* Chemistry [500]

1. Amongst the given options which of the following molecules/ ion acts as a Lewis acid?

   (A) OH^- (B) NH_3 (C) H_2O (D) BF_3

   Ans. : d

   Lewis acids are the one which accepts lone pair of electron due to presence of vacant orbital in outermost shell.

   $H_2\ddot{O}: \longrightarrow$ Lewis base

   $BF_3 \longrightarrow$ Lewis acid

   $:\ddot{O}:H \longrightarrow$ Lewis base

   $\ddot{N}H_3 \longrightarrow$ Lewis base

2. The pH of the solution containing $50\,mL$ each of $0.10\,M$ sodium acetate and $0.01\,M$ acetic acid is ....

   [Given pK_a of $CH_3COOH = 4.57$]

   (A) 3.57 (B) 4.57 (C) 2.57 (D) 5.57

   Ans. : d

   Weak acid (CH_3COOH) and salt of weak acid-strong base (CH_3COONa) form an acidic buffer.

   Sodium acetate $(CH_3COONa) = 0.10\,M$;

   Acetic acid $(CH_3COOH) = 0.01\,M$;

   pH of acidic buffer solution is given by

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

   $= 4.57 + \log\left(\frac{0.1}{0.01}\right)$

   $= 5.57$

3. Which among the following salt solutions is basic in nature?

   (A) Sodium acetate (B) Ammonium chloride

   (C) Ammonium sulphate (D) Ammonium nitrate

   Ans. : a

   Sodium acetate is a salt of strong base and weak acid therefore its salt solution is basic
   in nature.

4. The solubility product for a salt of the type AB is 4×10^{-8}. What is the molarity of its standard solution?

(A) $4 \times 10^{-4}\ mol/L$ (B) $2 \times 10^{-4}\ mol/L$ (C) $16 \times 10^{-16}\ mol/L$ (D) $2 \times 10^{-16}\ mol/L$

Ans. : b

For salt AB

$K_{ap} = (s)^2 \Rightarrow s = \sqrt{K_{sp}}$

$s = \sqrt{4 \times 10^{-8}} = 2 \times 10^{-4} M$

5. Find out the solubility of $Ni(OH)_2$ in $0.1 M NaOH$.

Given that the Ionic product of $Ni(OH)_2$ is 2×10^{-15}

(A) $1 \times 10^{8}\ M$ (B) $2 \times 10^{-13}\ M$

(C) $2 \times 10^{-8}\ M$ (D) $1 \times 10^{-13}\ M$

Ans. : b

$\alpha = 1$ for $NaOH$

$NaOH_{(aq)} \rightarrow \underset{0.1\ M}{Na^+}(aq) + \underset{0.1\ M}{OH^-}(aq)$

$Ni(OH)_2(s) \rightleftharpoons \underset{S'}{Ni^{+2}}(aq) + \underset{0.1+2S'}{2OH^-}(aq)$

Ionic product $= (S')(0.1 + 2S')^2$

$2 \times 10^{-15} = S'(0.1)^2$

$S' = 2 \times 10^{-13} M$

6. Conjugate base for Bronsted acids H_2O and HF are

(A) OH^- and H_2F^+ respectively

(B) H_3O^+ and F^-, respectively

(C) OH^- and F^-, respectively

(D) H_3O^+ and H_2F^+, respectively

Ans. : c

Conjugate base of H_2O is OH^-

Conjugate base of HF is F^-

7. The pH of $0.01\ M\ NaOH(aq)$ solution will be

(A) 7.01 (B) 2 (C) 12 (D) 9

Ans. : c

$\mathrm{NaOH(aq)}$ Is strong base solution

$\mathrm{So,} [\mathrm{OH^-}] = \mathrm{N} = 10^{-2}\mathrm{N}$

$\mathrm{pOH} = -\log[\mathrm{OH^-}] = -\log 10^{-2} = 2$

$\mathrm{pH} = 14 - \mathrm{pOH} = 14 - 2$

$\mathrm{pH} = 12$

8. pH of a saturated solution of $\mathrm{Ca(OH)_2}$ ts 9. The solublity product $(\mathrm{K_{sp}})$ of $\mathrm{Ca(OH)_2}$ is

(A) 0.5×10^{-15} (B) 0.25×10^{-10}

(C) 0.125×10^{-15} (D) 0.5×10^{-10}

Ans. : a

$\mathrm{Ca(OH)_2(s) \rightleftharpoons Ca^{+2}(aq) + 2OH^-(aq)}$

$\mathrm{pH = 9; pOH = 5; [OH]^- = 10^{-5} = 2S}$

$\mathrm{S = \frac{10^{-5}}{2}}$

$\mathrm{K}_{sp} - \left|\mathrm{Ca^{+2}}\right|\left|\mathrm{OH^-}\right|^2$

$\mathrm{K_{sp} = S \times (2S)^2}$

$\mathrm{K_{sp} = 4S^3}$

$\mathrm{K_{sp}} = 4 \times \left(\frac{10^{-5}}{2}\right)^3$

$\mathrm{K_{sp} = 0.5 \times 10^{-15}}$

9. Which will make basic buffer?

(A) $\mathrm{50mL}$ of $\mathrm{0.1MNaOH + 25mL}$ of $\mathrm{0.1MCH_3COOH}$

(B) $\mathrm{100mL}$ of $\mathrm{0.1MCH_3COOH + 100mL}$ of $\mathrm{0.1MNaOH}$

(C) $\mathrm{100mL}$ of $\mathrm{0.1MHCl + 200mL}$ of $\mathrm{0.1MNH_4OH}$

(D) $\mathrm{100mL}$ of $\mathrm{0.1MHCl + 100mL}$ of $\mathrm{0.1MNaOH}$

Ans. : c

Basic buffer is mixture of weak base and salt of weak base with strong acid

milli mole of $\mathrm{HCl = 100 \times 0.1 = 10}$ milli mole

milli mole of $\mathrm{NH_4OH = 200 \times 0.1 = 20mill}$ imole

$\mathrm{HCl + NH_4OH \rightarrow NH_4Cl + H_2O}$

10	20	–	–
–	10	10	–

10. Following solutions were prepared by mixing different volumes of $NaOH$ and HCl of different concentrations

(a) $\mathrm{60\ mL \frac{M}{10}\ HCl + 40\ mL \frac{M}{10}\ NaOH}$

(b) $\mathrm{55\ mL \frac{M}{10}\ HCl + 45\ mL \frac{M}{10}\ NaOH}$

(c) $\mathrm{75\ mL \frac{M}{5}\ HCl + 25\ mL \frac{M}{5}\ NaOH}$

(d) $\mathrm{100\ mL \frac{M}{10}\ HCl + 100\ mL \frac{M}{10}\ NaOH}$

pH of which one of them will be equal to 1 ?

(A) (b) (B) (a) (C) (d) (D) (c)

Ans. : d

As $\mathrm{N_1V_1 > N_2V_2}$

So acid is left at the end of reaction

$\mathrm{N_{final\ solution}} = \left[\mathrm{H^+}\right] = \frac{\mathrm{N_1V_1 - N_2V_2}}{\mathrm{V_1 + V_2}}$

$= \frac{\frac{1}{5} \times 75 - \frac{1}{5} \times 25}{75 + 25}$

$= \frac{1}{10} = 0.1$

$\mathrm{pH} = -\log[\mathrm{H}^+] = 1$

11. Concentration of the Ag^+ ions in a saturated solution of $Ag_2C_2O_4$ is $2.2 \times 10^{-4}\, mol\, L^{-1}$. Solubility product of $Ag_2C_2O_4$ is

(A) 2.66×10^{-12} (B) 4.5×10^{-11} (C) 5.3×10^{-12} (D) 2.42×10^{-8}

Ans. : c

$\mathrm{Ag_2C_2O_4} \rightleftharpoons 2\mathrm{Ag}^+ + \mathrm{C_2O_4^{2-}}$

$2.2 \times 10^{-4}\mathrm{M} \quad 1.1 \times 10^{-4}\mathrm{M}$

$\mathrm{K_{sp}} = [\mathrm{Ag}^+]^2 [\mathrm{C_2O_4^{2-}}]$

$= [2.2 \times 10^{-4}]^2 \cdot [1.1 \times 10^{-4}]$

$\mathrm{K_{sp}} = 5.3 \times 10^{-12}$

12. The solubility of $AgCl_{(s)}$ with solubility product 1.6×10^{-10} in $0.1\, M\, NaCl$ solution would be

(A) $1.26 \times 10^{-5}\, M$ (B) $1.6 \times 10^{-9}\, M$ (C) $1.6 \times 10^{-11}\, M$ (D) zero.

Ans. : b

$\mathrm{AgCl} \rightleftharpoons \mathrm{Ag}^+ + \mathrm{Cl}^-$

Ag^+ is S

Cl^- is 0.1

$\mathrm{K_{sp}} = [\mathrm{Ag}^+][\mathrm{Cl}^-]$

$1.6 \times 10^{-10} = S \times 0.1$

$1.6 \times 10^{-9} = S$

13. The percentage of pyridine (C_5H_5N) that forms pyridinium ion $(C_5H_5N^+H)$ in a $0.10\, M$ aqueous pyridine solution (K_b for $C_5H_5N = 1.7 \times 10^{-9}$) is

(A) 0.0060% (B) 0.013% (C) 0.77% (D) 1.6%

Ans. : b

$\mathrm{C_5H_5N + H_5O \longrightarrow C_5H_5N^+H + OH^-}$

$\alpha = \sqrt{\frac{K_b}{c}} = \sqrt{\frac{1.7 \times 10^{-9}}{0.1}} = \sqrt{1.7 \times 10^{-8}}$

$\% a = 1.3 \times 10^{-4} \times 100$

$= 1.3 \times 10^{-2} = 0.013$

14. What is the pH of the resulting solution when equal volumes of $0.1\, M\, NaOH$ and $0.01\, M\, HCl$ are mixed?

(A) 2 (B) 7 (C) 1.04 (D) 12.65

Ans. : d

One mole of $NaOH$ is completely neutralised by one mole of HCL.

Hence, 0.01 mole of $NaOH$ will be completely neutralised by 0.01 mole of HCL.

$\Rightarrow NaOH$ left unneutralised $= 0.1 - 0.01 = 0.09$ mol As equal volumes of two solutions are mixed,

$[\mathrm{OH}]^- = \frac{0.09}{2} = 0.045\mathrm{M}$

$\Rightarrow \mathrm{pOH} = -\log(0.045) = 1.35$

$\therefore \mathrm{pH} = 14 - 1.35 = 12.65$

15. Which one of the following pairs of solution is not an acidic buffer?

(A) CH_3COOH and CH_3COONa (B) H_2CO_3 and Na_2CO_3

(C) H_3PO_4 and Na_3PO_4 (D) $HClO_4$ and $NaClO_4$

Ans. : d

A buffer solution is one which resists changes in pH when small quantities of an acid or an alkali are added to it.

$\mathrm{HClO_4}$ and $\mathrm{NaClO_4}$ are not an acidic buffer because strong acid with its salt cannot form buffer solution.

16. Which of the following salts will give highest pH in water?

(A) KCl (B) $NaCl$ (C) Na_2CO_3 (D) $CuSO_4$

Ans. : c

$\mathrm{Na_2CO_3}$ which is a salt of NaOH(strong base) and $\mathrm{H_2CO_3}$ (weak acid) will produce a basic solution with pH greater than 7.

17. The pH of $1\ N\ H_2O$ is

(A) 7 (B) > 7 (C) < 7 (D) 0

Ans. : (a)Because pure water has a $7\,pH$.

18. The pH of $10^{-7} N\ HCl$ is

(A) 6 (B) 6.97 (C) 8 (D) 10

Ans. : (b) $10^{-7} N\ HCl$ means $(H^+) = 10^{-7}\ M$

$pH = -\log(H^+), \quad pH = 7$

19. What is the PH of $10^{-9}\ M$ aqueous HCl solution

(A) 9 (B) between 6 to 7 (C) 7 (D) can't determine

Ans. : b

Given that,

$HCl = 10^{-9}\ M = 0.01 \times 10^{-7}$

When HCl added to water the total H^+ ions

$H^+ = 0.01 \times 10^{-7} + 10^{-7}$

$H^+ = 1.01 \times 10^{-7}$

$pH = -\log\left(1.01 \times 10^{-7}\right)$

$pH = 6.9$

20. If a solution of $10^{-6}\ M\ HCl$ is diluted 100 times, the pH of solution is

(A) 8 (B) 6 (C) 6.96 (D) 7.04

Ans. : c

$[H^+] = 10^{-8} + 10^{-7}$ from HCl from H_2O

21. Which of the following has highest pH

(A) $0.1\,M\,HCl$ (B) $0.2\,M\,HCl$ (C) $0.1\,M\,CH_3COOH$ (D) $0.15\,M\,HNO_3$

Ans. : (C) $0.1\,M\,CH_3COOH$

22. If two acids of equimolar concentration are taken then which option is correct

(A) $\alpha_1^2\,K_{a_1} = \alpha_2^2\,K_{a_2}$ (B) $\alpha_1 K_{a_1}^2 = \alpha_2\,K_{a_2}^2$ (C) $\alpha_1^2\,K_{a_2} = \alpha_2^2\,K_{a_1}$ (D) $\alpha_1 K_{a_2}^2 = \alpha_2\,K_{a_1}^2$

Ans. : c

$\alpha = \sqrt{\frac{K_a}{C}}\,(C = \text{ Constant for equimolar solution })$

$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}}$

$\alpha_1\sqrt{K_{a_2}} = \alpha_2\sqrt{K_{a_1}}$

$\alpha_1^2 K_{a_2} = \alpha_2^2 K_{a_1}$

23. Number of equivalents of H_2SO_4 present in $100\,mL$ of its solution. Whose pH is 5.

(A) 10^{-4} (B) 10^{-6} (C) 10^{-2} (D) 10^{-5}

Ans. : $PH = 5 \Rightarrow [H^+] = 10^{-5}N = N_{H_2SO_4}$

Eq. of $H_2SO_4 = 10^{-5} \times 0.1 = 10^{-6}$ equivalents

24. The aqueous solution of ammonium chloride is

(A) Neutral (B) Basic (C) Acidic (D) Amphoteric

Ans. : (c)NH_4Cl is a salt of weak base $(NH_4OH)$and strong acid (HCl) .

25. Hydrolysis constant for a salt of weak acid and weak base would be

(A) $K_h = \frac{K_w}{K_a}$ (B) $K_h = \frac{K_w}{K_b}$ (C) $K_h = \frac{K_w}{K_a K_b}$ (D) None of these

Ans. : (c)$K_h = \frac{K_w}{K_a \times K_b}$

26. In hydrolysis of a salt of weak acid and strong base, $A^- + H_2O \rightleftharpoons HA + OH^-$, the hydrolysis constant (K_h) is equal to....

(A) $\frac{K_w}{K_a}$ (B) $\frac{K_w}{K_b}$ (C) $\sqrt{\frac{K_a}{C}}$ (D) $\frac{K_w}{K_a \times K_b}$

Ans. : (a) Hydrolysis constant $h = \frac{K_w}{K_a}$

27. On adding solid potassium cyanide to water

(A) pH will increase

(B) pH will decrease

(C) pH will not change

(D) Electrical conductance will not change

Ans. : (a) $CN^- + H_2O \rightarrow HCN + OH^-$

Because OH^- concentration is increased.

28. pH of $NaCl$ solution is

(A) 7 (B) 0 (C) > 7 (D) < 7

Ans. : (a)Because of $NaCl$ is a salt of strong acid and strong base. So that it is neutral.

29. A solution of sodium chloride in contact with atmosphere has a pH of about

(A) 3.5 (B) 5 (C) 7 (D) 1.4

Ans. : (c)When strong acid and strong base are react neutral salt are formed. So that $NaCl$ is a neutral salt.

30. By adding $20\ ml\ 0.1\ N\ HCl$ to $20\ ml\ 0.1\ N\ KOH$, the pH of the obtained solution will be

(A) 0 (B) 7 (C) 2 (D) 9

Ans. : (b) Neutralization reaction will takes place and form salt of strong acid and strong base. Which does not hydrolysed and thus $pH = 7$.

31. $50\ ml$ of $10^{-3}\ M\ HCl$ is mixed with $50\ ml$ of $0.05\ M\ CH_3COOH\ (pK_a = 4.7)$. The degree of dissociation of water in the resulting solution is

(A) 1.8×10^{-13} (B) 1.8×10^{-15} (C) 1.8×10^{-14} (D) 3.6×10^{-13}

Ans. : a

pH of $\mathrm{CH_3COOH} = \frac{1}{2}(\mathrm{pK}_a - \log c) = 3$

Mixing of HCl solution does not change the $[\mathrm{H}^+]$

Final solution has $[\mathrm{H}^+] = 10^{-3}\,\mathrm{M}$ therefore

$[\mathrm{OH}] = 10^{-11}\,\mathrm{M}$

Therefore $\alpha_{water} = \frac{10^{-11}}{1000/18} = 1.8 \times 10^{-13}$

32. If $1\,mL$ of $1\,M\ HCl$ solution is added to $99\,mL$ of aqueous solution of $NaCl$ so that pH of $NaCl$ solution change by X units. The value of X is

(A) 2 (B) 5 (C) 7 (D) 1

Ans. : (B) 5

33. Match the column I with column II and mark the appropriate choice

Column I	Column II
$(A)\ CH_3COONa$	(i) Almost neutral, $pH > 7$ or < 7
$(B)\ NH_4Cl$	(ii) Acidic, $pH < 7$
$(C)\ NaNO_3$	(iii) Alkaline, $pH > 7$

$(D)\ CH_3COONH_4$	(iv) Neutral, $pH=7$

(A) $A \rightarrow (i), B \rightarrow (ii), C \rightarrow (iii), D \rightarrow (iv)$

(B) $A \rightarrow (ii), B \rightarrow (iii), C \rightarrow (iv), D \rightarrow (i)$

(C) $A \rightarrow (iii), B \rightarrow (ii), C \rightarrow (iv), D \rightarrow (iv)$

(D) $A \rightarrow (iv), B \rightarrow (i), C \rightarrow (iii), D \rightarrow (ii)$

Ans. : c

$\mathrm{CH_3COONa}$ (salt of WA and SB) $\therefore \mathrm{pH} > 7$

$\mathrm{NH_4Cl}$(salt of SA and WB) $\therefore \mathrm{pH} < 7$

$\mathrm{NaNO_3}$ (salt of SA and SB) : Neutral $\mathrm{pH} = 7$

$\mathrm{CH_3CONH_4}$ (Salt of WA and WB) : Acidic or basic depends on $\mathrm{pK_a}$ and $\mathrm{pK_b}$

34. Correct order of pH of $0.1\,M$ solution is

(A) $NaCl < NH_4Cl < NaCN < HCl$

(B) $NaCN < NH_4Cl < NaCl < HCl$

(C) $HCl < NaCl < NaCN < NH_4Cl$

(D) $HCl < NH_4Cl < NaCl < NaCN$

Ans. : d

$$\underset{SA}{HCl} < \underset{Weekly\ Acid}{NH_4Cl} < \underset{Neutral}{NaCl} < \underset{Weekly\ Basic}{NaCN}$$

35. A weak acid HX has the dissociation constant $1 \times 10^{-5}\,M$. It forms a salt NaX on reaction with alkali. The percentage hydrolysis of $0.1\,M$ solution of NaX is .... %

(A) 0.0001 (B) 0.01 (C) 0.1 (D) 0.15

Ans. : b

NaX is a weak acid-strong base salt.

So, hydrolysis of NaX will be

$$h = \sqrt{\frac{K_w}{K_a \times C}} = \sqrt{\frac{10^{-14}}{10^{-5} \times 0.1}} = 10^{-4}$$

Therefore, % hydrolysis will be $= 10^{-4} \times 100 = 0.01\%$

36. Approximate pH of $0.01\ M\ NaHA$ is calculated by : ($K_{a_1} = 10^{-6}$ and $K_{a_2} = 10^{-8}$ are ionization constant of H_2A)

(A) $pH = 7 + \frac{pK_{a_1}}{2} + \frac{\log\ C}{2}$

(B) $pH = 7 - \frac{pK_{a_1}}{2} - \frac{\log\ C}{2}$

(C) $pH = \frac{pK_{a_1} + pK_{a_2}}{2}$

(D) $pH = \frac{pK_{a_1} - pK_{a_2}}{2}$

Ans. : c

Given,

$ka_1 = 10^{-6}$

$ka_2 = 10^{-8}$

$NaHA$ is the amphеprotic salt for ampheprotec sait,

$pH = \frac{pka_1 + pka_2}{2}$

37. The correct formula to calculate the hydroxyl ion concentration of an aqueous solution of NH_4NO_3 is

(A) $\sqrt{\frac{C \times K_w}{K_b}}$ (B) $\sqrt{\frac{K_w \times K_b}{C}}$ (C) $\sqrt{\frac{C \times K_w}{K_a}}$ (D) $\sqrt{\frac{K_a \times K_w}{C}}$

Ans. : (B) $\sqrt{\frac{K_w \times K_b}{C}}$

38. A weak acid HX has the dissociation constant $1 \times 10^{-5}\,M$. It forms a salt NaX on reaction with alkali. The percentage hydrolysis of $0.1\,M$ solution of NaX is .... %.

(A) 0.0001 (B) 0.01 (C) 0.1 (D) 0.15

Ans. : b

NaX is a weak acid-strong base salt.

So, hydrolysis of NaX will be

$h = \sqrt{\frac{K_w}{K_a \times C}} = \sqrt{\frac{10^{-14}}{10^{-5} \times 0.1}} = 10^{-4}$

Therefore, % hydrolysis will be $= 10^{-4} \times 100 = 0.01\,\%$

39. What will be the pH of an aqueous solution of $1.0\,M$ ammonium formate ? (Given: $pK_a = 3.8$ and $pK_b = 4.8$)

(A) 7.5 (B) 3.4 (C) 6.5 (D) 10.2

Ans. : c

$pH = 7 + \frac{1}{2}\,pK_a - \frac{1}{2}\,pK_b$

40. Which is the correct representation of the solubility product constant of Ag_2CrO_4

(A) $[Ag^+]^2\,[CrO_4^{-2}]$ (B) $[Ag^+]\,[CrO_4^{-2}]$

(C) $[2Ag^+]\,[CrO_4^{-2}]$ (D) $[2Ag^+]^2\,[CrO_4^{-2}]$

Ans. : (a) $Ag_2CrO_4 \rightleftharpoons [2Ag^+] + [CrO_4^{-\,-}]$

Hence $K_{sp} = [Ag^+]^2\,[CrO_4^{-\,-}]$

41. If K_{sp} for $HgSO_4$ is 6.4×10^{-5}, then solubility of the salt is

(A) 8×10^{-3} (B) 8×10^{-6} (C) 6.4×10^{-5} (D) 6.4×10^{-3}

Ans. : (a) $HgSO_4$ of $K_{sp} = S^2$

$S = \sqrt{K_{sp}}$; $S = \sqrt{6.4 \times 10^{-5}}$; $S = 8 \times 10^{-3}\,m/l$.

42. The solubility of $PbCl_2$ is

(A) $\sqrt{K_{sp}}$ (B) $\sqrt[3]{K_{sp}}$ (C) $\sqrt[3]{\frac{K_{sp}}{4}}$ (D) $\sqrt{8K_{sp}}$

Ans. : (c) $PbCl_2 \rightleftharpoons \underset{S}{Pb^{2+}} + \underset{S}{2Cl^-}$

K_{sp} of $PbCl_2 = [Pb^{2+}] \times [Cl^-]^2$;$K_{sp} = S \times (2S)^2$

$K_{sp} = S \times 4S^2 = 4S^3$; $S^3 = \frac{K_{sp}}{4}$; $S = \sqrt[3]{\frac{K_{sp}}{4}}$

43. If solubility of calcium hydroxide is $\sqrt{3}$, then its solubility product will be

(A) 27 (B) 3 (C) 9 (D) $12\sqrt{3}$

Ans. : (d) $Ca(OH)_2 \rightleftharpoons \underset{(S)}{Ca^{++}} + \underset{(2S)^2}{2OH^-}$

$K_{sp} = 4S^3 = 4 \times \sqrt{3} \times \sqrt{3} \times \sqrt{3} = 12\sqrt{3}$

44. K_{sp} for sodium chloride is $36\ mol^2/litre^2$. The solubility of sodium chloride is

(A) $\frac{1}{36}$ (B) $\frac{1}{6}$ (C) 6 (D) 3600

Ans. : (c) $NaCl \rightleftharpoons \underset{S}{Na^+} + \underset{S}{Cl^-}$

$K_{sp} = S^2$, $S = \sqrt{K_{sp}} = \sqrt{36} = 6$.

45. What is the minimum concentration of SO_4^{2-} required to precipitate $BaSO_4$ in a solution containing $1.0 \times 10^{-4}\ mol\ \ Ba^{2+}$ (K_{sp} for $BaSO_4$ is 4×10^{-10})

(A) $4 \times 10^{-10}\ M$ (B) $2 \times 10^{-7}\ M$ (C) $4 \times 10^{-6}\ M$ (D) $2 \times 10^{-3}\ M$

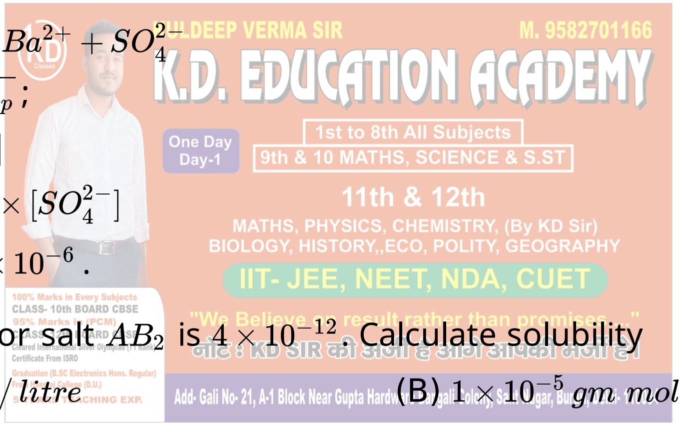
Ans. : (c) $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$

$K_{sp} = S^2 \Rightarrow S = \sqrt{K_{sp}}$;

$K_{sp} = [Ba^{2+}] \times [SO_4^{2-}]$

$4 \times 10^{-10} = [1 \times 10^{-4}] \times [SO_4^{2-}]$

$[SO_4^{2-}] = \frac{4\times10^{-10}}{1\times10^{-4}} = 4 \times 10^{-6}$.



46. Solubility product for salt AB_2 is 4×10^{-12}. Calculate solubility

(A) $1 \times 10^{-3}\ gm\ mol\ /\ litre$ (B) $1 \times 10^{-5}\ gm\ mol\ /\ litre$

(C) $1 \times 10^{-4}\ gm\ mol\ /\ litre$ (D) $1 \times 10^{-2}\ gm\ mol\ /\ litre$

Ans. : (c) $AB_2 \rightleftharpoons \underset{(S)}{A^{2+}} + \underset{(2S)^2}{2B^-}$

$K_{sp} = 4S^3$

$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4\times10^{-12}}{4}} = 1 \times 10^{-4}\ gm.\ mol/litre$

47. K_{sp} for $Cr(OH)_3$ is 2.7×10^{-31}. What is its solubility in moles / litre.

(A) 1×10^{-8} (B) 8×10^{-8} (C) 1.1×10^{-8} (D) 0.18×10^{-8}

Ans. : (a) $Cr(OH)_3 \rightarrow \underset{x}{Cr^{+3}} + \underset{3x}{3OH^-}$

$K_{sp} = x.(3x)^3 = 27x^4$

$x = \sqrt[4]{\frac{K_{sp}}{27}}$; $x = \sqrt[4]{\frac{2.7\times10^{-31}}{27}}$

$x = 1 \times 10^{-8}\ \ mole/litre.$

48. If the solubility product of lead iodide (PbI_2) is 3.2×10^{-8}, then its solubility in moles/litre will be

(A) 2×10^{-3} (B) 4×10^{-4} (C) 1.6×10^{-5} (D) 1.8×10^{-5}

Ans. : (a) $K_{sp} = 4S^3$

$4S^3 = 3.2 \times 10^{-8}$; $S = 2 \times 10^{-3} M$.

49. The values of K_{sp} for CuS, Ag_2S and HgS are 10^{-31}, 10^{42} and 10^{-54} respectively. The correct order of their solubility in water is

(A) $Ag_2S > HgS > CuS$ (B) $HgS > CuS > Ag_2S$

(C) $HgS > Ag_2S > CuS$ (D) $Ag_2S > CuS > HgS$

Ans. : (d) $Ag_2S > CuS > HgS$.

Solubility of $CuS = \sqrt{10^{-31}} = 3.16 \times 10^{-16}\ mol/lit.$

Solubility of Ag_2S

$= \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{10^{-42}}{4}} = 6.3 \times 10^{-5}\ mole/litre$

Solubility of $HgS = \sqrt{K_{sp}} = \sqrt{10^{-54}} = 10^{-27}\ mol/litre.$

50. The solubility product of As_2S_3 is 2.8×10^{-72}. What is the solubility of As_2S_3

(A) $1.09 \times 10^{-15}\ mole/litre$ (B) $1.72 \times 10^{-15}\ mole/litre$

(C) $2.3 \times 10^{-16}\ mole/litre$ (D) $1.65 \times 10^{-36}\ mole/litre$

Ans. : (a) $K_{sp} = [As^{3+}][S^{2-}]$, $S = \sqrt[5]{\frac{K_{sp}}{108}}$

$= \sqrt[5]{\frac{2.8 \times 10^{-72}}{108}} = 1.09 \times 10^{-15}$

51. If solubility product of $HgSO_4$ is 6.4×10^{-5}, then its solubility is

(A) $8 \times 10^{-3}\ mole/litre$ (B) $6.4 \times 10^{-5}\ mole/litre$

(C) $6.4 \times 10^{-3}\ mole/litre$ (D) $2.8 \times 10^{-6}\ mole/litre$

Ans. : (a) $HgSO_4 \rightarrow \underset{x}{Hg^{+2}} + \underset{x}{SO_4^{--}}$

$K_{sp} = x^2$; $x = \sqrt{K_{sp}}$; $x = \sqrt{6.4 \times 10^{-5}}$

$x = 8 \times 10^{-3}\ mole/litre.$

52. Solubility product (K_{sp}) of salts of type MX, MX_2, M_3X at temperature T are 4×10^{-8}, 3.2×10^{-14} and 2.7×10^{-15} respectively. Solubility of the salts at temperature T are in the order

(A) $MX > MX_2 > M_3X$

(B) $M_3X > MX_2 > MX$

(C) $MX_2 > M_3X > MX$

(D) $MX > M_3X > MX_2$

Ans. : d

$S_1 = 2 \times 10^{-4}$

$S_2 = \left(\frac{32}{4} \times 10^{-15}\right)^{1/3} = 2 \times 10^{-5}$

$S_3 = \left(\frac{27\times10^{-16}}{27}\right)^{19} = 10^{-4}$

$S_2 < S_3 < S_1$

53. MX is a salt formed by neutralisation of strong base, MOH, and weak acid , HX. If the dissociation constant of HX is K_a and solubility product of MX is k_{sp} , then the solubility of MX in aqueous acidic solution may be given as

(A) $\sqrt{k_{sp}}$ (B) $\sqrt{k_{sp}.k_a}$ (C) $\sqrt{k_{sp}.\left(1+\frac{[H^+]}{K_a}\right)}$ (D) $\sqrt{k_{sp}.\left(1+\frac{k_a}{[H^+]}\right)}$

Ans. : (C) $\sqrt{k_{sp}.\left(1+\frac{[H^+]}{K_a}\right)}$

54. A salt MX has $K_{sp} = 4 \times 10^{-10}$. What value of K_{sp} must another salt MX_3 have if the molar solubility of the two salts is to be identical

(A) 3.2×10^{-10} (B) 1.024×10^{-19} (C) 1.78×10^{-5} (D) 4.32×10^{-18}

Ans. : d

Solubility of MX $= \left(4 \times 10^{-10}\right)^{1/2}$

$MX_3 : K_{-p} = 27s^4$

$= 27\left(2 \times 10^{-5}\right)^4$

$= 4.32 \times 10^{-18}$

55. The solubility of $Ba_3(AsO_4)_2$ (molar mass $= 690$) is $6.9 \times 10^{-2}\, g/100\, mL$. What is the value of its K_{sp}?

(A) 1.08×10^{-11} (B) 1.08×10^{-13} (C) 1.0×10^{-15} (D) 6.0×10^{-13}

Ans. : b

$s = 6.9 \times 10^{-2}$ gm/100 ml

$= \frac{6.9\times10^{-2}}{690} = 10^{-4}$ mol/100 ml

$= 10^{-3}$ mol/lit

$K_{sp} = 108\, s^5$

$= 108 \times 10^{-15} = 1.08 \times 10^{-13}$

56. What is the molar solubility of :- $Ag_2CO_3 (K_{sp} = 4 \times 10^{-13})$ in $0.1\, M\, Na_2CO_3$ solution?

(A) 10^{-6} (B) 10^{-7} (C) 2×10^{-6} (D) 2×10^{-7}

Ans. : a

$Ag_2CO_3(s) \rightleftharpoons \underset{(2s)}{2Ag^+} + \underset{(s)}{CO_3^{-2}}$

$\mathrm{Na_2CO_3} \rightarrow \left[\mathrm{CO_3^{-2}}\right] = 0.1$

$\mathrm{K_{sp}} = \left[\mathrm{Ag^+}\right]^2 \times \left[\mathrm{CO_3^{-2}}\right]$

$4 \times 10^{-13} = (2s)^2 \times (s + 0.1)$

$\downarrow$

Neglected

$s = 10^{-6}$

57. K_{SP} of $Mg(OH)_2$ is 1×10^{-12}. $0.01\,M\,MgCl_2$ will show precipitation in a solution of pH reater than:-

(A) 3 (B) 9 (C) 6 (D) 8

Ans. : b

$\mathrm{K_{sp}} = \left[\mathrm{Mg^{2+}}\right][\mathrm{OH}]^2$

$1 \times 10^{-12} = 0.01[\mathrm{OH}]^2$

$1 \times 10^{-10} = \left[\mathrm{OH^-}\right]^2$

$1 \times 10^{-5} = [\mathrm{OH}]$

$5 = \mathrm{pH}$

$14 - 5 = pH$

$9 = pH$

58. Solubility product expression of salt MX_4 which is sparingly soluble with a solubility (S) can be given as

(A) $256S^5$ (B) $16S^3$ (C) $5S$ (D) $25S^4$

Ans. : a

$MX_4(s) \rightarrow M^{4+}_{(aq)} + 4X^-_{(aq)}$

$S(M) \quad 4S(M)$

$K_{SP} = S \times (4S)^4 = 256S^5$

59. Zirconium phosphate $[Zr_3(PO_4)_4]$ dissociates into three zirconium cations of charge $+4$ and four phosphate anions of charge -3. If molar solubility of zirconium phosphate is denoted by S and its solubility product by K_{sp} then which of the following relationship between S and K_{sp} is correct?

(A) $S = \{K_{sp}/144\}^{1/7}$

(B) $S = \left\{K_{sp}/(6912)^7\right\}$

(C) $S = (K_{sp}/6912)^{1/7}$

(D) $S = \{K_{sp}/6912\}^7$

Ans. : c

$(\mathrm{Zr_3(PO_4)_4}] \rightarrow 3\mathrm{Zr^{4+}} + 4\mathrm{PO_4^{3-}}$

$\mathrm{3S \quad 4S}$

$\mathrm{K_{sp}} = (4\mathrm{S})^4(3\mathrm{S})^3$

$S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$

60. The solubility product of $AgCl$ is 1.8×10^{-10}.Precipitation of $AgCl$ will occur by mixing which of the following solutions when mixed in equal volume ?

(A) $10^{-8}\, M\, Ag^{+}$ and $10^{-8}\, M\, Cl^{-}\, ions$

(B) $10^{-3}\, M\, Ag^{+}$ and $10^{-3}\, M\, Cl^{-}\, ions$

(C) $10^{-6}\, M\, Ag^{+}$ and $10^{-6}\, M\, Cl^{-}\, ions$

(D) $10^{-10}\, M\, Ag^{+}$ and $10^{-10}\, M\, Cl^{-}\, ions$

Ans. : b

$\text{I.P.} = [Ag^{+}][I^{-}]$

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

$= 0.25 \times 10^{-6}$

$= 2.5 \times 10^{-7}$

$IP > K_{SP}$ precipitation occur

61. Solid $Ba(NO_3)_2$ is gradually dissolved in $1.0 \times 10^{-4}\, M\, Na_2CO_3$ solution. At which concentration of Ba^{2+} the precipition will begin? ($K_{sp}\, BaCO_3 = 5.1 \times 10^{-9}$)

(A) $4.1 \times 10^{-5}\, M$ (B) $5.1 \times 10^{-5}\, M$ (C) $8.1 \times 10^{-8}\, M$ (D) $8.1 \times 10^{-7}\, M$

Ans. : b

$K_{sp} = [Ba^{2+}][CO_3^{-}]$

$5.1 \times 10^{-9} = [Ba^{2+}] \times 1 \times 10^{-4}$

$5.1 \times 10^{-5} = [Ba^{2+}]$

62. The expression for solubility product of mercurous iodide (Hg_2I_2) is

(A) $[2Hg^{+}]^2[I^{-}]^2$ (B) $[Hg^{+}]^2[I^{-}]^2$ (C) $[Hg^{+2}]^2[I^{-}]^2$ (D) $[Hg_2^{+2}][I^{-}]^2$

Ans. : d

Explanation :

Solubility product constant : It is defined as the product of the concentration of the ions present in a solution raised to the power by its stoichiometric coefficient in a solution of a salt. This takes place at equilibrium only. The solubility product constant is represented as, .

The dissociation of mercury(i) iodide is written as:

The expression for solubility constant for this reaction will be,

Therefore, the solubility product expression is

63. Calculate the minimum concentration of sulphate ion required to precipitate $BaSO_4$ in solution containing $10^{-4}\, mol/L$ of Ba^{+2} . $(K_{sp}\, BaSO_4 = 4 \times 10^{-10})$

(A) 4×10^{-6} (B) 4×10^{-10} (C) 2×10^{-6} (D) 2×10^{-10}

Ans. : a

$[Ba^{+2}][SO_4^{2-}] = 4 \times 10^{-10}$

for minimum or maximum

64. What is the molar solubility of $Fe(OH)_2$ $(K_{sp} = 8.0 \times 10^{-16})$ at pH 13.0 ?

(A) 8.0×10^{-18} (B) 8.0×10^{-15} (C) 8.0×10^{-17} (D) 8.0×10^{-14}

Ans.: (A) 8.0×10^{-18}

65. What is the molarity of F^- in a saturated solution of InF_3 ? $(K_{sp} = 7.9 \times 10^{-10})$

(A) 2.3×10^{-3} (B) 8.3×10^{-3} (C) 1.0×10^{-3} (D) 7.0×10^{-3}

Ans. : d

$\text{In} \quad F_3 \rightleftharpoons \text{In}^{3+} + 3F^-$

$t = 0 \quad C \quad 0 \quad 0$

$t = t \quad C - S \quad S \quad 3S$

$\therefore \quad K_{sp} = S \times (3S)^3$

$\therefore \quad 7.9 \times 10^{-10} = 27S^4$

$\therefore S^4 = 2.92 \times 10^{-11}$

$\therefore S = 2.32 \times 10^{-3}\, M$

where S is the concentration of F^- ion in terms of molarity.

66. The required amount of KBr (molar mass $= 119$) in gram to start the precipitation of $AgBr$ in $500\, mL$ solution of $0.05\, M\, AgNO_3$ will be $(K_{SP}$ of $AgBr = 5 \times 10^{-13})$

(A) $1.19 \times 10^{-9}\, g$ (B) $4 \times 10^{-11}\, g$ (C) $5.95 \times 10^{-10}\, g$ (D) $2.97 \times 10^{-10}\, g$

Ans. : (C) $5.95 \times 10^{-10}\, g$

67. Which of the following sets of concentrations will cause the precipitation of $ZnCl_2$ $(K_{sp} = 1.2 \times 10^{-12} M^3)$?

(A) $[Zn^{2+}] = 10^{-8}\, M; [Cl^-] = 10^{-8}\, M$ (B) $[Zn^{2+}] = 10^{-5}\, M; [Cl^-] = 10^{-4}\, M$

(C) $[Zn^{2+}] = 10^{-6}\, M; [Cl^-] = 10^{-5}\, M$ (D) $[Zn^{2+}] = 10^{-5}\, M; [Cl^-] = 10^{-3}\, M$

Ans. : d

For precipitation $[Zn^{2+}][Cl^-]^2 > K_{sp}$.

68. The solubility product of lead bromide is 8×10^{-5}. If the salt is 80% dissociated in saturated solution, find the solubility of the salt

(A) $1.7 \times 10^{-4}\, M$ (B) $2.3 \times 10^{-6}\, M$ (C) $1.8 \times 10^{-4}\, M$ (D) $3.4 \times 10^{-2}\, M$

Ans. : d

$PbBr_2(s) \rightleftharpoons \underset{0.8\, S}{Pb^{2+}} + \underset{2 \times 0.8\, S}{2Br^-}$

$\mathrm{K_{sp}} = [\mathrm{Pb^{2+}}][\mathrm{Br^-}]^2$

$\Rightarrow 8 \times 10^{-5} = (0.8\,\mathrm{S}) \times (2 \times 0.8\,\mathrm{S})^2 = 2.048\,\mathrm{S}^3$

$\Rightarrow \mathrm{S} = 3.4 \times 10^{-2}\,\mathrm{M}$

69. $150\,mL$ of $0.0008\,M$ ammonium sulphate is mixed with $50\,mL$ of $0.04\,M$ calcium nitrate. The ionic product of $CaSO_4$ will be ($K_{SP} = 2.4 \times 10^{-5}$ for $CaSO_4$)

(A) $< K_{SP}$ (B) $> K_{SP}$ (C) $\approx K_{SP}$ (D) None of these

Ans. : a

For $[\mathrm{Ca^{+2}}]$, $\mathrm{M_1V_1 = M_2V_2}$

$0.04 \times 50 = M_2 \times 200$

$M_2 = 0.01$

Similarly, for $[\mathrm{SO_4^{-2}}]$

$0.0008 \times 150 = \mathrm{M_2} \times 200$

$\mathrm{M_2} = 0.0006$

ionic product (Q) $= [\mathrm{Ca^{+2}}] \times [\mathrm{SO_4^{-2}}]$

$= 6 \times 10^{-6}$

So, $Q < K_{\mathrm{SP}}$

70. The solubility product of $Mg(OH_2)$ is 1×10^{-12}. At what pH precipitation will occur in $0.01\,M$ $MgCl_2$ solution?

(A) 8 (B) 9 (C) 10 (D) 12

Ans. : b

The solubility product expression is:

$\mathrm{K_{SP}} = [\mathrm{Mg^{2+}}][\mathrm{OH^-}]^2$

Rearrange the above expression:

$[\mathrm{OH^-}] = \sqrt{\frac{\mathrm{K_{sp}}}{[\mathrm{Mg^{2+}}]}}$

Substitute values in the above expression and calculate hydroxide ion concentration-

$[\mathrm{OH^-}] = \sqrt{\frac{1 \times 10^{-12}}{0.01}}$

$[\mathrm{OH^-}] = 1 \times 10^{-5}\mathrm{M}$

Calculate pOH from the hydroxide ion concentration:

$\mathrm{pOH} = -\log[\mathrm{OH^-}] = -\log(1 \times 10^{-5})$

$\mathrm{pOH} = 5$

Calculate pH from pOH-

$\mathrm{pH} = 14 - \mathrm{pOH}$

$\mathrm{pH} = 14 - 5$

$\mathrm{pH} = 9$

71. The solubility product of AB_2 is $8 \times 10^{-5}\, M^3$. If the salt is $80\,\%$ dissociated in saturated solution, find the solubility of the salt.
Given [Molar mass $(AB_2) = 360$]

(A) 0.050 (B) 0.065 (C) 0.034 (D) 0.074

Ans. : d

$$AB_2 \xleftrightarrow{80\%} \underset{0.8S}{A^{+2}} + \underset{2\times 0.8S}{2B^-}$$

$$\mathrm{K_{sp}} = (0.8 \times \mathrm{S})(2 \times 0.8\mathrm{S})^2$$

$$8 \times 10^{-5} = 2.048\mathrm{S}^3 \left(= 2\mathrm{S}^3\right)$$

$$\mathrm{S}^3 = \frac{8\times 10^{-5}}{2}$$

$$\mathrm{S}^3 = 40 \times 10^{-6}$$

$$S = \sqrt[3]{40 \times 10^{-6}}$$

$$= 3.4 \times 10^{-2}\,\mathrm{M}$$

$$= 0.034\,\mathrm{M}$$

72. The molar solubility of PbI_2 in $0.2\,M$ $Pb(NO_3)_2$ solution in terms of solubility product of PbI_2 is

(A) $\sqrt{(K_{sp}/0.2)}$ (B) $\sqrt{K_{sp}/0.4}$ (C) $\sqrt[3]{K_{sp}/0.8}$ (D) $\sqrt{K_{sp}/0.8}$

Ans. : d

The solution contains:

$$PbI_2 \implies Pb^{+2} + 2I^-$$

solubitity S $\quad S \quad 2S$

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

$0.2 \quad 0.2 \quad 2 \times 0.2$

$$K_{sp} = \left[pb^{2+}\right]\left[I^-\right]^2$$

$$K_{sp} = (S + 0.2)(2S)^2$$

$$S << 2$$

$$K_{sp} = 0.2 \times 4S^2$$

$$K_{sp} = 0.8S^2$$

$$S = \sqrt{\frac{Ksp}{0.8}}$$

$$= (K_{sp}/0.8)^{1/2}$$

73. The required amount of KBr (molar mass $= 119$) in gram to start the precipitation of $AgBr$ in $500\,mL$ solution of $0.05\,M$ $AgNO_3$ will be :- (K_{SP} of $AgBr = 5 \times 10^{-13}$)

(A) $1.19 \times 10^{-9}\,g$ (B) $4 \times 10^{-11}\,g$ (C) $5.95 \times 10^{-10}\,g$ (D) $2.97 \times 10^{-10}\,g$

Ans. : (C) $5.95 \times 10^{-10}\,g$

74. In an aqueous solution SCN^{-1}, B^{-1}, I^{-1} and Cl^{-} are present. Which will get precipitated first, when $AgNO_3$ is mixed with each of them? Given that

Ksp of $AgCl = 1.2 \times 10^{-10}$,

Ksp of $AgI = 1.7 \times 10^{-16}$

Ksp of $AgSCN = 7.1 \times 10^{-7}$,

Ksp of $AgBr = 3.5 \times 10^{-13}$

(A) I^{-} (B) Cl^{-} (C) Br^{-} (D) SCN^{-}

Ans.: (A) I^{-}

75. Solubility of $AgCl$ in $0.2\, M\, NaCl$ is x and that in $0.1\, M\, AgNO_3$ is y then which of the following is correct ?

(A) $x = y$ (B) $x > y$

(C) $x < y$ (D) we can not predict

Ans. : c

In $0.2\, M\, NaCl, [Cl^{-}] = 0.2$

$(K_{sp})_{AgCl} = [Ag^{+}][Cl^{-}]$

$\therefore (K_{sp})_{AgCl} = x \times 0.2$

$\Rightarrow x = \frac{(K_{sp})_{AgCl}}{0.2} = 5(K_{sp})_{AgC1}$

In $0.1\, M\, AgNO3, [Ag^{+}] = 0.1\, M$

$\Rightarrow (K_{sp})_{AgCl} = 0.1 \times y$

$\Rightarrow y = \frac{(K_{sp})_{AgCl}}{0.1} = 10(K_{sp})_{AgCl}$

$\therefore x < y$

76. Solubility of an MX_2 type electrolyte is $0.5 \times 10^{-4}\, mol/litre$, then K_{sp} of electrolyte is

(A) 5×10^{-12} (B) 2.5×10^{-10} (C) 1×10^{-13} (D) 5×10^{-13}

Ans. : d

An electrolyte MX_2 undergoes dissociation as follows :-

$MX_2 \rightleftharpoons M^{+2} + 2X^{-}$

Concentration	MX_2	M^{+2}	X^{-}
Initial concentration	1	0	0
Concentration at Equilibrium	$1-s$	s	$2s$

Thus from the above condition we can say that,

$K_{sp} = s \times (2s)^2 = 4 \times (s)^3$

Here, s (the solubility) is $0.5 \times 10^{-4}\, mole/lit.$

$\therefore K_{\mathrm{sp}} = 4 \times \left(0.5 \times 10^{-4}\right)^3$

$\therefore K_{sp} = 5 \times 10^{-13}$

77. What is the pH of a saturated solution of $Cu(OH)_2$? $(K_{sp} = 4.0 \times 10^{-6}\ )$

(A) 6.1 (B) 12.3 (C) 8.42 (D) 11.8

Ans. : a

$$S = \left[\frac{k_{sp}}{4}\right]^{\frac{1}{3}}, \therefore \{\mathrm{OH}^-\} = 2\mathrm{S} = 2 \times 10^{-2}$$

$$\therefore \mathrm{pOH} = -\log 2 \times 10^{-2} = 1.7$$

$$\therefore \mathrm{pH} = 14 - 1.7 = 12.3$$

78. What will be the solubility of $AlCl_3$ in solution of $CaCl_2$ with concentration C

(A) $\frac{K_{sp}}{2C}$ (B) $\frac{K_{sp}}{8C^3}$ (C) $\frac{K_{sp}}{3C}$ (D) $\frac{K_{sp}}{4C^2}$

Ans. : (B) $\frac{K_{sp}}{8C^3}$

79. What is the molar solubility of $Ag_2CO_3\,(K_{sp} = 4 \times 10^{-13})$ in $0.1\,M\ Na_2CO_3$ solution ?

(A) 10^{-6} (B) 10^{-7} (C) 2×10^{-6} (D) 2×10^{-7}

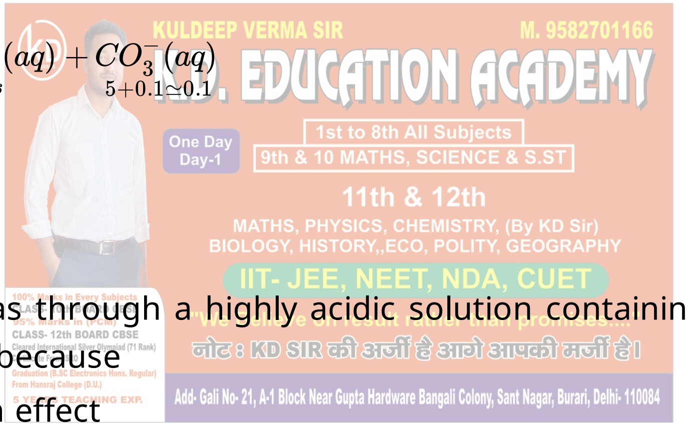
Ans. : a

$$Ag_2CO_3(s) \xleftrightarrow{K_q} \underset{2s}{2Ag^+(aq)} + \underset{5+0.1\simeq 0.1}{CO_3^-(aq)}$$

$$\mathrm{K_{sp}} = 4\mathrm{s}^2 \times 0.1$$

$$4 \times 10^{-13} = 45^2 \times 0.1$$

$$\mathrm{s} = 1 \times 10^{-6}$$



80. On passing H_2S gas through a highly acidic solution containing Cd^{2+} ions, CdS is not precipitated because

(A) Of common ion effect

(B) The solubility of CdS is low

(C) Cd^{2+}ions do not form complex with H_2S

(D) The solubility product of CdS is low

Ans. : a

It's Obvious.

81. When NH_4Cl is added to NH_4OH solution, the dissociation of ammonium hydroxide is reduced. It is due to

(A) Common ion effect (B) Hydrolysis

(C) Oxidation (D) Reduction

Ans. : (a)Due to common ion effect.

82. In the reaction: $H_2S \rightleftharpoons 2H^+ + S^{2-}$, when NH_4OH is added, then

(A) S^{2-} is precipitate

(B) No action takes places

(C) Concentration of S^{2-} decreases

(D) Concentration of S^{2-} increases

Ans. : (d) In IV^{th} group the S^{2-} concentration increase when added the NH_4OH because

$NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$H_2S \rightleftharpoons 2H^+ + S^{2-}$

$OH^- + H^+ \rightleftharpoons H_2O$. So that S^{2-} is increased.

83. Addition of conc. HCl to saturated $BaCl_2$ solution precipitates $BaCl_2$; because

(A) It follows from Le Chatelier's principle

(B) Of common-ion effect

(C) Ionic product (Ba^{++}),(Cl^-)remains constant in a saturated solution

(D) At constant temperature, the product (Ba^{2+}), $(Cl^-)^2$ remains constant in a saturated solution

Ans. : c

When you add a compound having a common ion, It decreases the solubility of the solute. This is because Le Chatelier's principle states the reaction will shift toward the left (toward the reactants) to relieve the stress of the excess product.

This is called common-ion effect. due to this effect, some solute becomes insoluble and precipitate.

In the above case HCl and $BaCl_2$ have Cl^- common ions and thus $BaCl_2$ precipitates.

84. The solubility of $AgCl$ is minimum in

(A) $AgNO_3\,(0.1\,M)$ (B) $H_2O\,(l)$ (C) $NaCl\,(0.4\,M)$ (D) $BaCl_2\,(0.3\,M)$

Ans. : b

Due to common ion /rea^n will proceed in backward direction that why solubility is decreases in the presence of common ion.

Concentration of common ion increases, solubility decreases. Max^n common ion is obtained from $BaCl_2$

$\underset{0.3}{BaCl_2} \rightarrow Ba^{2+} + \underset{0.6\,M}{2Cl^-}$

$\underset{0.1}{AgNO_3} \rightarrow Ag^+ + \underset{0.1}{NO_3^-}$

$\underset{0.4}{NaCl} \rightarrow Na^+ + \underset{0.4}{Cl^-}$

85. The pH of a simple sodium acetate buffer is given by $pH = pK_a + \log\frac{[Salt]}{[Acid]}$ K_a of acetic acid $= 1.8 \times 10^{-5}$ If [Salt] = [Acid] = $0.1\,M$, the pH of the solution would be about

(A) 7 (B) 4.7 (C) 5.3 (D) 1.4

Ans. : (b) [Salt]= $0.1\,M$, [Acid]= $0.1\,M$

$K_a = 1.8 \times 10^{-5}$; $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log 1.8 \times 10^{-5} + \log \frac{0.1}{0.1} = -\log\ 1.8 \times 10^{-5}$

$pH = 4.7.$

86. ......m sodium acetate should be added to a $0.1\ m$ solution of CH_3COOH to give a solution of $pH = 5.5\,(pK_a$ of $CH_3COOH = 4.5)$

(A) 0.1 (B) 0.2 (C) 1 (D) 10

Ans. : (c) $pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$; $5.5 = 4.5 + \log \frac{[\text{Salt}]}{[0.1]}$

$\log \frac{[\text{Salt}]}{0.1} = 5.5 - 4.5 = 1$

$\frac{[\text{Salt}]}{0.1} = antilog\,1 = 10$; $[\text{Salt}] = 1$

87. Which of the following is a buffer

(A) $NaOH + CH_3COONa$ (B) $NaOH + Na_2SO_4$

(C) $K_2SO_4 + H_2SO_4$ (D) $NH_4OH + CH_3COONH_4$

Ans. : (d) Buffer is mixture of weak base and its acid salt.

88. When a buffer solution of sodium acetate and acetic acid is diluted with water

(A) Acetate ion concentration increases

(B) H^+ion concentration increases

(C) OH^-ion concentration increases

(D) H^+ion concentration remain unaltered

Ans. : (d)Buffer solution have constant pH. When we add the water into this buffer solution. So no effect on it.

89. Which of the following solutions cannot act as a buffer

(A) $NaH_2PO_4 + H_3PO_4$ (B) $CH_3COOH + CH_3COONa$

(C) $HCl + NH_4Cl$ (D) $H_3PO_4 + Na_2HPO_4$

Ans. : (c)A strong acid is not used to make a buffer.

90. A solution which is resistant to change of pH upon the addition of an acid or a base is known as

(A) A colloid (B) A crystalloid (C) A buffer (D) An indicator

Ans. : (c)Because buffer solution have a constant pH.

91. A certain buffer solution contains equal concentration of X^- and HX. The K_a for HX is 10^{-8}. The pH of the buffer is

(A) 3 (B) 8 (C) 11 (D) 14

Ans. : (b)The equal conc. of salt and acid.

92. The pK_a of equimolecular sodium acetate and acetic acid mixture is 4.74. If pH is

(A) 7 (B) 9.2 (C) 4.74 (D) 14

Ans. : (c) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$ equimolar means $\frac{[\text{Salt}]}{[\text{Acid}]} = 1$; $pH = 4.74 + 0 = 4.74$

93. The condition for minimum change in pH for a buffer solution is

(A) Isoelectronic species are added

(B) Conjugate acid or base is added

(C) $pH = pK_a$

(D) None of these

Ans. : c

Buffers work best when $pH = pK_a$

From the Henderson-Hasselbalch equation

$pH = pK_a + \log 10\frac{[A-]}{[HA]}$

If $\text{pH} = pK_a \implies \log 10\frac{[A-]}{[HA]} = 0 \implies \frac{[A-]}{[HA]} = 1$

buffer exhibits the highest resistance to acid and base addition for the equimolar solution (when $pH = pKa$). From the plot, it is also obvious that buffer capacity has reasonably high values only for pH close to pKa value. The further from the optimal value, the lower buffer capacity of the solution

94. The ionization constant of a certain weak acid is 4. What should be the [salt] to [acid] ratio if we have to prepare a buffer with $pH = 5$ using this acid and one of the salts

(A) $1:10$ (B) $10:1$ (C) $5:4$ (D) $4:5$

Ans. : (b) $pH = -\log K_b + \log\frac{[\text{salt}]}{[\text{acid}]}$

$5 = -\log 10^{-4} + \log\frac{[\text{salt}]}{[\text{acid}]}$

$\log\frac{[\text{salt}]}{[\text{acid}]} = 1$

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

95. $50\ ml$ of $2\,N$ acetic acid mixed with $10\,ml$ of $1\ N$ sodium acetate solution will have an approximate pH of

(A) 4 (B) 5 (C) 6 (D) 7

Ans. : (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]} = -\log 2\times10^{-5} + \log\frac{10\times1}{50\times2} = 4$.

96. Which of the following will not function as a buffer solution

(A) $NaCl$ and $NaOH$

(B) $NaOH$ and NH_4OH

(C) CH_3COONH_4 and HCl

(D) All of above

Ans. : (d) Because buffer solution are mixture of weak acid or weak base and their salt.

97. pH of $0.1\,M$ solution of a weak acid (HA) is 4.50. It is neutralised with $NaOH$ solution to decrease the acid content to half pH of the resulting solution

(A) 4.50 (B) 8 (C) 7 (D) 10

Ans. : (b)$pH = pK_a + \log \frac{[Salt]}{[Acid]}$

$pH = pK_a$

$K_a = 0.1 \times (10^{-3.5})^2 = 0.1 \times 10^{-7} = 10^{-8} \Rightarrow pH = 8$

98. If $50\,ml$ of $0.2\,M\,KOH$ is added to $40\,ml$ of $0.5\,M\,HCOOH$, the pH of the resulting solution is $(K_a = 1.8 \times 10^{-4})$

(A) 3.4 (B) 7.5 (C) 5.6 (D) 3.75

Ans. : (a) $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$[\text{Salt}] = \frac{0.2 \times 50}{1000} = 0.01$; $[\text{Acid}] = \frac{0.5 \times 40}{1000} = 0.02$

$pH = -\log\,(1.8 \times 10^{-4}) + \log \frac{0.01}{0.02}$

$pH = 4 - \log\,(1.8) + \log 0.5$

$pH = 4 - \log\,(1.8) - 0.301$

$pH = 3.4$

99. The pH of a buffer solution containg 0.2 mole per litre CH_3COONa and 1.5 mole per litre CH_3COOH is (Ka for acetic acid is 1.8×10^{-5})

(A) 4.87 (B) 5.8 (C) 2.4 (D) 9.2

Ans. : (a) $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log\,[1.8 \times 10^{-5}] + \log \frac{0.2}{0.1} = 4.87$

100. For preparing a buffer solution of $pH\,5$ by mixing sodium acetate and acetic acid, the ratio of the concentration of salt and acid should be $(K_a = 10^{-5})$:-

(A) $1:10$ (B) $1:1$ (C) $10:1$ (D) $1:100$

Ans. : (B) $1:1$

101. What is the pH of solution obtained by mixing $5.076\,gm$ of methyl ammonium nitrate $(CH_3NH_3NO_3)$ to $120\,ml$, $0.225\,M$ methylamine (CH_3NH_2 ; $K_b = 4 \times 10^{-4}$).

(A) 3.7 (B) 4.3 (C) 10.3 (D) 11

Ans. : c

m. moles of salt $= \frac{5.076}{94} \times 1000 = 54$

m. moles of weak base $= 120 \times 0.225 = 27$

$\text{pOH} = \text{pK}_\text{b} + \log \frac{[\text{Salt}]}{[\text{Base}]} \Rightarrow 4 - \log 4 + \log \frac{54}{27}$

$\text{pOH} = 3.7$

$\mathrm{pH} = 10.3$

102. A buffer that is a mixture of acetic acid $(K_a = 2 \times 10^{-5})$ and potassium acetate has $pH = 5.18$. The $\frac{[CH_3COO^-]}{[CH_3COOH]}$ ratio in this buffer is approx

(A) $1:1$ (B) $3:1$ (C) $5:1$ (D) $1:3$

Ans. : b

$\mathrm{pH} = \mathrm{pKa} + \log \frac{[\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]}$

$5.18 = (5 - \log 2) + \log \frac{[\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]}$

$5.18 - 4.7 = 0.48 = \log 3 = \log \frac{[\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]}$

$\therefore \frac{[\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]} = 3:1$

103. Which buffer is present in human blood :-

(A) $NaH_2PO_4 + Na_2HPO_4$ (B) $H_3PO_4 + NaH_2PO_4$

(C) $CH_3COOH + CH_3COONa$ (D) $H_2CO_3 + HCO_3^-$

Ans. : d

H_2CO_3/HCO_3^-

104. Which of the following is a buffer solution ?

(A) $500\,mL$ of $0.1\,N\,CH_3COOH + 500\,mL$ of $0.1\,N\,NaOH$

(B) $500\,mL$ of $0.1\,N\,CH_3COOH + 1000\,ml$ of $0.1\,N\,NaOH$

(C) $500\,ml$ of $0.1\,N\,CH_3COOH + 500\,mL$ of $0.2\,N\,NaOH$

(D) $500\,mL$ of $0.2\,N\,CH_3COOH + 500\,ml$ of $0.1\,N\,NaOH$

Ans. : d

Buffer solution contains both weak acid and its salt

105. The pK_a of weak acid (HA) is 4.5 . The pOH of an aqueous buffered solution of HA in which 50% of the acid is neutralised in its titration with strong base (BOH) is

(A) 7 (B) 4.5 (C) 2.5 (D) 9.5

Ans. : d

$\mathrm{pH} = \mathrm{pK_a} + \log \frac{[\mathrm{A^-}]}{[\mathrm{HA}]}$

$= 4.5 + \log \frac{50}{50}$

$= 4.5$

$\mathrm{pOH} = 14 - 4.5 = 9.5$

106. Calculate pH of solution obtained by mixing $50\,ml$ $0.2\,M$ NH_4Cl solution and $75\,ml$. $0.1\,M\,NaOH$ solution. pKb for aqueous NH_3 is 4.74

(A) 4.26 (B) 5.22 (C) 8.78 (D) 9.74

Ans. : d

Given that :-

$NH_4Cl = 50\,ml$ and $0.2\,M$

$NaOH = 75\,ml$ and $0.1\,M$

$$NH_4Cl + NaOH \longrightarrow NH_4OH + NaCl$$

Initial	0.2×50	0.1×75		
	$= 10$	$= 7.5$		
Final	$= 2.5$	$= 0$	$\longrightarrow$	7.5
	(Salt)			(Weak Base)

$pOH = pK_b + \log \frac{[\,\text{salt}\,]}{[\,\text{Base}\,]}$

$= 4.74 + \log \frac{[2.5]}{[7.5]}$ $\quad [\because pK_b$ of $NH_{3(aq)} = 4.74]$

$= 4.74 + \log \frac{1}{3}$

$= 4.74 + \log 3$

$= 4.74 - 0.47$

$pOH = 4.27$

$pH = 14 - pOH$

$14 - 4.27$

$pH = 9.73$

107. In which of the following combinations, is buffer action expected

1. $NH_3 + NH_4Cl$
2. $HCl + NaCl$
3. $NH_3 + HCl$ in $2:1$ $mole$ ratio

Select the correct answer using the code given below

(A) 1 and 2 (B) 1 and 3 (C) 2 and 3 (D) $1,2$ and 3

Ans. : (B) 1 and 3

108. The pK_a of weak acid HA is 4.5. The pOH of an aqueous buffered solution of HA in which 50% of the HA acid is ionised is

(A) 4.5 (B) 2.5 (C) 9.5 (D) 7

Ans. : c

$pH = pK_a + \log \frac{[\text{Conjugate base}\,]}{[\,\text{acid}\,]}$

$pKa = 4.5$

[Conjugate base $]=[$ acid], as $50\,\%$ ionification

$pH = 4.5 + \log \frac{0.5}{0.5}$

$pH = 4.5$

Then, $pOH = 14 - 4.5 = 9.5$

109. $20\,ml$ solution contains $0.1\,M, NH_4Cl$ and $0.01\,M\ NH_4OH$. By adding which one its pH will not change ?

(A) Addition of $1\,ml$ water

(B) Addition of $5\,ml$, $0.1\,M\,NH_4Cl$

(C) Addition of $5\,ml$, $0.1\,M\,NH_4OH$

(D) Addition of $10\,ml$, $0.1\,M\,NH_4Cl$

Ans. : a

By addition of small amount of water pH of buffer solution does not change

110. Which of the following mixture of solution can function as a buffer solution ?

(A) $50\,ml$ of $0.1\,M\,CH_3COOH + 50\,ml$ of $0.1\,M\,NaOH$

(B) $50\,ml$ of $0.2\,M\,HCl + 50\,ml$ of $0.2\,M\,NaOH$

(C) $50\,ml$ of $0.2\,M\,NH_3 + 50\,ml$ of $0.2\,M\,HCl$

(D) $50\,ml$ of $0.2\,M\,NH_3 + 50\,ml$ of $0.1\,M\,HCl$

Ans. : d

It will behave as a mixture of NH_4^+ and NH_3, because some of the NH_3 would react with HCl to form NH_4^+.

111. The pH of an acidic buffer mixture is

(A) > 7 (B) < 7

(C) $= 7$ (D) Depends upon K_a of acid

Ans. : d

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

112. In which of the following combination, is buffer action expected?

$(A)\ NH_4OH + NH_4Cl$ $(B)\ HCl + NaCl$

$(C)\ NH_4OH + HCl$ in $2:1\,mole$ ratio

Select the correct answer using the code given below

(A) A and B (B) A and C (C) A, B and C (D) None of these

Ans. : b

Combination of weak base and its salt with strong acid act as basis bufter solution

$$NH_4OH + HCl \rightarrow NH_4Cl + H_2O$$

inintial	2 mol	1 mol	0	0
eqn	$2-1=1$	$1-1=0$	1	1

Final solution have $\mathrm{NH_4OH + NH_4Cl}$ then its buffer solution

113. If $20\,mL$ of $0.1\,M\,NaOH$ is added to $30\,mL$ of $0.2\,M\,CH_3COOH$ $(pK_a = 4.74)$, the pH of the resulting solution is

(A) 4.44 (B) 9.56 (C) 8.96 (D) 9.26

Ans. : a

$$CH_3COOH + NaOH \rightarrow CH_3COONA + H_2O$$

Initial milli moles	6	2	$-$	$-$
after reaction	4	$-$	2	$-$

$\mathrm{pH} = \mathrm{pK_a} + \log \frac{[\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]} = 4.44$

114. Which of the following is a buffer solution?

(A) $100\,ml\,0.1\,M\,CH_3COOH + 100\,ml, 0.05\,M NaOH$

(B) $200\,ml\,0.1\,M\,NH_4OH + 200\,ml, 0.08\,M HCl$

(C) $300\,ml\,0.1\,M\,NaOH + 500\,ml, 0.1\,M C_6H_5COOH$

(D) All of these

Ans. : (D) All of these

115. A certain buffer solution contains equal concentration of X^- and HX. The K_b for X^- is 10^{-10}. The pH of the buffer is

(A) 4 (B) 7 (C) 10 (D) 14

Ans. : a

$pH = pKa + \log \frac{[X^-]}{[HX]}$

$K_a \times K_b = K_w$

116. Which is a buffer solution

(A) $CH_3COOH + CH_3COONa$ (B) $CH_3COOH + CH_3COONH_4$

(C) $CH_3COOH + NH_4Cl$ (D) $NaOH + NaCl$

Ans. : a

($WA+$ salt of the wA with SB) $=$ Acidic buffer

117. Phenolphthalein does not act as an indicator for the titration between

(A) $NaOH$ and CH_3COOH

(B) $H_2C_2O_4$ and $KMnO_4$

(C) $Ba(OH)_2$ and HCl

(D) KOH and H_2SO_4

Ans. : b

It's Obvious.

118. If very small amount of HPh (phenolphthalein) added to $0.1\,M$ aqueous solution of CH_3COONa. The fraction of indicator that exists in colourless form is Given $[K_b\,(CH_3COO^-) = 10^{-9}, pK_a\,(HPh) = 9.6, log\,2 = 0.3]$

(A) 0.2 (B) 0.8 (C) 0.5 (D) 0.4

Ans. : b

$\mathrm{pH} = \frac{1}{2}(\mathrm{pKw} + \mathrm{pKa} + \log \mathrm{C})$

$= \frac{1}{2}(14 + 5 - 1) = 9$

$\mathrm{pH} = \mathrm{pKa} + \log \frac{\mathrm{In^-}}{\mathrm{HIn}}$

$\log \frac{\mathrm{HIn}}{\mathrm{In^-}} = 0.6 = 2 \times 0.3 = 2 \log 2$

$\frac{\mathrm{HIn}}{\mathrm{In}^-} = 4, \mathrm{HIn} + \mathrm{In}^- = 1.0, \mathrm{HIn} = 0.80$

119. A $20\,ml$ sample of a $0.1\,M$ solution of the weak monoprotic acid HA is titrated with $0.25\,M\,NaOH$, volume of $NaOH$ solution required till equivalent point is :- .... mL

(A) 10 (B) 6 (C) 12 (D) 8

Ans. : d

mili gram equivalent of HA = mili gram equivalent of $NaOH$

$0.1\,\mathrm{N} \times 20\,\mathrm{mL} = 0.25\,\mathrm{N} \times \mathrm{V}$

120. CrO_4^{-2} (Yellow) changes to $Cr_2O_7^{-2}$ (orange) in $pH = x$ and vice verses in $pH = y$, hence x and y are respectively

(A) $6,5$ (B) $8,6$ (C) $6,8$ (D) $7,7$

Ans. : (C) $6,8$

121. The solubility product of CuS, Ag_2S, HgS are 10^{-31}, 10^{-44}, 10^{-54} respectively. The solubilities of these sulphides are in the order

(A) $Ag_2S > CuS > HgS$ (B) $Ag_2S > HgS > Cus$

(C) $HgS > Ag_2S > Cus$ (D) $CuS > Ag_2S > HgS$

Ans. : (a) $Ag_2S > CuS > HgS$

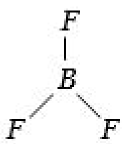
122. Boron halides behave as Lewis acids, because of their

(A) Ionic nature (B) Acidic nature

(C) Covalent nature (D) Electron deficient nature

Ans. : (d) Boron halides behave as Lewis acid because of their electron deficient nature eg., as

(Deficiency of two electron for inert configuration)



123. pH value of $N/10\ NaOH$ solution is

(A) 10 (B) 11 (C) 12 (D) 13

Ans. : (d) In $\frac{N}{10}\,NaOH$ have $[OH^-] = 10^{-1}M$ means $pOH = 1$ and then $pH + pOH = 14$
$pH = 14 - pOH = 13$.

124. The pH of blood does not appreciably change by a small addition of an acid or a base because blood

(A) Contains serum protein which acts as buffer

(B) Contains iron as a part of the molecule

(C) Can be easily coagulated

(D) It is body fluid

Ans. : (a) pH of blood does not change because it is a buffer solution.

125. Consider the following reaction involving two acids shown below : formic acid and.HF.

Which of the following statements about this reaction are true ?

(A) Formic acid is the strongest Bronsted acid in the reaction

(B) HF is the strongest Bronsted acid in the reaction

(C) KF is the strongest Bronsted base in the reaction

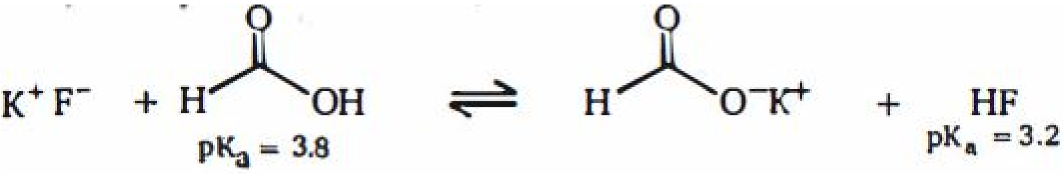
(D) KO_2CH is the strongest Bronsted base in the reaction

(E) The equilibrium favours the reactants

(F) The equilibrium favours the products

(G) Formic acid has a weaker conjugate base

(H) HF has a weaker conjugate base



(A) A, D and F (B) $B, D,$ and H

(C) $A, C,$ and H (D) B, D, E and H

Ans. : (d) Acid base reactions favours equilibrium towards weak acid & weak base.

----- "काक चेष्टा बको ध्यानं, श्वान निद्रा तथैव च। अल्पाहारी गृहत्यागी, विद्यार्थी पंच लक्षणं।" -- कौवे की तरह प्रयास, बगुले की तरह ध्यान, कुत्ते की तरह नींद, कम भोजन, और घर त्याग, ये विद्यार्थियों के पांच लक्षण हैं। -----