

<H1>Additional Objective Questions:

<H2>Single Correct Choice Type

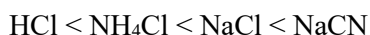
1.(C) Acetic acid is incompletely dissociated and therefore, pH cannot be calculated from its stoichiometric concentration.

2.(B) As oxidation no of chlorine is +7.

3.(B) pH defines as the concentration of H^+ ions in the solution. The order of H^+ ions in the solution is as follows:



Therefore, the order of pH is as follows:



4.(C) Since Arrhenius theory is only applicable to aqueous media.

5.(B) Since $\alpha = 1.8 \times 10^{-9}$ and for water $c = \frac{1000}{18} = 55.56$

$$[H^+] = [OH^-] = c\alpha = 55.56 \times 1.8 \times 10^{-9} = 1 \times 10^{-14}$$

$$K_w = (1 \times 10^{-7})^2 = 10^{-14}$$

$$K_d = \frac{K_w}{[H_2O]} = \frac{10^{-14}}{55.56} = 1.8 \times 10^{-16}$$

6.(D) $BeCl_2$ exists in polymeric form and so cannot act as a Lewis acid.

7.(A) Since $K_a(CH_3COOH) = K_b(NH_4OH)$ and concentration are equal so $pH(CH_3COOH) = pOH(NH_4OH)$. Given that $pH = 2.873$. Therefore, $pH = 14 - pOH = 14 - 2.873 = 11.127$.

8.(C) We know

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

$$K_h = \frac{10^{-14}}{10^{-6}} = 10^{-8}$$

9.(A) Solubility of $Al(OH)_3$ is lesser than $Zn(OH)_2$.

10.(C) $K_{sp} = [Ag^+][Cl^-]$. If S' is the solubility in presence of 0.1 M HCl.

$K_{sp} = [S'] [S' + 0.1] \approx S' \times 0.1$. (Neglecting S' compared to 0.1 in the second term).

11.(C) Since for strong acids (completely ionized) only concentration is the measure of strength but for weak (incompletely ionized) acids both degree of ionization (or K_a) and concentration will be required.

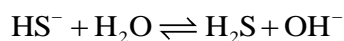
12.(A) As one can conclude from the increasing values of K_a .

13.(D) $Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$ $K_{sp} = (2 \times 2 \times 10^{-14})^2 (2 \times 10^{-4}) = 32 \times 10^{-12} = 3.2 \times 10^{-11}$.

14.(C) $pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C = 7 + \frac{1}{2} \times 4.74 + \frac{1}{2}\log(0.1)$

$pH = 8.87$ and $pOH = 14 - 8.87 = 5.13$

15.(A) From the anionic hydrolysis



$$[OH^-] = \sqrt{K_h \times C}$$

$$[OH^-] = \sqrt{\frac{K_w}{K_a} \times C}$$

At 25°C, $K_w = [H^+][OH^-]$, therefore,

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{K_w}{\sqrt{\frac{K_w}{K_a} \times C}} = \sqrt{\frac{K_w \times K_a}{C}}$$

\Rightarrow

$$[H^+] = \left(\frac{K_w \times K_a}{C} \right)^{1/2}$$

Since, $pH = -\log[H^+]$

\Rightarrow

$$pH = -\log \left(\frac{K_w \times K_a}{C} \right)^{1/2}$$

\Rightarrow

$$pH = -\log \left(\frac{K_w \times K_a}{C} \right)^{1/2}$$

\Rightarrow

$$pH = 7 + \frac{pK_a}{2} + \frac{\log C}{2}$$

16.(B) We know

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

$$K_h = \frac{10^{-14}}{1 \times 10^{-5}} = 10^{-19}$$

17.(D) $[\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.5 \times 10^{-9} (K_{sp})$ and $[\text{Ba}^{2+}] = 0.01 \text{ M}$

So, required $[\text{SO}_4^{2-}] = \frac{1.5 \times 10^{-9}}{0.01} = 1.5 \times 10^{-7}$

So, $[\text{H}_2\text{SO}_4] > 1.5 \times 10^{-7}$ for precipitation of BaSO_4 .

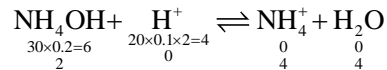
18.(D) The solubility equilibrium for the salt is $\text{M}_2\text{X}_3 \rightleftharpoons 2\text{M}^{3+} + 3\text{X}^{2-}$. Using the equation

$$K_{sp} = [\text{M}^{3+}]^2 [\text{X}^{2-}]^3$$

we get $K_{sp} = [\text{M}^{3+}]^2 [\text{X}^{2-}]^3 = (2S)^2 \times (3S)^3 = 108S^5$.

$$19.(A) h = \sqrt{\left[\frac{K_h}{C} \right]} = \sqrt{\left[\frac{K_w}{K_a \times C} \right]} = \sqrt{\left[\frac{10^{-14}}{10^{-5} \times 0.001} \right]} = 10^{-3}$$

20.(B) From the reaction involved, we have



Since, the solution is a buffer solution, therefore,

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_4\text{OH}]}$$

$$\text{pOH} = 4.7 + \log \frac{4}{2}$$

$$\text{pH} = 14 - \text{pOH} = 14 - 4.7 - \log 2$$

$$\text{pH} = 14 - 5 = 9.0$$

21.(C) Solubility of AgCl in water $= \sqrt{K_{sp}} = S_1$

In 0.01 M CaCl_2 it is given by

$$K_{sp} = S \times (0.01 \times 2 + S) \text{ therefore } S_2 = \frac{K_{sp}}{0.02}$$

In 0.01 M NaCl it is given by

$$K_{sp} = S \times (0.01 + S) \text{ therefore } S_3 = \frac{K_{sp}}{0.01}$$

In 0.05 M AgNO_3 it is given by

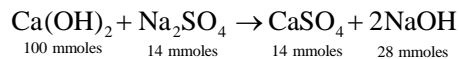
$$K_{sp} = S \times (0.05 + S) \text{ therefore } S_4 = \frac{K_{sp}}{0.05}$$

The solubilities are derived by neglecting S in comparison to 0.02, 0.01, and 0.05.

Hence, the correct order is $S_3 > S_2 > S_4 > S_1$

22.(A) Millimoles of $\text{Na}_2\text{SO}_4 = \frac{2}{142} = 14$ mmoles

From the stoichiometry of the reaction,



Millimoles of $\text{CaSO}_4 = 14$ m moles.

Therefore, the mass of calcium sulphate formed (W) = $14 \times 10^{-3} \times 136 = 1.9$ g

And Concentration of $[\text{OH}^-] = \frac{28}{100} = 0.28 \text{ mol L}^{-1}$

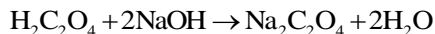
23.(C) For conjugate acid–base pair $K_a K_b = K_w$

$$K_a = \frac{10^{-14}}{10^{-10}} = 10^{-4}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$5 = 4 + \log \frac{[\text{A}^-]}{[\text{HA}]} \Rightarrow 1 = \log \frac{[\text{A}^-]}{[\text{HA}]} \Rightarrow 10 = \frac{[\text{A}^-]}{[\text{HA}]} \Rightarrow \frac{[\text{HA}]}{[\text{A}^-]} = 0.1$$

24.(A) The reaction involved is



milliequiv. of $\text{H}_2\text{C}_2\text{O}_4 =$ milliequiv. of NaOH

$$50 \times 0.5 \times 2 = 25 \times M_{\text{NaOH}} \times 1$$

$$M_{\text{NaOH}} = 2 \text{ M}$$

Now, 1000 mL solution contains = 2×40 g = 80 g NaOH

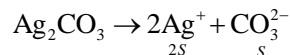
Therefore, 50 mL solution contains = $\frac{80}{1000} \times 50 = 4$ g NaOH

25.(A) $\text{CaC}_2\text{O}_4 \rightleftharpoons \text{Ca}^{++} + \text{C}_2\text{O}_4^-$; $K_{\text{sp}} = s \times s = s^2$

$$s = \sqrt{K_{\text{sp}}} = (2.5 \times 10^{-9})^{1/2} = 5 \times 10^{-5} \text{ mol L}^{-1}$$

Therefore , $\frac{w}{128} = 5 \times 10^{-5}$ Therefore $w = 6.4 \times 10^{-3}$ g

26.(C) We have



In 0.1 M AgNO_3

Concentration of $\text{Ag}^+ = 2S + 0.1$

Now,

$$K_{\text{sp}(\text{Ag}_2\text{CO}_3)} = (2S + 0.1)^2 S \tag{1}$$

On solving Eq. (1), we get

$$S = 8.0 \times 10^{-10} \text{ M}$$

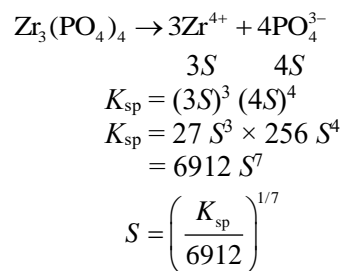
27.(A) Let H^+ ion concentration changes by x factor.

Since $pH = -\log [H_3O^+]$

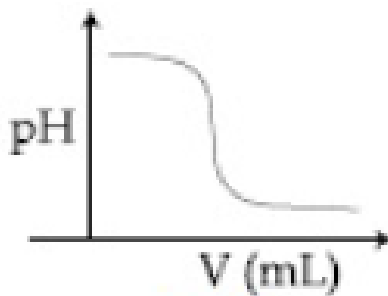
Therefore, $pH + \Delta pH = -\log \{x[H_3O^+]\} = -\log x - \log[H_3O^+]$

or $pH = -\log x = 0.3 \Rightarrow x = 0.5$.

28.(B) For $Zr_3(PO_4)_4$



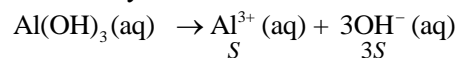
29.(A) A typical curve for the change of pH with volume of acid added is depicted as:



30.(D) We know

$$\begin{aligned} pH &= 7 - \frac{1}{2} pK_b - \frac{1}{2} \log C \\ &= 7 - \frac{5}{2} - \frac{1}{2} \log(0.02) \\ &= 7 - 2.5 + \frac{1.7}{2} \\ &= 5.35 \end{aligned}$$

31.(C) For the dissociation of aluminium hydroxide



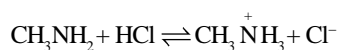
In 0.2 M NaOH, concentration of OH^- ions will be $(0.2 + 3S)$, due to common ion effect, effective OH^- ions concentration is 0.2 M.

Therefore, solubility product of Al(OH)_3 is

$$K_{\text{sp}} = S \times (0.2)^3$$

$$S = \frac{2.4 \times 10^{-24}}{0.8 \times 10^{-2}} = 3 \times 10^{-22} \text{ M}$$

32.(B) Consider the given reaction



Initial (moles)	0.1	0.08	0	0
Final (moles)	0.02	0	0.08	0.08

$$[\text{OH}^-] = K_b = \frac{[\text{CH}_3\text{NH}_2]}{[\text{CH}_3\text{NH}_3^+]}$$

$$[\text{OH}^-] = \frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{10^{-14}}{5 \times 10^{-4}} = 8 \times 10^{-11}$$

33.(B) 0.08 mol of HCl will react with 0.08 mol CH_3NH_2 and 0.02 mol CH_3NH_2 is left over, producing 0.08 mol salt $\text{CH}_3\text{NH}_3^+\text{Cl}^-$

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

$$\text{pOH} = -\log 5 \times 10^{-4} + \log \frac{0.08}{0.02} = 3.3010 + 0.6021 = +3.9031$$

$$\text{pH} = \text{pOH} = 14 - 3.9031 = 11.0969$$

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-11.0969} = 10^{-11.0969} \times 10^{-12}$$

$$= 0.8019 \times 10^{-12} = 8.019 \times 10^{-11} \text{ molL}^{-1}$$

34.(D) We know

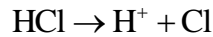
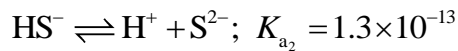
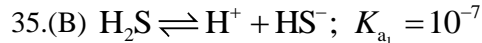
$$\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]} \quad (1)$$

Substituting the values in Eq. (1), we get

$$\text{pH} = 4.74 + \log \left[\frac{0.02 \times 200}{0.01 \times 100} \right] = 4.74 + \log 4$$

$$= 4.74 + 0.6 = 5.34$$

pH of buffer solution remain same after dilution.



Due to common ion effect the dissociation of H_2S is suppressed and the $[\text{H}^+]$ in solution is due to HCl.

$$K_{a_1} = \frac{[\text{H}^+][\text{HS}^-]}{[\text{H}_2\text{S}]}$$

$$10^{-7} = \frac{[0.3][\text{HS}^-]}{[0.1]} \quad [\text{Since } [\text{H}^+] \text{ from HCl} = 0.3]$$

$$[\text{HS}^-] = \frac{10^{-7} \times 0.1}{0.3} = 3.3 \times 10^{-8} \text{ M}$$

$$K_{a_2} = \frac{[\text{H}^+][\text{S}^{2-}]}{[\text{HS}^-]}$$

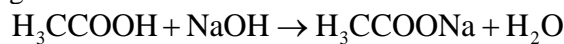
$$\text{Further } 1.3 \times 10^{-13} = \frac{[0.3][\text{S}^{2-}]}{3.3 \times 10^{-8}}$$

$$[\text{S}^{2-}] = \frac{1.3 \times 10^{-13} \times 3.3 \times 10^{-8}}{0.03} = 1.43 \times 10^{-20} \text{ M}$$

36.(D) Before NaOH addition,

$$\text{pH} = \text{p}K_a = 4.74 \quad [\text{Since } [\text{CH}_3\text{COO}^-] = [\text{CH}_3\text{COOH}]]$$

The following reaction occurs due to NaOH addition.



Initially 0.01 mol 0.01 mol 0 mol 0

Finally (0.01 - 0.001) mol 0 mol 0.001 mol

$$n_{\text{H}_3\text{CCOONa}} = 0.01 \text{ mol}$$

$$\text{After reaction, } n_{\text{CH}_3\text{COOH}} = 0.009 \text{ mol}$$

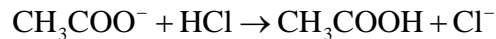
$$n_{\text{CH}_3\text{COOH}} = (0.01 + 0.001) \text{ mol} = 0.011 \text{ mol}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 4.74 + \log \frac{0.011/V}{0.009/V} = 4.74 + \log \frac{11}{9} = 4.74 + 0.872$$

$$\text{Change in pH} = \log \frac{11}{9} = 0.872$$

37.(B)

(a) The reaction is

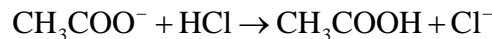


$$\begin{array}{ccc} 1 & 0.2 & 1 \\ 0.8 & 0 & 0.2 \end{array}$$

Total concentration of $\text{CH}_3\text{COOH} = 1 + 0.2 = 1.2$

$$\begin{aligned} \text{pH} &= \text{p}K_s + \log\left(\frac{\text{salt}}{\text{acid}}\right) = -\log 1.8 \times 10^{-5} + \log \frac{[0.8]}{[1.2]} \\ &= 4.7447 - 0.1761 = 4.5686 \approx 4.57 \end{aligned}$$

(b) Now for the reaction



$$\begin{array}{ccc} 0.1 & 0.2 & 0.1 \\ 0 & 0.1 & 0.2 \end{array}$$

Hence the concentration of $[\text{H}^+]$ arises only from free $\text{HCl} = 0.1\text{M} = 10^{-1}$. Hence $\text{pH} = 1$. <

38.(B) pH of buffer is given by

$$\text{pH} = -\log K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$\text{Case I: } 4 = -\log 1.0 \times 10^{-5} + \log \frac{[\text{salt}]}{(0.5)}$$

$$\text{or } [\text{salt}] = 0.1 \times 0.5 = 0.05 \text{ M}$$

$$\text{Case II: } 6 = -\log 1.0 \times 10^{-5} + \log \frac{[\text{salt}]}{(0.5)}$$

$$\text{As } \log \frac{[\text{salt}]}{(0.5)} = 1, [\text{salt}] = 10 \times 0.5 = 5 \text{ M}$$

Now the two buffer [(I. NaA = 0.05 M and HA = 0.5 M) and (II. NaA = 5 M and HA = 0.5 M)] are mixed in equal proportion.

$$\text{Thus, new concn to concentration of NaA is mixed buffer} = \frac{0.05 \times V + 5 \times V}{2V} = \frac{5.05}{2}$$

$$\text{New concentration of HA in mixed buffer} = \frac{0.5 \times V + 0.5 \times V}{2V} = 0.5 \text{ M}$$

$$\text{pH} = -\log(1.0 \times 10^{-5}) + \log \frac{[5.05/2]}{[0.5]}$$

$$\text{Thus, } \text{pH} = 5 + 0.7033 = 5.7033.$$

39.(C) NH_4^+ is a proton donor.

$$40.(C) [\text{H}^+] = C\alpha = 0.1 \times 8 \times 10^{-2}$$

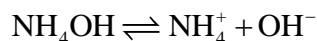
$$[\text{OH}^-] = \frac{10^{-14}}{[\text{H}^+]} = 1.25 \times 10^{-12}$$

41.(D) The minimum $[\text{OH}^-]$ at which there will be no precipitation of $\text{Mg}(\text{OH})_2$ can be obtained by

$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 9.0 \times 10^{-12} = (0.05) \times [\text{OH}^-]^2 \Rightarrow [\text{OH}^-] = 1.34 \times 10^{-5} \text{ M}$$

Thus, solution having $[\text{OH}^-] = 1.34 \times 10^{-5} \text{ M}$ will not show precipitation of $\text{Mg}(\text{OH})_2$ in 0.05 M Mg^{2+} . These hydroxyl ions are to be derived by basic buffer of NH_4Cl and NH_4OH .

$$\text{pOH} = \text{p}K_{\text{b}} + \log \frac{[\text{salt}]}{[\text{acid}]} = \text{p}K_{\text{b}} + \log \frac{[\text{NH}_4^+]}{[\text{NH}_4\text{OH}]}$$



In presence of $[\text{NH}_4\text{Cl}]$, all the NH_4^+ are provided by NH_4Cl as due to common ion effect, dissociation of NH_4OH will decrease.

$$-\log[\text{OH}^-] = -\log 1.8 \times 10^{-5} + \log \frac{[\text{NH}_4^+]}{[0.05]}$$

Therefore $\text{NH}_4^+ = 0.067 \text{ M}$ or $[\text{NH}_4\text{Cl}] = 0.067 \text{ M}$

42.(A) One can calculate ionic product from given data and for precipitation, ionic product $> K_{\text{sp}}$.

For part (A), ionic product $= (10^{-4}) \times (10^{-4}) = 10^{-8} \text{ M}$ which is greater than K_{sp} .

43.(C) H_2SO_4 is a strong acid, so it will completely dissociate and have the highest concentration of $[\text{H}^+]$. H_2S is a weak acid, so will have lower $[\text{H}^+]$. NaCl is a salt of strong acid and strong base so will give rise to neutral solution. NaNO_2 will alkaline solution on hydrolysis, so $[\text{H}^+]$ is the least.

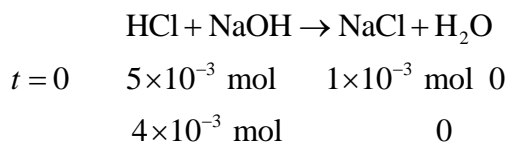
44.(B) Before adding HCl solution

$$\text{pH} = 1 \text{ [Since } [\text{HCl}] = [\text{H}_3\text{O}^+] = 10^{-1} \text{ M]}$$

$$n_{\text{HCl}} \text{ (initially)} = MV = 0.1 \text{ M} \times 0.05 \text{ L} = 5 \times 10^{-3} \text{ mol}$$

$$n_{\text{HCl}} \text{ added} = MV = 0.1 \text{ M} \times 0.01 \text{ L} = 1 \times 10^{-3} \text{ mol}$$

The reaction involved is



$$V_{\text{final}} = 100 \text{ mL} = 0.1 \text{ L}$$

$$[\text{HCl}] = \frac{n}{V} = \frac{4 \times 10^{-3} \text{ mol}}{0.1 \text{ L}} = 4 \times 10^{-2} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = 2 - \log 4 = 2 - 2 \log 2 = 2 - 2 \times 0.301 = 2 - 0.602 = 1.398$$

So, Increase in $\text{pH} = (1.398 - 1) = 0.398$

45.(D) Since the resulting solution be the acidic buffer, one may use Henderson equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

Let n mol of CH_3COONa be added to do so

or
$$\text{pH} = 4.74 + \log \frac{n \text{ mol L}^{-1}}{0.1 \text{ mol L}^{-1}}$$

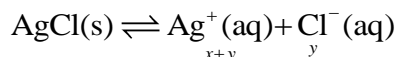
$$5 = 4.74 + \log \frac{n}{0.1} \Rightarrow \log \frac{n}{0.1} = 0.26$$

$$\frac{n}{0.1} = \text{antilog } 0.26 = 1.8197$$

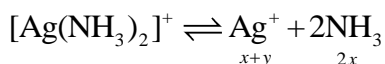
Therefore, $n = 0.18197 \text{ mol} \cong 0.182 \text{ mol}$. Amount of sodium acetate = $0.182 \times 82 \text{ g} = 14.924 \text{ g}$.

46.(B) All salt are AB type so solubility will be $\sqrt{K_{sp}}$

47.(A) Let the solubility of AgCl and $[\text{Ag}(\text{NH}_3)_2]^+$ be x and y , then



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (x+y)y$$



Given that

$$K_c = \frac{[\text{Ag}^+][\text{NH}_3]^2}{[\text{Ag}(\text{NH}_3)_2]^+} = 6.2 \times 10^{-8}$$

In case of simultaneous solubility, Ag^+ remains same in both the equilibrium

$$K_c = \frac{(x+y) \times (2x)}{[\text{Ag}(\text{NH}_3)_2]^+}$$

$$K_{sp} = (x+y)y$$

Therefore,
$$\frac{K_c}{K_{sp}} = \frac{2x}{[\text{Ag}(\text{NH}_3)_2]^+ y}$$

Given that $[\text{NH}_3] = 2x = 1 \text{ M}$.

$[\text{Ag}(\text{NH}_3)_2]^+ = [\text{Cl}^-] = y$ because Ag^+ obtained from AgCl passes in $[\text{Ag}(\text{NH}_3)_2]^+$ state

$$\frac{K_c}{K_{sp}} = \frac{1}{y \times y} \text{ or } y^2 = \frac{1.8 \times 10^{-10}}{6.2 \times 10^{-8}} = 0.29 \times 10^{-2}$$

Therefore $y = 0.539 \times 10^{-1} = 0.0539$, that is, $[\text{Ag}(\text{NH}_3)_2]^+ = 0.539$.

$$48.(B) [\text{CH}_3\text{COOH}] = \frac{6.0}{60} \times 1 = 0.1 \text{ M and } [\text{CH}_3\text{COONa}] = \frac{8.2}{82} \times 1 = 0.1 \text{ M}$$

$$\text{So, } \text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]} = 4.5 + \log \left(\frac{0.1}{0.1} \right) = 4.5$$

49.(B) Dissolution equilibria shift toward right side due to hydrolysis of cation or anion.

50.(D) Since sodium acetate is salt of (WA + SB) and its $\text{pH} > 7$.

51.(A) The buffer action of a buffer mixture is effective in the pH range $\text{p}K_a \pm 1$. It is maximum when $\text{pH} = \text{p}K_a$.

$$52.(B) K_a = 1 \times 10^{-5} \quad K_w = 1 \times 10^{-14} \quad c = 0.1$$

$$h = \sqrt{\frac{K_w}{K_b \times c}} = \sqrt{\frac{10^{-14}}{10^{-5} \times 10^{-1}}} = \sqrt{10^{-8}} = 10^{-4} = 0.0001$$

$$\text{Percentage of hydrolysis} = 0.0001 \times 100 = 0.01\%$$

$$53.(D) [\text{H}^+] \text{ from HCl} = 10^{-8}$$

$$[\text{H}^+] \text{ from H}_2\text{O} = 10^{-7}$$

$$\text{Total } [\text{H}^+] = 10^{-8} + 10^{-7} = 10^{-7} [0.1 + 1] = 1.1 \times 10^{-7} \text{ g ions L}^{-1}$$

$$\text{pH} = -\log_{10} [\text{H}^+] = -\log_{10} (1.1 \times 10^{-7}) = -\log_{10} 1.1 + 7 \log_{10} 10 = 7 - 0.0414 = 6.9$$

$$\text{pOH} = 14 - 6.9 = 7.1 \text{ which is between 7 and 8.}$$

54.(B) Given that $\text{pH} = 6$, $K_a = 10^{-5}$. So,

$$\text{p}K_a = -\log K_a = -\log_{10} 10^{-5} = 5 \times \log_{10} 10 = 5$$

Now, $\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$. Therefore,

$$6 = 5 + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$\text{or } \frac{[\text{salt}]}{[\text{acid}]} = \frac{10}{1} = 10 : 1$$

55.(A) pH of $\text{Ba}(\text{OH})_2 = 12$ and pOH of $\text{Ba}(\text{OH})_2 = 2$

$$[\text{OH}^-] = 10^{-2} \text{ and } [\text{OH}^-] = 2[\text{Ba}^{2+}] \Rightarrow [\text{Ba}^{2+}] = \frac{1}{2}[\text{OH}^-]$$

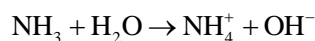
Therefore,

$$K_{\text{sp}} = \left[\frac{10^{-2}}{2} \right] [10^{-2}]^2 = \frac{10^{-6}}{2} = 5 \times 10^{-7} \text{ M}^3$$

56.(B) From the density of the ammonia, the mass of the one liter solution = $1000 \times 0.99 = 990 \text{ g}$
 Since, solution contains 10% ammonia, therefore, mass of ammonia present in solution = 99 g

$$\text{Moles of ammonia present in 1 L of solution} = \frac{99}{17} = 5.82$$

For the solution of ammonia,



Dissociation constant is

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{C \times C}{5.8}$$

We know

$$K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{10^{-14}}{5 \times 10^{-10}} = 0.2 \times 10^{-4} \quad \left[\text{Given, } K_a = 5 \times 10^{-10} \right]$$

Therefore,

$$0.2 \times 10^{-4} = \frac{C^2}{5.8}$$

After solving, we get

$$C = [\text{OH}^-] = 1.079 \times 10^{-2}$$

We know

$$[\text{H}^+] = \frac{10^{-14}}{1.079 \times 10^{-2}} = 9.27 \times 10^{-13} \text{ molL}^{-1}$$

The pH of the solution is

$$\text{pH} = -\log[\text{H}^+] = -\log[9.27 \times 10^{-13}] = 12.03$$

57.(C) Total Meq. of acid = Meq. of base = $26.6 \times 0.1 = 2.66$

Now for partial neutralization of acid

	HA	+	BOH	→	BA	+	H ₂ O
Meq. before reaction	2.66		1.2		0		0
Meq. after reaction	1.46		0		1.2		1.2

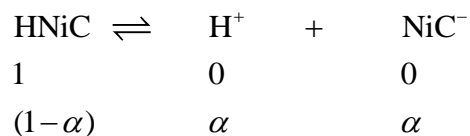
The resultant mixture acts as a buffer or

$$\text{pH} = -\log K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$5 = -\log K_a + \log \frac{1.2}{1.46}$$

Solving, we get $K_a = 8.219 \times 10^{-6}$.

58.(C) The reaction involved is

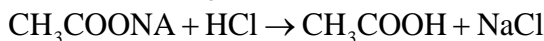


Also, $C = \frac{0.1}{2} = 5 \times 10^{-2} \text{ mol L}^{-1}$; $K_s = 1.4 \times 10^{-5}$

$$K_a = \frac{C\alpha^2}{(1-\alpha)} = C\alpha^2 \quad (\text{Since } 1-\alpha = 1)$$

Therefore,
$$\alpha = \sqrt{\left(\frac{K_a}{C}\right)} = \sqrt{\left(\frac{1.4 \times 10^{-5}}{5 \times 10^{-2}}\right)} = 1.67 \times 10^{-2} \text{ or } 1.67\%$$

59.(A) Case I: pH when 1 mol of CH_3COONa and 1 mol of HCl are present.



Before reaction 1 1 0 0

After reaction 0 0 1 1

Given that $[\text{CH}_3\text{COOH}] = 1\text{M}$

Therefore,
$$[\text{H}^+] = C\alpha = C \sqrt{\left(\frac{K_a}{C}\right)} = \sqrt{(K_a C)} = \sqrt{(K_a)} \text{ therefore } C = 1$$

or
$$\text{pH}_I = -\frac{1}{2} \log K_a$$

Case II: pH when 1 mol of CH_3COONa and 1 mol of CH_3COOH ; a acidic buffer solution forms.

Given that $[\text{salt}] = 1\text{M}$, $[\text{acid}] = 1\text{M}$. Therefore,

$$\text{pH}_{II} = -\log K_a + \log \frac{[\text{salt}]}{[\text{acid}]} = -\log K_a$$

Therefore
$$\frac{\text{pH}_I}{\text{pH}_{II}} = \frac{1}{2}$$

60.(B) The reaction involved is



When $\frac{[\text{In}^-]}{[\text{HIn}]} = 10$, we have

$$K_a = [\text{H}^+] \times 10 \Rightarrow [\text{H}^+] = \frac{1 \times 10^{-5}}{10} = 10^{-6}$$

or
$$\text{pH} = 6$$

When $\frac{[\text{In}]}{[\text{HIn}]} = \frac{1}{10}$, we have

$$K_a = [\text{H}^+] \times \frac{1}{10} \Rightarrow [\text{H}^+] = 1 \times 10^{-5} \times 10 = 10^{-4}$$

or $\text{pH} = 4$

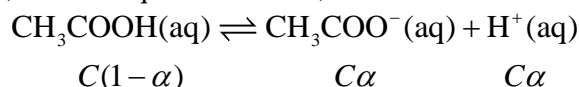
Therefore minimum change in $\text{pH} = 6 - 4 = 2$

<H2> Multiple Correct Choice Type

1. (C, D) Since, pairs in option (A) both are acid and in option (B), both are base, therefore they will not form buffer upon mixing.

2. (C, D) For weak acids and weak bases $K_a \times K_b = K_w$ of $\text{p}K_a + \text{p}K_b = \text{p}K_w$.

3. (A, B, C) $K_a(\text{CH}_3\text{COOH}) = 1.74 \times 10^{-5}$. If α is the degree of ionization of the acid and C its concentration in the solution, then for equilibrium state, we can write



We have $K_a = \frac{C\alpha \times C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$. Substituting values, we get

$$1.74 \times 10^{-5} = \frac{0.05 \times \alpha^2}{(1-\alpha)} = 0.05 \alpha^2 \quad (\alpha \ll 1)$$

Solving, we get $\alpha = \sqrt{\frac{1.74 \times 10^{-5}}{0.05}} = 1.86 \times 10^{-2}$.

$$[\text{CH}_3\text{COO}^-] = [\text{H}^+] = C\alpha = 0.05 \times 1.86 \times 10^{-2} \text{ mol L}^{-1} = 9.3 \times 10^{-4} \text{ mol L}^{-1}$$

Therefore, $\text{pH} = -\log(9.3 \times 10^{-4}) = 3.03$.

4. (A, B, C, D) After dilution $[\text{H}^+] = 10^{-2}$ or $\text{pH} = 2$

Let V liter solution of $\text{pH} = 2$ is added in original solution so that pH remains fixed.

Therefore,
$$[\text{H}^+] = \frac{10^{-2}x + V \times 10^{-2}}{X + V} = 10^{-2}$$

This result is independent of volume taken.

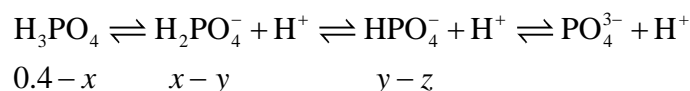
5. (B, D)

6.(A, C, D) (A) CH^- is strong conjugate base and undergoes hydrolysis so S is not equal to $\sqrt{K_{sp}}$ but more than that

(B) Weak base strong acid salt is not neutral.

(C) At first equivalence $\text{pH} = \frac{\text{p}K_{a_1} + \text{p}K_{a_2}}{2}$ not depend upon concentration

(D) $[\text{H}_3\text{PO}_4]_{\text{initial}} = \frac{0.1}{250} \times 1000 = 0.4 \text{ M}$. The reaction is



7.(A, C) $\alpha = \sqrt{\frac{K_a}{C}}$ for weak acids; and $K_a = \frac{K_w}{K_b}$

8.(B, C, D)

(A) Basic buffer is formed

(B) Slope will be minimum at $\text{pH} = \text{p}K_a$ (maximum) buffer capacity.

(C, D) Initial decrement is due to consumption of free OH^- ions, then slow decrement in pH is due to basic buffer solution and minimum slope will be there when there is best buffer action ($[\text{salt}]/[\text{base}] = 1$)

9.(A, C, D) $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$

In water, $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ $K_w = [\text{H}_3\text{O}^+]^2$

$\log K_w = 2 \log [\text{H}_3\text{O}^+]$

$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\frac{1}{2} \log K_w = -\frac{1}{2} \log \{5.6 \times 10^{-13}\} = 6.1259$ (for neutral solution)

10.(B, C) For an indication



$$[\text{H}^+] = K_a \frac{[\text{HIn}]}{[\text{In}^-]} \Rightarrow \text{pH} = \text{p}K_a + \log \frac{[\text{In}^-]}{[\text{HIn}]}$$

Substituting values, we get $9.6 = 9.0 + \log \frac{[\text{In}^-]}{[\text{HIn}]}$

Hence, $\log 4 = \log \frac{[\text{In}^-]}{[\text{HIn}]} \Rightarrow \frac{[\text{In}^-]}{[\text{HIn}]} = 4$

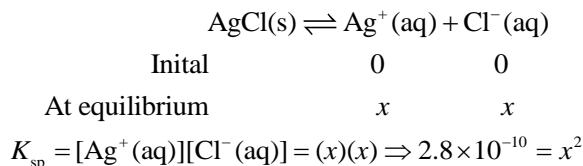
So, percentage of $[\text{In}^-] = 80\%$; percentage of $[\text{HIn}] = 20\%$.

11.(A, B) Between 0.1 M and 0.01 M there is a dilution factor of 10. If this is the only change that happens, then the $[\text{H}^+]$ should have gone down by a factor of 10 and pH would go up by one unit, that is, 5.128–

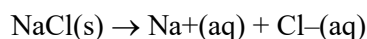
6.128. Actually, it is of 5.628 which is less and would thus correspond to a more acidic solution than expected. This means that the acid NH_4^+ ion dissociates more in a more dilute solution.

The degree of hydrolysis for 0.1 M and 0.01 M solution can be shown to be 7.5×10^{-5} and 2.4×10^{-4} , respectively.

12.(A, B) The reaction is



Solving, we get $x = 1.673 \times 10^{-5} \text{ mol L}^{-1}$.

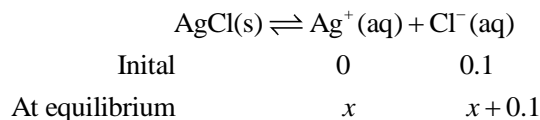


The concentrations of sodium and chloride ions are equal to the concentration of the NaCl(aq) solution.

$[\text{Na}^+(\text{aq})] = 0.10 \text{ mol/L}$ before adding AgCl(s)

$[\text{Cl}^-(\text{aq})] = 0.10 \text{ mol/L}$ before adding AgCl(s)

Now,



$$K_{\text{sp}} = [\text{Ag}^+(\text{aq})][\text{Cl}^-(\text{aq})] = (x)(x + 0.1) \Rightarrow \frac{2.8 \times 10^{-10}}{x + 0.1} = x$$

AgCl(s) has a very low solubility, so $x + 0.1 \cong 0.1$. Solving, we get $x = 2.8 \times 10^{-9} \text{ mol L}^{-1}$.

13.(A, B, D) For A, B, D ionic product > solubility product.

14.(B, C) These being strong acids would dissociate more readily thus increasing H^+ ion concentration. The ionization of phthalic acid will be suppressed due to common ion effect.

$$15.(A, C) (\text{pH})_1 = \text{p}K_a + \log\left(\frac{y}{x}\right) \text{ and } (\text{pH})_2 = \text{p}K_a + \log\left(\frac{x}{y}\right)$$

$$(\text{pH})_2 + (\text{pH})_1 = \log\left(\frac{x}{y}\right) - \log\left(\frac{y}{x}\right) = 1 \text{ or } \log\left(\frac{x^2}{y^2}\right) = 1$$

$$\text{so } \frac{x^2}{y^2} = 10 \text{ or } \frac{x}{y} = 3.162$$

$$(\text{pH})_2 + (\text{pH})_1 = 2\text{p}K_a + \log\left(\frac{y}{x}\right) + \log\left(\frac{x}{y}\right) = 2\text{p}K_a = 9.5$$

$$\text{so } \text{p}K_a = \frac{9.5}{2} = 4.75$$

<H2>Assertion–Reasoning Type

1.(B) The buffer consists of two solutes, one providing a weak Brønsted acid and the other a weak Brønsted base. Usually, the acid and base represent a conjugate pair.

2.(D) When equal volumes of the 0.002 M solutions of sodium iodate and copper sulphate are mixed, their effective concentration in the mixture becomes 0.001 M. Thus in the solution,

$$[\text{iodate ion}] = [\text{IO}_3^-] = 0.001 \text{ M}$$

$$[\text{copper ion}] = [\text{Cu}^{2+}] = 0.001 \text{ M}$$

Then, the ionic product of Cu^{2+} and IO_3^- ions in the solution relative to $\text{Cu}(\text{IO}_3)_2$ is given by,

$$\text{Ionic product in solution} = [\text{Cu}^{2+}][\text{IO}_3^-]^2 = (0.001) \times (0.001)^2 = 4 \times 10^{-9}$$

$$\text{But } K_{\text{sp}}(\text{Cu}(\text{IO}_3)_2) = 7.4 \times 10^{-8}$$

Since the ionic product in the solution relative to $\text{Cu}(\text{IO}_3)_2$ is less than the K for copper iodate, no precipitation will occur in the solution.

3.(C) Upon dilution, K_h of a salt decreases.

4.(A) Liquid NH_3 undergoes self-ionization according to the reaction,



According to the solvent system concept of acid and base, the strongest acid and base that can exist in a self-ionizing solvent are the characteristic cation and anion respectively. Thus, in liquid ammonia the strongest acid is NH_4^+ and the strongest base is NH_2^- .

Thus, all ammonium salts act as acids, and amides act as bases in liquid NH_3 .

5.(C) The value of K_{sp} of $\text{AgCl} > K_{\text{sp}}$ of AgBr , that is why AgCl is more soluble.

6.(A)

7.(A)

8.(A)

<H2>Comprehension Type

1.(B) For ZnS : $K_{\text{sp}} = [\text{Zn}^{2+}][\text{S}^{2-}] = 1 \times 10^{-21}$

For CdS : $K_{\text{sp}} = [\text{Cd}^{2+}][\text{S}^{2-}] = 8 \times 10^{-27}$

For the same concentration of sulphide ions, cadmium ions require lower concentration of sulphide ions for precipitation; hence CdS will precipitate first.

2.(A) $[\text{H}^+]$ must be greater than $[\text{Cd}^{2+}]$.

3.(A) ZnS starts precipitating when concentration of sulphide ion concentration reaches

$$[S^{2-}] = \frac{1 \times 10^{-21}}{[Zn^{2+}]}$$

The concentration of Cd^{2+} remaining when ZnS starts to precipitate is

$$= \frac{K_{sp} \text{ CdS}}{\text{sulphide concentration at which ZnS precipitates}}$$

$$= \frac{8 \times 10^{-27}}{1 \times 10^{-21}} \times [Zn^{2+}]$$

$$= 8 \times 10^{-7}$$

4.(C) At the equivalence $pH = 7$ on adding HCl, pH decreases to 5.

5.(B) For titration of weak acid with strong base,

$$pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$$

	NaOH	+	HA	\rightleftharpoons	NaA	+	H ₂ O
Initial	0.2 × 10		0.2 × 10		0		0
After	0		0		2		2

hydrolysis

$$\text{So } [NaA] = \frac{0.2}{20} = 0.1 \text{ M}$$

and given that $pK_a = 5$.

The pH of the solution will be determined by hydrolysis of salt of strong base and weak acid and is given by

$$pH = \frac{1}{2} [pK_w + pK_a + \log C] = \frac{1}{2} [14 + 5 - 1] = 9$$

6.(D) Given that [Acetic acid] = 0.5 M, [Acetate ion] = 0.5 M. Addition of 0.5 M of HCl to this buffer will lead to the following changes:

	HCl	+	CH ₃ COONa	\rightleftharpoons	CH ₃ COOH	+	NaCl
Initial	0.5		0.5		0.5		0
After hydrolysis	0		0		1.0		0.5

The concentration of the salt in the solution is $0.5/10 = 0.05 \text{ M}$. Therefore,

$$pH = 7 + \frac{1}{2} pK_a + \frac{1}{2} \log C = 7 + \frac{1}{2} [-\log(1.8 \times 10^{-5})] + \frac{1}{2} \log(0.05)$$

$$= 7 - \frac{1}{2} (4.7448 - 1.3010) = 7 - 1.7219 = 5.27$$

$$7.(A) \frac{\alpha_1}{\alpha_2} = \sqrt{\frac{C_2}{C_1}} \Rightarrow (\alpha_2)^2 = \frac{C_1(\alpha_1)^2}{C_2} = \frac{0.01 \times (0.0424) \times (0.0424)}{0.1}$$

Solving, we get $\alpha_2 = 0.0134$ or 1.34%.

$$8.(A) [H^+] = \sqrt{K_a \times C} = \sqrt{2 \times 10^{-4} \times 0.005} = 10^{-3} \text{ M. So, pH} = 3.$$

9.(A) Less is pH, stronger is the acid.

$$10.(A) \frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}} \Rightarrow K_{a_2} = 4 \times 2 \times 10^{-4} = 8 \times 10^{-4}.$$

<H2>Integer Answer Type

1. (6) Since, H_2SO_4 H_3PO_3 H_2CO_3 $H_2S_2O_7$ H_2CrO_4 H_2SO_3 have two replaceable hydrogen atoms, therefore, they are diprotic acids and H_3PO_4 H_3BO_3 are triprotic acid while H_3PO_2 is a monobasic acid.

2.(3) Since, NaCl is a strong salt, so it just provides dilution to the HCl solution, therefore, the concentration of HCl is

$$[HCl] = \frac{10 \times 0.1}{10 + 990} = \frac{1}{1000} = 10^{-3}.$$

We know

$$pH = -\log[H^+]$$

\Rightarrow

$$pH = -\log[10^{-3}] = 3$$

3.(7) From the solubility product of the base, we have

$$K_{sp} = [M^{2+}][OH^-]^2 = [0.05] \times [OH^-]^2 = 5 \times 10^{-16}$$

$$[OH^-] = (5 \times 10^{-14} \text{ M})^{1/2} = 10^{-7}$$

We know

$$pOH = -\log[OH^-] = 7$$

$$pH = 14 - 7 = 7$$

4.(9) We know

$$pOH = pK_b + \log \frac{[NH_4Cl]}{[NH_4OH]}$$

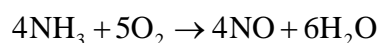
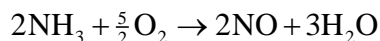
$$5 + \log \frac{0.01}{0.01} = 5$$

$$pH = 14 - pOH = 14 - 5 = 9$$

5.(8) We know

$$\begin{aligned} \text{pH} &= \left[\frac{\text{p}K_w + \text{p}K_a + \log[\text{Conc.}]}{2} \right] \\ \Rightarrow \text{pH} &= \left[\frac{14 + 4 + \log[10^{-2}]}{2} \right] \\ \Rightarrow \text{pH} &= \left[\frac{14 + 4 - 2}{2} \right] = \frac{16}{2} = 7 \end{aligned}$$

6.(1) For the reaction,



The equilibrium constant is

$$K = \frac{(\text{bar})^4 (\text{bar})^6}{(\text{bar})^4 (\text{bar})^5} = (\text{bar})^1 \Rightarrow n = 1$$

$$7.(6) \text{pH} = \frac{1}{2}(14 + \text{p}K_a - \text{p}K_b)$$

$$\frac{1}{2}(14 + 4 - 6) = 6$$

<H2>Matrix-Match Type

1.(A) \rightarrow (q), (r), (t); (B) \rightarrow (p); (C) \rightarrow (q); (D) \rightarrow (s)

NH_4Cl which is a salt of a strong acid (HCl) and a weak base (NH_4OH). It undergoes complete dissociation in

aqueous medium. The hydrolysis constant is $K_h = K_w/K_b$, and so $h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_w}{CK_b}}$

The salt is acidic so $\text{pH} < 7$ at 25°C .

NaCl is salt of strong acid and strong base, so it does not undergo hydrolysis.

CH_3COONa which is a salt of weak acid (CH_3COOH) and a strong base (NaOH). It undergoes complete

dissociation in aqueous solution. The degree of hydrolysis is $h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_w}{K_a C}}$

$\text{CH}_3\text{COONH}_4$ is a salt of a weak acid (CH_3COOH) and a weak base (NH_4OH). The salt will be partially dissociated

in aqueous medium. The degree of hydrolysis is $h = \sqrt{\frac{K_w}{K_a \times K_b}} = K_h$.

2.(A) \rightarrow (r); (B) \rightarrow (s); (C) \rightarrow (p); (D) \rightarrow (q)

(A) Diprotic weak acid, H^+ only comes from first ionization (K_1).

(B) HCrO_4 is amphoteric anion.

(C) $(\text{NH}_4)_2$ it is salt of weak base and weak acid.

(D) Hydrolysis of CrO_4^{2-} (polyvalent anion), but only first hydrolysis is appreciable.

3.(A) \rightarrow (s); (B) \rightarrow (q); (C) \rightarrow (r); (D) \rightarrow (p)

