Ionic Equilibrium

1. The pH of a solution obtained by dissolving 5 × 10−4 moles of $Ca(OH)2$ (strong electrolyte) to 100 ml solution at 298 K will be

(a) 11 (b) 12

(c) 9.8 (d) 2

Answer: (b)

$[Ca(OH)]_2 = \frac{5\times10-4\times1000}{100} = 5\times10-3$ mole L^{-1}

$[OH-] = 2 \times 5 \times 10-3 = 10-2$ mole L−1

$[H+] = \frac{10-14}{10-2} = 10^{-12}$

pH = −log 10−12 = 12

2. One litre of a saturated solution of $CaCO3$ is evaporated to dryness due to which 7.0 mg of residue is left. The solubility product for $CaCO3$ is

(a) 4.9 × 10−8 (b) 4.9 × 10−5

(c) 4.9 × 10−9 (d) 4.9 × 10−7

Answer: (C)

Moles of $CaCO3$ in residue $= \frac{7\times10-3}{100} = 7 \times 10-5$

Moles of $CaCO_3$ in 1 litre solution = 7 × 10−5

$CaCO_3(s) \rightleftharpoons Ca^{2+} + CO_3^{2-}$

$7 \times 10^{-5} \quad 7 \times 10^{-5}$

$K_{sp} = [Ca^{2+}] \times [CO_3^{2-}]$

$= 7 \times 10^{-5} \times 7 \times 10^{-5} = 4.9 \times 10^{-9}$

3. The hydrolysis constant for $ZnCl_2$ will be

(a) $K_h = \frac{K_w}{K_b}$ (b) $K_h = \frac{K_w^2}{K_b}$

(c) $K_h = \frac{K_w}{K_b^2}$ (d) $K_h = \frac{K_b^2}{K_w}$

Where, K_b is effective dissociation constant of base Zn^{++}

Answer: (b)

$Zn^{++} + 2H_2O \rightleftharpoons Zn(OH)_2 + 2H^+$

$\therefore K_h = \frac{[Zn(OH)_2][H^+]^2}{[Zn^{++}]}$

$Zn(OH)_2 \rightleftharpoons Zn^{++} + 2OH^-$

$\therefore K_b = \frac{[Zn^{++}][OH^-]^2}{[Zn(OH)_2]}$

$K_w = [H^+][OH^-]$

$\therefore \frac{K_w^2}{K_b} = K_h$

4. M(OH)x has KSP 4 × 10^{-12} and solubility 10−4M. Then the value of x is

(a) 1 (b) 2

(c) 3 (d) −4

Answer: (b)

M(OH)x will ionize in the way

M(OH)x $\rightleftharpoons$ M+x + x OH$^-$

10−4 x × 10−4

∴ K_{sp} = [M+x] [O−x H]

∴ (10−4) (x × 10−4)x = 4 × 10−12

by inspection we get this relation will hold good when x = 2

5. For a weak electrolyte α1 and α2 are in ratio of 1: 2, for a given concentration ka_1=2×10−4. What will be value of k_{a2}?

(a) 8 × 10−4 (b) 2 × 10−4

(c) 4 × 10−4 (d) 1 × 10−4

Answer: (a)

$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{ka_1}{ka_2}}$$

$$\frac{1}{2} = \sqrt{\frac{2\times10^{-4}}{k_{a2}}}$$

$$\frac{1}{4} = \frac{2\times10-4}{ka_2}$$

$ka_2 = 8\times10^{-4}$

6. Consider the reaction $A^- + HO_3^+ \rightleftharpoons HA + H_2O$. The Ka value for acid HA is 1.0×10^{-6}. What is the value of K for this reaction.

(a) 1.0×10^{6} (b) 1.0×10^{-8}

(c) 1.0×10^{8} (d) 1.0×10^{-6}

Answer: (a)

The reaction is the reverse of the ionization reaction of HA, hence the equilibrium constant is the reciprocal of Ka.

$$K = \frac{[HA]}{[A^-]\,[H^+]} = \frac{1}{Ka} = \frac{1}{1.0 \times 10^{-6}} = 1.0 \times 10^{6}$$

7. The ionization constant of HCO_2H is 1.8×10^{-4}. What is the percent ionization of a 0.001 M solution?

(a) 66% (b) 42%

(c) 34% (d) 58%

Answer: (C)

$$HCO_2H \rightleftharpoons H^+ + HCO_2^-$$

$0.001 - x$ x x

α is > 0.1 which means is not ignorable.

$\therefore K_a = 1.8 \times 10^{-4} = \frac{x^2}{0.001 - x}$

Solving by quadratic,

$x = 3.4 \times 10^{-4}$

check for α

$\alpha = \sqrt{\frac{Ka}{C}}$

$\alpha = \sqrt{\frac{1.8 \times 10^{-4}}{0.001}} = 0.42$

$\therefore$ % ionization = $\frac{\text{ionized } HCO_2H}{\text{total } HCO_2H} \times 100 = \frac{3.4 \times 10^{-4}}{0.001} \times 100 = 34\%$

8. At −50°C, the self−ionization constant (ion product) is $K_{NH_3} = [NH_4^+][NH_2^-] = 10^{-30}$. How many amide ions are present per mm^3 of pure liquid ammonia?

(a) 600 ions/mm^3

(b) 6×10^6 ions/mm^3

(c) 6×10^4 ions/mm^3

(d) 60 ions/mm^3

Answer: (a)

$$2NH_3(l) \rightleftharpoons NH_4^+ + NH_2^-$$

$$K_{NH_3} = x^2 = 10^{-30}$$

thus, $x = 10^{-15}$ M = $[NH_2^-]$

$[NH_2^-] = 10^{-15}$ moles/lit = $\frac{10^{-15}}{10^6}$ moles/mm^3

$[NH_2^-] = 10^{-21} \times 6 \times 10^{23}$ ions/mm3 = 600 ions/mm3

9. The degree of dissociation of anilinium acetate is

(a) independent of initial concentration

(b) directly proportional to initial concentration

(c) inversely proportional to initial concentration

(d) inversely proportional to square root of initial concentration

Answer: (d)

$$C_6H_5NH_3^+CH_3CO_2^- \rightleftharpoons C_6H_5NH_3^+ + CH_3CO_2^-$$

$$C(1-\alpha) \qquad C\alpha \qquad C\alpha$$

$K_{Ceq} = \frac{\alpha^2}{(1-\alpha)}$ assuming α to be small, $(1-\alpha) \approx 1$

$$\therefore \alpha = \sqrt{\frac{K_{eq}}{C}}$$

10. 10 ml of 0.2 M acid is added to 250 ml of a buffer solution with pH = 6.34 and the pH of the solution becomes 6.32. The buffer capacity of the solution is

(a) 0.1 (b) 0.2

(c) 0.3 (d) 0.4

Answer: (d)

$$\text{Buffer capacity} = \frac{\text{moles of acid or base added per litre}}{\text{change in pH}}$$

$$\text{Buffer capacity} = \frac{\frac{0.2\times10}{1000}\times\frac{1000}{250}}{6.34-6.32}=0.4$$

Hence, (d) is the correct answer.

11. A weak base $B(OH)_2$ has dissociation constant 10^{-8}. The equilibrium constant for its reaction with excess of strong acid will be

(a) 10^{20} (b) 10^{-6}

(c) 10^{-20} (d) 10^{-16}

Answer: (a)

$$B(OH)_2 + 2H^+ \rightleftharpoons B^{2+} + 2H_2O$$

$$K = \frac{1}{K_h}$$

K_h for hydrolysis of B^{2+}

$$= \frac{[B(OH)_2][H^+]^2 \times [OH^-]^2}{[B^{2+}][OH^-]^2} = \frac{K_w^2}{K_b}$$

$$K_h = \frac{10^{-28}}{10^{-8}} = 10^{-20}$$

$$\frac{1}{K_h} = 10^{20}$$

Hence, (a) is the correct answer.

12. A weak monobasic acid (0.1 M) has a pH of 3 at a particular temperature (25°C). When this acid is neutralized by strong base (NaOH), what is the value of equilibrium constant at equivalent point at 25°C?

(a) 10^{-5} (b) 10^{-4}

10^{-14}

(c) 10^{-7} (d)

Answer: (a)

Let, the weak acid be HA

$$HA \rightleftharpoons H^{+} + A^{-}$$

$0.1(1-\alpha)$ 0.1α 0.1α

$0.1\alpha = 10^{-3}$

$\alpha = 1 \times 10^{-4}$

$$\therefore K_a = \frac{0.1\alpha \times 0.1\alpha}{0.1(1-\alpha)} = 0.1\alpha^2 = 10^{-9}$$

At equivalence point,

$$A^{-} + H_2O \rightleftharpoons HA + OH^{-}$$

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-9}} = 10^{-5}$$

Hence, (a) is the correct answer.

13. The concentration of $[OH^-]$ left in a solution after mixing 50 ml of 0.2 M ACl_2 with 50 ml 0.4 m NaOH. [Ksp of $A[OH]_2 = 1.2 \times 10^{-11}$]

(a) 2.8×10^{-4} m/L (b) 1.4×10^{-4} m/L

(c) 2×10^{-2} M (d) 4×10^{-2} M

Answer: (a)

$$ACl_2 + 2NaOH \longrightarrow A(OH)_2 + 2NaCl$$

Initially	10	20	0	0
After reaction	0	0	10	20

$[OH^-]^2 [A^+]^2 = K_{sp}$

$4s^3 = K_{sp}$

$$s = \sqrt[3]{\frac{1.2 \times 10^{-11}}{4}} = 1.4 \times 10^{-4}$$

$[OH^-] = 2.8 \times 10^{-4}$

Hence, (a) is the correct answer.

14. A weak acid HA after treatment with 12 ml of 0.1 M strong base BOH has a pH of 5. At the end point, the volume of same base required is 20 ml, then Ka of acid is

(a) 1.5×10^{-5} (b) 8.12×10^{-6}

(c) 1.8×10^{-6} (d) 8.2×10^{-5}

Answer: (a)

Eq. of acid = Eq. of base

= 20 × 0.1 = 2

$$HA + BOH \longrightarrow BA + H_2O$$

2	1.2	0	0

0.8   0   1.2   1.2

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

$5 = -\log K_a + \log\frac{[1.2]}{[0.8]}$

$\Rightarrow K_a = \text{antilog } 4.8 = 1.5 \times 10^{-5}$

Hence, (a) is the correct answer.

15. The H^+ ion concentration in 0.001 M CH_3COOH ($K_a = 1.8\times10^{-5}$) is 1.34×10^{-4} g ion/L. The H^+ ion concentration if 0.164 g of CH_3COONa is added to a litre of 0.001 M CH_3COOH will be

(a) 9×10^{-6} (b) 18×10^{-6}

(c) 4.5×10^{-6} (d) 5×10^{-6}

Answer: (a)

$pH = -\log 1.8\times10^{-5} + \log\frac{0.164/84}{0.001} = 4.74 + 0.3010 = 5.0457 = -\log[H^+]$

$\Rightarrow [H^+] = 9\times10^{-6}$

Hence, (a) is the correct answer.

16. The pH at which an acid indicator with $K_a = 1 \times 10^{-5}$ changes colours when indicator concentration is 1×10^{-5} M is

(a) 4 (b) 5

(c) 6 (d) 3

Answer: (b)

For acid indicator HIn,

$HI_n \rightleftharpoons H^+ + I_n^-$

Colour(A) Colour(B)

$K_{In} = \frac{[H^+][I_n^-]}{[HI_n]}$ when $[I_n^-] = [HI_n]$ indicator changes colour.

$\therefore K_{I_n} = [H^+] = 1 \times 10^{-5}$

$\therefore$ pH = 5

Hence, (b) is the correct answer.

17. If Kb for fluoride ion at 25°C is 1.48×10^{-11}, the ionization constant of hydrofluoric acid in water at this temperature is

(a) 1.7×10^{-5} (b) 3.52×10^{-3}

(c) 6.75×10^{-4} (d) 5.38×10^{-2}

Answer: (c)

$K_a \times K_w = K_\omega \Rightarrow \quad K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.48 \times 10^{-11}} = 6.75 \times 10^{-4}$

Hence, (c) is the correct answer.

18. 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) 3.75 (b) 5.6

(c) 7.5 (d) 3.4

Answer: (a)

Meq. of KOH = 50 × 0.2 = 10

Meq. of HCOOH = 40 × 0.2 = 20

∴ Meq. of KOH left = 10 Meq. of HCOOK formed = 10

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

$\Rightarrow pH = -\log K_a + \log \frac{10}{10} \Rightarrow pH = -\log Ka + \log \frac{10}{10}$

$\Rightarrow pH = -\log(1.8 \times 10-4) = 3.75$

Hence, (a) is the correct answer.

19. If the degree of ionization of water be 1.8 × 10-9 at 298 K. Its ionization constant will be

(a) 1.8 × 10-16 (b) 1 × 10-14

(c) 1 × 10-16 (d) 1.67 × 10-14

Answer: (a)

$K = \frac{K_w}{[H2O]} = \frac{10^{-14}}{55.5} = 1.8 \times 10^{-16}$

Hence, (a) is the correct answer.

20. When a solution of benzoic acid was titrated with NaOH, the pH of the solution, when half of the acid was neutralized, will be 4.2. Dissociation constant of the acid is

(a) 6.31 × 10-5 (b) 3.2 × 10−5

(c) 8.7 × 10−8 (d) 6.42 × 10−4

Answer: (a)

$$C6H5COOH + NaOH \rightleftharpoons C6H5COONa + H2O$$

Final 0.5 0.5

It is a buffer solution of weak acid and its salt.

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

pKa = 4.2

$K_a = 6.31 \times 10{-}5$

Hence, (a) is the correct answer.

21. 10-2 mole of NaOH was added to 10 litre of water. The pH will change by

4 (a) (b) 3

11 (c) (d) 7

Answer: (a)

pHold = 7

$[OH−]new = 10^{-2} \times \frac{1}{10} = 10^{-3}$

$\therefore [H^+]_{new} = 10^{-11}$

pHnew = 11, Change in pH = 4

Hence, (a) is the correct answer.

22. If an acidic indicator HIn Ionizes as HIn $\rightleftharpoons$ H+ + In−. To which maximum pH value its solution has distinct colour characteristic of HIn?

(a) pK in−1 (b) pKin ± 1

(c) pKin (d) 7

Answer: (b)

For colour characteristic of H indicator

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

Indicator’s colour range

pH = pKln ± 1

Hence, (b) is the correct answer.

23. The H+ ion concentration in 0.001M CH_3CO_2H (K_a = 1.8×10−5) is 1.34×10−4gion/L. The H+ ion concentration if 0.164g of CH_3COONa is added to a litre of 0.001M CH3CO2H will be

(a) 9×10−6 (b) 18×10−6

4.5×10−6

(c) (d) 5×10−6

Answer: (a)

$pH = -\log 1.8\times10^{-5} + \log\frac{0.1\ 64/84}{0.001}$

$= 4.74 + 0.3010 = 5.0457 = -\log[H^+]$

$\Rightarrow [H^+] = 9\times10^{-6}$

Hence (a) is the correct answer.

24. What is the minimum pH required to prevent the precipitation of ZnS in a solution that is 0.01 M $ZnCl_2$ and saturated with 0.10 M H_2S ?

[Given $K_{sp} = 10^{-21}$, $K_{a_1} \times K_{a_2} = 10^{-20}$]

(a) 0 (b) 1 (c) 2 (d) 4

Answer: (b)

$K_{sp} = [Zn^{2+}][S^{2-}]$

$[S^{2-}] = \frac{10^{-21}}{0.01} = 10^{-19}$

for $K_{a_1} . K_{a_2} = \frac{[H^+]^2[S^{2-}]}{[H_2S]}$

$10^{-20} = \frac{[H^+]^2 \times 10^{-19}}{0.1} \Rightarrow [H^+] = 0.1$

or pH = 1

25. Ammoniacal solution of Ag^+ + Acidified solution of N_2H_4 → ppt of (X)

(a) (X) is Ag (b) (X) is Ag2O (c) (X) is AgNH2

(d) (X) is Ag2SO4

Answer: (a)

4[Ag(NH)]+32 + H2N – NH2. H2SO4 →

$4Ag\downarrow + N_2 + 6NH_4^+ + 2NH_3 + SO_4^{2-}$

26. 100 ml solution (I) of buffer containing 0.1(M) HA and 0.2 (M) A−, is mixed with another solution (II) of 100 ml containing 0.2(M) HA and 0.3(M) A− .

After mixing what is the pH of resulting solution ?

Given pKa of HA = 5

(a) 5 – log 5/3

(b) 5 + log 5/3

(c) 5 + log 2/5

(d) 5 – log 5/2

Answer: (b)

After mixing total moles of

A− = 100 × 0.2 × 10−3 + 100 × 0.3 × 10−3 = 100 × 10−3 × 0.5 moles

After mixing total moles of

HA = 100 × 0.1 × 10−3 + 100 × 0.2 × 10−3

= 100 × 0.3 × 10−3 moles

After mixing resulting $pH = 5 + \log\frac{5}{3}$

27. Which of the following is/are soft bases?

H_2O, H^-, CO, CO_2, C_2H_4, CN^-

(a) C_2H_4

(b) H^-, CN^-, CO

(c) H^-, CN^-, CO, C_2H_4

(d) H^-, CO, CO_2, C_2H_4

Answer: (c)

Soft bases have the donor atom of highly polarisable and they preferably combine with the metal ions of lower oxidation states.

28. Calcium lactate is a salt of weak organic acid and strong base represented as $Ca(LaC)_2$. A saturated solution of $Ca(LaC)_2$ contains 0.6 mole in 2 litre solution. pOH of solution is 5.60. If 90% dissociation of the salt takes place, then what is pKa of lactic acid ?

(a) 2.8 – log (0.54)

(b) 2.8 + log (0.54)

(c) 2.8 + log (0.27)

(d) None of these

Answer: (a)

$Ca(LaC)_2 \rightarrow Ca2+ + 2\,LaC^-$

Initial concentration 0.6/2 = 0.3 M

After dissociation [LaC] = 2 × 0.3 × 0.9 = 0.54 M

$pH = \frac{1}{2}[pK_w + pKa + \log(La\,C)]$

$14 - 5.6 = \frac{1}{2}[14 + pK_a + \log(0.54)]$

$pKa = 2.8 - \log(0.54)$

29. Fixed volume of 0.1 M benzoic acid (pKa = 4.2) solution is added into 0.2 M sodium benzoate solution and formed a 300 ml, resultant acidic buffer solution. If pH of this buffer solution is 4.5 then find added volume of benzoic acid -

(a) 100 ml (b) 150 ml

(c) 200 ml (d) None of these

Answer: (c)

$pH = pK_a + \log\frac{[C_6H_5COO^-]}{[C6H5COOH]}$

$\therefore \frac{[C_6H_5COO^-]}{[C6H5COOH]} = 2$

Let volume of acid is V ml.

$\frac{0.2\times(300-V)}{0.1\times V} = 2$

⇒ V = 200 ml.

30. The Ksp of FeS = 4 × 10−19 at 298 K. The minimum concentration of H+ ions required to prevent the precipitation of FeS from a 0.01 M solution Fe2+ salt by passing H2S is –

(given [H+]2[S−2] = 1 × 10−21)
0.1

(a) 1.6 × 10−3 M (b) 2.5 × 10−4 M

(c) 2.0 × 10−2 M (d) 1.2 × 10−4 M

Answer: (a)

[Fe2+] [S−2] = 4 × 10−19 ⇒ [S−2] = 4×10−19 = 4×10−17 M
1×10−2

In order to precipitate FeS, [S−2] required is 4 × 10−17 M from 0.01 M Fe2+ salts.

Now [H+]2[4×10−17] = 1 × 10−21
0.1

⇒ [H+]2 = 2.5 × 10−6

⇒ [H+] = 1.6 × 10−3

31. The equilibrium constant for the given reaction is approximately 10−3

HPO2−4(aq) + HCO−3(aq) H2PO−4(aq) + CO2−3(aq)

Which is strongest conjugate base in the given reaction?

(a) HPO2−4(aq) (b) HCO−3(aq)

(c) H_2PO_4−(aq) (d) CO32−(aq)

Answer: (d)

$CO_3$2− + H_2PO−4⇌ HPO2 4 + HCO32− Kc = 10 ^{3}

Therefore CO2−3 is the stronger conjugate base than HPO2−4

32. Which of the following constitutes a set amphiprotic species.

(a) H +3O , H_2PO4,− HCO−^{3} (b) H2O , HPO2−4, ^{2}H2PO−

(c) H2O , H2PO−3, HPO2−4 (d) HC2O−4, H2PO−4, SO2−4

Answer: (c)

$H3O^+$ cannot take up proton; H2PO−2 cannot give up proton, SO2−4 cannot give proton

33. Which of the following order represent the order for the strength of base?

(a) CH_3CH−2 > NH−2 > HC ≡ C− > OH−

(b) H − C ≡ C− >CH−−−3CH2 > NH2> OH

(c) OH− > NH−2 > HC ≡ C− > CH− 3CH2

(d) NH−≡− −−2 > HC C> OH > CH3CH2

Answer: (a)

The strength of acid is in the order of

CH3 − CH3 < NH3 < HC ≡ CH < H2O

34. Which of the following order of acid strengths are correct?

(a) HOI < HOBr < HOCl

(b) HCl < HBr < HI

(c) HIO4 < HBrO4 < HClO4

(d) All of the above are correct order

Answer: (d)

In HOX with increase in electronegativity acid strength increases

In H – X acid strength increases with decrease in bond energies.

In HXO4 acid strength increases with decrease in size of X.

35. The pH of Ba(OH)2 solution is 13. The number millimoles of Ba(OH)2 present in 10 ml of solution would be

(a) 1.00 (b) 0.50 (c) 10.00 (d) 15.00

Answer: (b)

[OH−] = 0.1

And conc. of Ba(OH) 2 = $\frac{0.1}{2}$

∴ The no. of millimoles of Ba(OH)2 present in 10 ml solution

$= \frac{0.1}{2} \times 10 \times 10-3 \times 103 = 0.5$

36. For 10−2(M) H3PO3 solution which of the following relations is correct?

(a) $[H3PO3] + [H_2PO_3^-] + [HPO_3^{2-}] + [PO_3^{3-}] = 10^{-2}$

(b) $[H3PO3] + [H_2PO_3^-] + [HPO_3^{2-}] = 10^{-2}$

(c) $[H_2PO_3^-] + [HPO_3^{2-}] + [PO_3^{3-}] = 10^{-2}$

(d) $[H_3PO_3] + [H_2PO_3^-] + 2[HPO_3^{2-}] = 10^{-2}$

Answer: (b)

$H_3PO_3 \rightleftharpoons H_2PO_3^- + H^+$

$H_2PO_3^- \rightleftharpoons HPO_3^{2-} + H^+$

H_3PO_3 is a dibasic acid

37. The pH of 0.01 (M) KOH is 12; if the temperature of the given KOH solution is increased which of the following would occur ?

(a) Both pH and pOH would remain constant

(b) pH would be decreased but pOH remain constant

(c) pH would be increased while pOH remain constant

(d) pH would be increased while pOH would be decreased

Answer: (b)

Since, KOH is a strong electrolyte its dissociation remain unaffected, hence, $[OH^-] = 10^{-2}$
or pOH = 2 but on increasing temperature pKw < 14

or pH + pOH < 14

or pH < 12

Therefore, pH would be decreased.

38. 4 ml of HCl solution of pH = 2 is mixed with 6 ml of NaOH solution of pH = 12. What would be the final pH of solution? log 2 = 0.3

(a) 10.3 (b) 11.3 (c) 11 (d) 4.3

Answer: (b)

pH = 2, [HCl] = 10−2 (M) & pOH = 2

or [NaOH] = 10−2 (M)

4 ml of 10−2 (M) HCl 4 × 10−5 moles HCl.

6 ml of 10−2 (M) NaOH 6 × 10−5 moles NaOH

After mixing excess moles of OH− = 2 × 10−5

$[\mathrm{OH}-] = \frac{2\times10}{10}-5\times 103 = 2 \times 10-3$

or pOH = 3 − log 2 = 3 − 0.3 = 2.7

or pH = 11.3

39. The pH of 0.1(M) H2SO4 in water is

(a)is equal to one

(b)is more than one

(c)is less than one

(d)may be equal to or greater than or smaller than one depending upon temperature

Answer: (c)

H2SO4 → H+ + HSO−4

$HSO_4^- \rightleftharpoons H+ + SO_4 2-$

In solution [H+] > 0.1(M), hence pH < 1

40. A solution of NaCl is

(a)acidic in nature

(b)alkaline in nature

(c)neutral in nature

(d)amphoteric in nature

Answer: (c)

A solution of NaCl is neutral as both Na+ and Cl− ions are weak conjugate acid and base respectively.

41. A solution of ammonium cyanide is

(a) acidic in nature (b) alkaline in nature

(c) neutral in nature (d) amphoteric in nature

Answer: (b)

A solution of NH_4CN is alkaline because CN− is more strong conjugate base than NH+4, conjugate acid. The hydrolysis of CN− proceeds more that of NH+4.

42. The pH of 0.1 (M) solution of the following salts increases in the order

(a) NaCl < NH4Cl < NaCN < HCl

(b) HCl < NH4Cl < NaCl < NaCN

(c) NaCN < NH4Cl < NaCl < HCl

(d) HCl < NaCl < NaCN < NH4Cl

Answer: (b)

pH of NaCl = 7 ; pH of NH4Cl < 7 ; pH of NaCN > 7 and pH of 0.1 (M) HCl = 1

43. HA is a weak acid and BOH is a weak base. For which of the following salts the extent of hydrolysis is independent on the concentration of the salt in its aqueous solution

(a) NaA (b) NaB (c) BCl (d) BA

Answer: (c)

The extent of the hydrolysis (h) of the salt made by weak acid and weak base is independent on the concentration of salt, h = $\frac{\sqrt{Kh}}{\sqrt{Kh}+1}$

K_h = hydrolytic constant = $\frac{K_w}{K_aK_b}$

44. In alkaline buffer solution containing NH4OH and NH4Cl , if some HCl is added its pH will

(a) increase a little

(b) decrease a little

(c) remains constant

(d) change but cannot be predicted

Answer: (b)

The addition of HCl causes the replacement of base (NH4OH) with salt (NH4Cl). The ratio of salt to base increases. Hence the pOH of buffer solution will increase and its pH will decrease.

45. A solution contains 0.05 (M) each of NaCl and Na2CrO4. Solid AgNO3 is gradually added to it. Ksp (AgCl) = 1.7 × 10−10 and K_{sp} (Ag_2CrO 4) =1.9 × 10^{-12}

Which of the following would occur?

(a) Cl− ions are precipitated first

(b) CrO2−4 ions are precipitated first

(c) Both Cl− and CrO2−4 ions are precipitated together

(d) The second ion starts precipitating when half of the first ion is precipitated

Answer: (a)

To precipitate Cl−, required [Ag+] = $\frac{(K_{sp})_{AgCl}}{0.05}$

To precipitate CrO_4^{2-}, required [Ag+] = $\left[\frac{(K_{sp})_{Ag_2CrO_4}}{0.05}\right]^{1/2}$

Since, for the precipitation of Ag2CrO4 required concentration of Ag+ is greater, hence Cl− ions precipitate first in the form of AgCl (s).