $\alpha = \sqrt{\dfrac{K_a}{c}} = \sqrt{1.8 \times 10^{-4}} = 1.34 \times 10^{-2}$

$pH = -\log_{10}(c\alpha) = -\log_{10}(\sqrt{1.8} \times 10^{-3}) = 2.87$

(ii) $\alpha = \sqrt{\dfrac{K_a}{c}} = \sqrt{1.8 \times 10^{-3}} = 4.24 \times 10^{-2}$

$pH = -\log_{10} c\alpha = -\log_{10}(\sqrt{1.8} \times 10^{-4}) = 3.38$

6. Mmoles $HCl = 1 \times 50 = 50$

Mmoles $NaOH = 1 \times 30 = 30 \Rightarrow$ Mmoles HCl (left) $= 20 \Rightarrow [HCl] = \dfrac{20}{80} = \dfrac{1}{4} = [H^+] \Rightarrow pH = 0.6$

7. $K_a = 4.9 \times 10^{-8};\ c = 0.1\,M \Rightarrow K_a = \dfrac{c\alpha^2}{1-\alpha}$

$\Rightarrow \alpha = \sqrt{\dfrac{K_a}{c}} = 7 \times 10^{-4};\ pH = -\log_{10}(C\alpha) = -\log_{10}(7 \times 10^{-5}) = 4.15$

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

8. **(i)** $c = 0.1\ N, \alpha = 1.3\%$

$[H^+] = c\alpha = 1.3 \times 10^{-3} M \quad \Rightarrow \quad pH = 2.89$

(ii) $\alpha = \sqrt{\frac{K_a}{c}}$ can't be used as c is low.

Solve : $K_a = \frac{c\alpha^2}{1-\alpha}$ to get $\alpha = 0.13$

$pH = -\log_{10} c\alpha = -\log_{10}(10^{-3} \times 0.13) = 3.89$

9. $pH = 2.5 \quad \Rightarrow \quad [H^+] = 10^{-2.5} = \sqrt{10} \times 10^{-3}\ M$

$$K_a = \frac{[H^+]^2}{c - [H^+]} = 4 \times 10^{-5}$$

10. Mmoles NaOH $= 0.1 \times 25 = 2.5$

(i) Mmoles HCl $= 0.1 \times 20 = 2.0$

$$\Rightarrow \quad [NaOH]_{left} = \frac{0.5}{45} = \frac{1}{90}\ M = [OH^-]$$

$$\Rightarrow \quad [H^+] = \frac{K_w}{[OH^-]} = 9 \times 10^{-13}\ M \quad \Rightarrow \quad pH = 12.04$$

(ii) Mmoles HCl $= 0.1 \times 24 = 2.4$

$$\Rightarrow \quad [NaOH]_{left} = \frac{0.1}{49} = \frac{1}{490}\ M = [OH^-]$$

$$\Rightarrow \quad [H^+] = 4.9 \times 10^{-12}\ M \quad \Rightarrow \quad pH = 11.31$$

11. **(i) (C)** To get conjugate acid, add proton (H^+) $\Rightarrow$ Ans: NH_3

(ii) (A) $pH = 6 \quad \Rightarrow \quad [H^+] = 10^{-6}\ M \quad \Rightarrow \quad [H^+]_{new} = 10^{-8}\ M$ (100 times dilution)

$\Rightarrow$ Dissociation of water has to be taken into account. $\Rightarrow$ pH will be close to 7.

(iii) (D) For pH close to 1,

$\Rightarrow$ $[H^+]$ should remain after neutralisation.

In **(D)** : Mmoles $HCl = \frac{1}{5} \times 75 = 15$ and Mmoles $NaOH = \frac{1}{5} \times 25 = 5$

$\Rightarrow$ Mmoles $HCl_{left} = 10$; $V_{sol^n} = 100$ ml $\Rightarrow$ $[HCl] = [H^+] = 0.1\ M \quad \Rightarrow \quad pH = 1$

(iv) (B) $[H^+] = [OH^-]$ (Pure water) $= 10^{-6}\ M$ (given) $\Rightarrow$ $K_w = [H^+][OH^-] = 10^{-12}$

(v) (D) $K_a \approx c\alpha^2$ [If α is small] $\Rightarrow$ $\frac{K_1}{K_2} = \frac{c_2}{c_1}$ $\Rightarrow$ $\frac{c_1}{c_2} = \frac{1}{4}$

$$\approx \frac{[H^+]^2}{c}$$

(vi) (C) $K_a[H_2O] = K_w \quad \Rightarrow \quad K_a = \frac{K_w}{[H_2O]} = \frac{10^{-14}}{55.5} = 1.8 \times 10^{-16}$

(vii) (C) pH = 13 $\Rightarrow$ $[H^+] = 10^{-13}$ M $\Rightarrow$ 1000 ml solution contains 10^{-13} moles $H^+ \equiv 10^{-13} \times 6.02 \times 10^{23}$ H^+ ions
$\Rightarrow$ 1 ml solution contains 6.02×10^7 ions of H^+.

(viii) (D) Since, K_w changes with T so, $[H^+]$ and $[OH^-]$ will change with T.

(ix) (D) $K_w \Uparrow$ as T $\Uparrow$ $\Rightarrow$ $[H^+]$ and $[OH^-] \Uparrow$ $\Rightarrow$ pH and pOH will decrease.

(x) (C) $A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$

	c	–	–
	$c - c\alpha$	$c\alpha\, x$	$c\alpha\, y$

$$K_a = \frac{[A^{y+}]^x\,[B^{x-}]^y}{[A_xB_y]} = \frac{(c\alpha\, x)^x \cdot (c\alpha\, y)^y}{c} \quad \Rightarrow \quad \alpha = \left(\frac{K_a}{x^x \cdot y^y \cdot c^{x+y-1}}\right)^{1/(x+y)}$$

(xi) (B) HCl + HCl
[pH = 1 $\Rightarrow$ $[H^+] = 10^{-1}$ M] [pH = 5 $\Rightarrow$ $[H^+] = 10^{-5}$ M]
$\Rightarrow$ when two sol^n. are mixed, H^+ contribution from the second solution, will be small.

Thus, $[H^+]_{new} \approx \frac{[H^+]_1}{2}$ [Equal volumes are mixed] $\Rightarrow$ pH will increase but will be between 1 and 2.

(xii) (A) Higher is K_a, higher is α (for same 'c') $\Rightarrow$ Higher is $c\alpha$ [= $[H^+]$] $\Rightarrow$ Lower is the pH.

(xiii) (B) $K_b = \frac{c\alpha^2}{1-\alpha}$; $K_b = 1.8 \times 10^{-5}$; c = 0.1 M $\Rightarrow$ $[OH^-] = c\alpha$.

(xiv) (C) pH = 3 to 6 $\Rightarrow$ $[H^+] = 10^{-3}$ M to 10^{-6} M $\Rightarrow$ $[H^+]$ reduced by 1000 times.

(xv) (D) HCl, Cl^- : conjugate acid - base.

CH_3COOH, $CH_3COOH_2^+$: conjugate acid - base.

EXERCISE-B

1. pH = 9.6 $\Rightarrow$ $pOH = 4.4 = pK_b + \log_{10}\frac{[Salt]}{[Base]}$

Here: $pK_b = 4.7$, [Base] = 0.18 M $\Rightarrow$ [Salt] = 0.09 M $\Rightarrow$ Moles = $0.09 \times 0.2 = 0.018$

2. $pOH = pK_b + \log_{10}\frac{[Salt]}{[Base]}$; $pK_b = 4.7$; [Salt] = 0.3 M; [Base] = 0.25 M $\Rightarrow$ pOH = 4.78 $\Rightarrow$ pH = 9.22

When NaOH is added to basic buffer, pOH $\Downarrow$. Thus, pH $\Uparrow$

$\Rightarrow$ $pOH_{new} = 4.78 - 0.6 = pK_b + \log_{10}\frac{0.3 - x}{0.25 + x}$

NH_4^+	+ OH^-	$\longrightarrow$ NH_4OH
0.3	x	0.25
0.3 – x	–	0.25 + x

Solve to get, x = 0.173 moles [V = 1 L]

3. $[CH_3COOH] = \frac{15/60}{1} = \frac{1}{4}$M ; $[CH_3COONa] = \frac{20.5}{82} = \frac{1}{4}$M

$pH = pK_a + \log_{10}\frac{[Salt]}{[Acid]} = pK_a = 4.74$

Mmoles: $CH_3COONa + H^+ \longrightarrow CH_3COOH$

250 10 250

240 – 260

$pH_{new} = pK_a + \log_{10}\frac{240/V}{260/V} \Rightarrow \Delta pH = \log_{10}\frac{24}{26} = -0.035$

4. **(i) (D)** In small intestine, pH = 8 (Basic solution) $\Rightarrow$ Shifts the dissociation of aspirin in forward direction.

In stomach, pH = 2 – 3 (Highly acidic) $\Rightarrow$ Will supress the degree of dissociation of aspirin due to common ion.

(ii)(C) $pH = pK_a + \log_{10}\frac{[Salt]}{[Acid]} \Rightarrow 5 = 6 + \log_{10}\frac{[Salt]}{[Acid]} \Rightarrow \frac{[Salt]}{[Acid]} = \frac{1}{10}$

(iii)(C) Clearly, solution of $CH_3COOH + CH_3COO^- Na^+$ is an acidic buffer.

(iv)(A) For maximum buffer capacity, [salt] = [acid]

(v)(A) Adding sodium acetate to acetic acid, will supress the dissociation of acetic acid $\Rightarrow [H^+]\Downarrow \Rightarrow$ pH increases.

(vi)(B) Adding HCl will cause the conversion of some of sodium acetate to acetic acid i.e., $pH_{new} = pK_a + \log_{10}\frac{[Salt - x]}{[Acid + x]}$

$\Rightarrow$ pH will decrease.

(vii) (C) pH of a buffer remains same on dilution.

(viii) (ABCD) Solution containing NH_4OH/CH_3COOH may also behave as a buffer if unequal Mmoles of either NH_4OH or CH_3COOH are used.

EXERCISE-C

1. $K_h = \frac{K_w}{K_b} = \frac{10^{-14}}{2\times10^{-5}} = 5\times10^{-10}$; $h = \sqrt{\frac{K_h}{c}} = \sqrt{\frac{5\times10^{-10}}{0.1}} = \sqrt{50}\times10^{-5} = 7.07\times10^{-5}$

$[H^+] = ch = 7.07\times10^{-6}$ M

2. NH_4Cl ; salt of SA - WB : $pH = 7 - \frac{1}{2}(pK_b + \log_{10} c)$

$pH = 4.5$; $pK_b = 4.7 \Rightarrow c = 2M \Rightarrow g_{NH_4Cl} = (2\times0.5)\times53.5 = 53.5$ gm

3. $K_b = 10^{-6} \Rightarrow K_h = \frac{K_w}{K_b} = 10^{-8}$

$N_2\overset{+}{H}_5Cl^-$: salt of SA - WB : $\Rightarrow h = \sqrt{\frac{K_h}{c}} = \sqrt{10^{-5}} = 3.16\times10^{-3}$

4. $C_5H_5N\overset{+}{H}Cl^-$: SA - WB: $pH = \frac{pK_w}{2} - \frac{1}{2}(pK_b + \log_{10} c) \Rightarrow pK_b = 8.82$

5. **(i)(A)** $K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{1.8\times10^{-5}} = 5.5\times10^{-10}$

(ii)(D) Basic solution $\Rightarrow$ salt of WA - SB:

(iii)(D) CH_3COONa : Salt of WA - SB: $pH = 7 + \frac{1}{2}(pK_a + \log[salt])$

(iv)(B) $NaNO_3$ and NaCl : Salt of SA - SB $\Rightarrow$ pH = 7

H_2SO_4 : Acid (pH < 7)

H_2S : VERY Weak Acid. (but pH < 7)

(V)(D) Since pH = 11 (Basic sol^n.) $\Rightarrow$ Salt must be WA - SB.

EXERCISE-D

1. Solubility of $AgCl = \sqrt{K_{SP\ AgCl}} = 1.25\times10^{-5}\ M = 1.25\times10^{-5}\times143.5\ g/L = 1.79\ mg/L$

2. Solubility of $Ca(IO_3)_2 = \frac{2.1}{390} M = 5.38\times10^{-3}\ M = x$

$$K_{SP\,Ca(IO_3)_2} = 4x^3 = 6.24\times10^{-7}\ M$$

3. $Ca(IO_3)_2(s) \rightleftharpoons Ca^{2+} + 2IO_3^-$

x

– x 2x

$Ca(NO_3)_2 \longrightarrow Ca^{2+} + 2NO_3^-$

0.1M 0.2M

$K_{SP\,Ca(IO_3)_2} = [Ca^{2+}][IO_3^-]^2 = (0.1+x)(2x)^2 \approx 0.1(2x)^2 \Rightarrow x = 1.25\times10^{-3}\ M$

4. $K_{SP} = 1.5\times10^{-9}$

(i) Solubility in water $= \sqrt{K_{SP}} = 3.87\times10^{-5}\ M = 3.87\times10^{-5}\times233.4\ gm/L = 9.03\ mg/L$

(ii) In 0.1 M $BaCl_2$ $\Rightarrow$ $[Ba^{2+}] = 0.1\ M$, $[SO_4^{2-}] = \frac{K_{SP}}{[Ba^{2+}]_{Total}} \approx \frac{K_{SP}}{[Ba^{2+}]_{Ext.}} = 1.5\times10^{-8}\ M$

$\Rightarrow$ Solubility $= 1.5\times10^{-8}\times233.4\ g/L = 3.5\times10^{-6}\ g/L$

5. $[Cl^-]_{\substack{min\\To\ start\\PPt}} = \frac{K_{SP}\ AgCl}{[Ag^+]} = \frac{2.8\times10^{-10}}{2.3\times10^{-3}} M = 1.22\times10^{-7}\ M$

6. $Ag_2C_2O_4 \rightleftharpoons 2Ag^+ + C_2O_4^{2-}$ $\quad K_{SP} = 4x^3 = 4(1.1\times10^{-4})^3 = 5.32\times10^{-12}$ $[\because\ 2x = 2.2\times10^{-4}]$

2x x

In 0.1 M $Na_2C_2O_4 \equiv [C_2O_4^{2-}] = 0.1M$ $\Rightarrow$ $(2y)^2(y+0.1) = K_{SP}$ $\quad [y \equiv$ New solubility$]$

$\Rightarrow$ $y = 3.64\times10^{-6}\ M$

7. $K_{SP} = 4\times10^{-6}$

$PbBr_2 \rightleftharpoons Pb^{2+} + 2Br^-$ $\quad K_{SP} = [Pb^{2+}][Br^-]^2 = (x)(2x)^2 \Rightarrow x = 0.01\ M$

x – –

– x 2x

8. $[Ca^{2+}] = 2.5 \times 10^{-4}$ M [After equal volumes, mixing] $\Rightarrow$ $I.P._{CaF_2} = (2.5 \times 10^{-4})(10^{-4})^2$

$[F^-] = 10^{-4}$ M $= 2.5 \times 10^{-12} < K_{SP}$ [No PPt.]

9. Use: $K_1K_2 = \dfrac{[H^+]^2[S^{2-}]}{[H_2S]} = K_a = 10^{-21}$; $[H^+] = 1$ M; $[H_2S] = 0.1$ M

$\Rightarrow$ $[S^{2-}] = 10^{-22}$ M At saturation

$I.P._{Zns} = [Zn^{2+}][S^{2-}] = 0.02 \times 10^{-22} = 2 \times 10^{-24}$

$I.P._{Cus} = [Cu^{2+}][S^{2-}] = 0.02 \times 10^{-22} = 2 \times 10^{-24} > K_{SP}$ Cus $\Rightarrow$ CuS will get precipitated.

10. $[Ca^{2+}] = 0.01$ M
$[II_2SO_4]_1 = 0.001$ M $\longrightarrow$ $[Ca^{2+}] = 0.005$ M; $[H_2SO_4] = 0.0005$ M (After mixing)
$[H_2SO_4]_2 = 0.02$ M $\longrightarrow$ $[Ca^{2+}] = 0.005$ M; $[H_2SO_4] = 0.01$ M (After mixing)

Before mixing: $I.P._{CaSO_4\,1} = (5 \times 10^{-3})(5 \times 10^{-4}) = 2.5 \times 10^{-6} < K_{SP}$

and $I.P._{CaSO_4\,2} = (5 \times 10^{-3})(10^{-2}) = 5 \times 10^{-5} < K_{SP}$ $\Rightarrow$ No precipitation.

11. **(i) (D)** $A_2X_3 \rightleftharpoons \underset{2y}{2A^{3+}} + \underset{3y}{3X^{2-}}$ $K_{SP} = (2y)^2 \cdot (3y)^3 = 108y^5$

(ii) (B) Check the I.P. of $CaF_2 = [Ca^{2+}][F^-]^2$ after halving the concentrations of Ca^{2+} and F^- (Equal volumes mixed)

(iii) (B) In water : Solubility $= \sqrt{K_{SP}} = S_0$

In 0.01 M $CaCl_2$: $S_1 = \dfrac{K_{SP}}{0.02}$; In 0.01 M NaCl : $S_2 = \dfrac{K_{SP}}{0.01}$

In 0.05 M $AgNO_3$: $S_3 = \dfrac{K_{SP}}{0.05}$

(iv) (C) $pH = 10.5$ $\Rightarrow$ $[OH^-] = 10^{-3.5}$ M $= 2x$ $\Rightarrow$ $2x = \sqrt{10} \times 10^{-4}$ M

$Mg(OH)_2 \rightleftharpoons \underset{x}{Mg^{2+}} + \underset{2x}{2OH^-}$ $K_{SP} = 4x^3 = 1.57 \times 10^{-11}$

(v) (C) $[Sr^{2+}] = \dfrac{K_{SP}\,SrCO_3}{[CO_3^{2-}]} = \dfrac{7 \times 10^{-10}}{1.2 \times 10^{-3}} = 5.83 \times 10^{-7}$ M and $[F^-] = \sqrt{\dfrac{K_{SP}\,SrF_2}{[Sr^{2+}]}} = \sqrt{\dfrac{7.9 \times 10^{-10}}{5.83 \times 10^{-3}}} = 3.68 \times 10^{-2}$ M

(vi) (A) $K_a = \dfrac{[H^+]^2[S^{2-}]}{[H_2S]}$ $\Rightarrow$ $[S^{2-}] = \dfrac{K_a \cdot [H_2S]}{[H^+]^2} = \dfrac{10^{-21} \times 0.1}{0.1^2} = 10^{-20}$ M

$\Rightarrow$ 1L contains 10^{-20} moles $\equiv 6.02 \times 10^{23} \times 10^{-20}$ ions of S^{2-}

(vii) (B) I^- from NaI will exert common ion effect on AgI.

(viii) (B) For any polyprotic acid, $K_{a_1} >> K_{a_2} >> K_{a_3}$ ....

(ix) (B) $[OH^-]_{min} = \sqrt{\frac{K_{SP}Mg(OH)_2}{[Mg^{2+}]}} = \sqrt{\frac{10^{-12}}{0.01}} = 10^{-5}M \quad \Rightarrow \quad pOH = 5 \quad \Rightarrow \quad pH = 9$

(x) (A) Look for a salt requiring min. value of $[M^+]$.

(xi) (A) $A_pB_q \rightleftharpoons pA^{q+} + qB^{p-} \qquad K_{SP} = (px)^p \cdot (qx)^q = p^p \cdot q^q \cdot x^{p+q}$

x	–	–
–	px	qx

(xii) (A) Find I.P of AgCl using I.P. = $[Ag^+]\,[Cl^-]$ with halved concentrations as equal volumes were mixed.

EXERCISE-E

1. **(i)** Mmoles NaOH = $0.1 \times 50 = 5$

Mmoles $CH_3COOH = 0.1 \times 50 = 5 \quad \Rightarrow$ Complete neutralisation $\Rightarrow \quad pH = 7 + \frac{1}{2}(pK_a + \log_{10} c) = 8.72$

$\Downarrow$

Hydrolysis

Here, $c = \text{conc.of salt} = \frac{5}{100}$

(ii) Mmoles NaOH = $0.1 \times 50 = 5$

Mmoles $CH_3COOH = 0.05 \times 50 = 2.5 \quad \Rightarrow$ Formation of basic solution.

$pOH = -\log_{10}\left(\frac{2.5}{100}\right) = 1.6 \quad \Rightarrow \quad pH = 12.4$

(iii) Mmoles NaOH = $0.05 \times 50 = 2.5$

Mmoles $CH_3COOH = 0.1 \times 50 = 5 \quad \Rightarrow$ Formation of an acidic buffer

$[Salt] = [CH_3COOH] = \frac{2.5}{100} \quad \Rightarrow \quad pH = pK_a = 4.74$

(iv) Mmoles $NH_4OH = 0.1 \times 50 = 5$

Mmoles $HCl = 0.05 \times 50 = 2.5 \quad \Rightarrow$ Buffer (Basic) Formation

$pOH = pK_b \qquad [\because [Salt] = [Base]]$

$= 4.74 \quad \Rightarrow \quad pH = 9.26$

(v) Mmoles $NH_4OH = 0.05 \times 50 = 2.5$

Mmoles $HCl = 0.1 \times 50 = 5 \quad \Rightarrow$ Acidic Sol^n : $pH = -\log_{10}\left(\frac{2.5}{100}\right) = 1.6$

(vi) Mmoles $NH_4OH = 0.05 \times 50 = 2.5$

Mmoles $CH_3COOH = 0.05 \times 50 = 2.5 \quad \Rightarrow$ WA - WB Salt: $pH = 7 + \frac{1}{2}(pK_a - pK_b) = 7$

2. Theoretically, Indicator changes its color when $[In^-] = [HIn] \quad \Rightarrow \quad pH = pK_a = 5$

3. Indicator will change its color at $pH = pK_a = 3.4$

It will be useful where pH at equivalence point is close to $pK_a \quad \Rightarrow$ Acidic solution at equivalence point.

4. **(i) (A)** $NH_4OH + HCl \quad \Rightarrow \quad NH_4Cl \quad \Rightarrow \quad pH < 7$ at equivalence point.

(ii) (C) pH indicator are weak acids/bases.

(iii) (B) $pH > 8$ at equivalence point $\quad \Rightarrow \quad$ Salt of WA - SB.

(iv) (C) Learn as a result.

(v) (D) Weak Acid - Strong Base titration $\quad \Rightarrow \quad$ pH at equivalence point > 7.

$\Rightarrow$ Choose indicator having pK_a close to pH at equivalence point.