Q1: NTA Test 01 (Single Choice)

In an aqueous solution, the ionization constants for carbonic acid are

$K_1 = 4.2 \times 10^{-7}$ and $K_2 = 4.8 \times 10^{-11}$

Select the correct statement for a saturated 0.034 M solution of the carbonic acid:

(A) The concentration of H^+ is double that of CO_3^{2-}.

(B) The concentration of CO_3^{2-} is 0.034 M.

(C) The concentration of CO_3^{2-} is greater than that of HCO_3^-.

(D) The concentration of H^+ and HCO_3^- are approximately equal.

Q2: NTA Test 02 (Single Choice)

At 25°C, pH range of phenolphthalein is $8 - 10$. At 100°C, pH range of phenolphthalein would be -

(A) pH range remain unaffected by the temperature

(B) pH range is altered to 7 - 9

(C) pH range is altered to 7 - 11

(D) pH range is altered to 8 - 11

Q3: NTA Test 03 (Single Choice)

The equilibrium constant for :

$CN^- + CH_3COOH \rightleftharpoons HCN + CH_3COO^-$ is :

$\left(\text{Given } pK_b \text{ for } CN^- = 4.69 \text{ and } pK_b \text{ for } CH_3COO^- = 2.25\right)$

(A) 3.7×10^4

(B) 2.8×10^{-5}

(C) 1.97×10^4

(D) 0.5×10^{-5}

Q4: NTA Test 04 (Numerical)

Bromine in excess is dropped to a 0.01 M SO_2. All of SO_2 is oxidized to H_2SO_4 and the excess Br_2 is removed by flushing with gaseous N_2. Determine the pH of the resulting solution assuming K_{a1} of H_2SO_4 very large & $K_{a2} = 10^{-2}$. Take the value of $\log(3.24) = 0.51$.

Q5: NTA Test 05 (Single Choice)

The K_{sp} of FeS $= 4 \times 10^{-19}$ at 298 K. The minimum concentration of H^+ ions required to prevent the precipitation of FeS from a 0.01 M solution Fe^{2+} salt by passing H_2S

(Given H_2S $k_{a_1} \times k_{b_1} = 10^{-21}$)

(A) 1.6×10^{-3} M

(B) 2.5×10^{-4} M

(C) 2.0×10^{-2} M

(D) 1.2×10^{-4} M

Q6: NTA Test 07 (Single Choice)

K_a for HCN is 5×10^{-10} at 25° C. For maintaining a constant pH of 9, the volume of 5 M KCN solution required to be added to 10 mL of 2 M HCN solution is:

(A) 4 mL

(B) 7.95 mL

(C) 9.3 mL

(D) 2 mL

Q7: NTA Test 08 (Single Choice)

Minimum moles of NH_3 required to be added to 1L solution so as to dissolve 0 .1 mol of AgCl $\left(K_{sp} = 1.0 \times 10^{-10}\right)$ by the reaction is:

$AgCl(s) + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]^+ + Cl^-$

Given K_f of $\left[Ag(NH_3)_2\right]^+ = 10^8$

(A) 0 .5 mol

(B) 1 .0 mol

(C) 1 .1 mol

(D) 1 .2 mol

Q8: NTA Test 08 (Single Choice)

For the indicator, HIn; the ratio $\frac{[\mathrm{In}^-]}{[\mathrm{HIn}]}$ is 7 .0 at pH of 4 .3. K_{In} for the indicator is

[Given: log $7 = 0.845$ and Antilog $(-3.455) = 3.5 \times 10^{-4}$]

(A) 3.5×10^{-4}

(B) 3.5×10^{-5}

(C) 3.5×10^{-2}

(D) 3.5×10^{-3}

Q9: NTA Test 10 (Single Choice)

If the pK_a of acetic acid and pK_b of NH_4OH are 4.76 and 4.75 respectively, what will be the pH of ammonium acetate solution?

(A) 9.51

(B) 7.005

(C) 7.00

(D) 6.9

Q10: NTA Test 11 (Single Choice)

What is $[H^+]$ in a solution that is 0.01 M in HCN and 0.02 M in NaCN?
br /> ($K_a$for HCN $= 6.2 \times 10^{-10}$)

(A) 3.1×10^{10}

(B) 6.2×10^{5}

(C) 6.2×10^{-10}

(D) 3.1×10^{-10}

Q11: NTA Test 13 (Numerical)

If the concentration of $[NH_4^+]$ in a solution having 0.02 M NH_3 and 0.005 M $Ca(OH)_2$ is $a \times 10^{-6}$ M , determine a .
$[k_b\ (NH_3) = 1.8 \times 10^{-5}]$

Q12: NTA Test 15 (Single Choice)

What is the $[OH^-]$ concentration of a 0.04 M solution of CH_3COONa?

[K_a of $CH_3COONa = 2 \times 10^{-5}$, $\log 2 = 0.30$, $\log 5 = 0.69$, $10^{0.7} \cong 5$]

(A) 5×10^{-6}

(B) 6×10^{-6}

(C) 2×10^{-9}

(D) 3×10^{-9}

Q13: NTA Test 16 (Single Choice)

Calculate dissociation constant for an acid HA having $[H^+]$ equals to 0.1 M (Given: degree of dissociation of HA is 1%)

(A) 10^{-8}

(B) 10^{-6}

(C) 10^{-3}

(D) 10^{-7}

Q14: NTA Test 17 (Single Choice)

The Solubility of AB_2 is 0.05 g per 100 mL at 25° C. Calculate K_{sp} of AB_2 at 25° C? [Atomic mass of A = 20 amu, atomic mass of B = 40 amu]

(A) 10^{-3}

(B) 5×10^{-7}

(C) 10^{-6}

(D) 5×10^{-3}

Q15: NTA Test 18 (Numerical)

One litre of 1 M solution of an acid HA $\left(K_a = 10^{-4} \text{ at } 25\ ^{o}C\right)$ has pH = 2. It is diluted by water so the new pH becomes double. The solution was diluted to $y \times 10^{z}$ ml . The value of $\frac{y+z}{2}$ is:

Q16: NTA Test 19 (Single Choice)

When 0.01 moles of the following acids are dissolved in 1 L of H_2O , the $\left[H^+\right]$ will be greatest in: -

(A) HNO_2 ;pka= 3.0 (B) HCOOH ; pka = 3.75

(C) HCN; pka = 9.4 (D) CH_3COOH; pka = 4.75

Q17: NTA Test 20 (Single Choice)

pH of a 1000 cc solution is 2. It will not change if

(A) 100 cc of water is added to it (B) 100 cc of 0.1 M HCl is added to it

(C) 100 cc (N/100) HCl is added to it (D) 1 cc of 0.1 M HCl is added to it.

Q18: NTA Test 21 (Single Choice)

When 10 mL of 0.1 M acetic acid (pK_a = 5) is titrated against 10 mL of 0.1 M ammonia solution (pK_b = 5), the equivalent point will occur at pH:

(A) 5 (B) 6

(C) 7 (D) 9

Q19: NTA Test 22 (Single Choice)

If the pK_a of a weak acid HA is 4.80 and the pK_b of a weak base BOH is 4.78. Then, the pH of an aqueous solution of the corresponding salt, BA will be

(A) 8.23 (B) 9.41

(C) 7.01 (D) 5.91

Q20: NTA Test 23 (Single Choice)

The ionization constant of a weak monobasic acid is 25×10^{-6} while the equivalent conductance of its 0.01 M solution is 19.6 Scm^2eq^{-1} . The equivalent conductance of the electrolyte at infinite dilution $\left(in\ Scm^2eq^{-1}\right)$ will be

(A) 250 (B) 196

(C) 392 (D) 384

Q21: NTA Test 25 (Single Choice)

The solution of Na_2CO_3 has pH

(A) greater than 7 (B) less than 7

(C) equal to 7 (D) equal to zero

Q22: NTA Test 26 (Single Choice)

Using the Gibbs energy change, $\Delta G^\circ = +63.3$ kJ , for the following reaction,

$Ag_2CO_3 \rightarrow 2Ag + (aq) + CO_3^{2-}(aq)$

the K_{sp} of $Ag_2CO_3(s)$ in water at 25°C is

$(R = 8.314\ J\ K^{-1} mol^{-1})\{Given:\ 10^{0.91} \cong 8\}$

(A) 7.9×10^{-2} (B) 8.0×10^{-12}

(C) 2.9×10^{-3} (D) 3.2×10^{-26}

Q23: NTA Test 27 (Single Choice)

The solubility of a sparingly soluble salt $A(OH)_2$ (mol. wt. 192.3) is 19 .23 g/ litre at 300 K. The pH of its saturated solution assuming 80% ionisation at 300 K is :

(A) 1.0970 (B) 12.9030

(C) 13.2041 (D) 12.0000

Q24: NTA Test 28 (Single Choice)

A 50 mL solution of pH = 1 is mixed with a 50 mL solution of pH = 2. The pH of the mixture is

(A) 0.86

(B) 1.26

(C) 1.76

(D) 2.26

Q25: NTA Test 30 (Single Choice)

Find out the percentage dissociation of an acid having conc. of 10 M and dissociation constant 1.0×10^{-3} .

(A) 0.1

(B) 0.5

(C) 1.0

(D) 2.0

Q26: NTA Test 31 (Single Choice)

A 0.010 M solution of maleic acid, a monoprotic organic acid, is 14% ionized. What is K_{a_1} for maleic acid?

(A) 2.3×10^{-3}

(B) 2.3×10^{-4}

(C) 2.0×10^{-4}

(D) 2.0×10^{-6}

Q27: NTA Test 32 (Single Choice)

The solubility product of AgCl is 1.8×10^{-10}. Precipitation of AgCl will occur only when equal volumes of which of the following solutions are mixed?

(A) 10^{-4} M Ag^+ and 10^{-4} M Cl^-

(B) 10^{-7} M Ag^+ and 10^{-7} M Cl^-

(C) 10^{-5} M Ag^+ and 10^{-5} M Cl^-

(D) 10^{-10} M Ag^+ and 10^{-10} M Cl^-

Q28: NTA Test 33 (Single Choice)

1M NH_4OH and 1M HCl are mixed to make a total volume of 300 mL. If pH of the mixture is 9.26 and $pK_a\left(NH_4^+\right) = 9.26$ then what would be the volume ratio of NH_4OH and HCl .

(A) 2:1

(B) 1:2

(C) 2:3

(D) 3:2

Q29: NTA Test 34 (Single Choice)

0.1 M NaOH is titrated with 0.1 M, 20 ml HA till the end point, $K_a(HA) = 6\times10^{-6}$ and degree of dissociation of HA is small as compared to unity. What is the pH of the resulting solution at the end point?

(A) 6.23

(B) 9.22

(C) 7.21

(D) 8.95

Q30: NTA Test 34 (Single Choice)

The degree of dissociation of a weak monoprotic acid of concentration 1.2×10^{-3} M having $K_a = 1.0 \times 10^{-4}$ is

(A) 1

(B) 10

(C) 15

(D) 25

Q31: NTA Test 35 (Single Choice)

The degree of dissociation of acetic acid in a 0.1 M solution is 1.0×10^{-2} . The pK_a of acetic acid value.

(A) 3

(B) 4

(C) 5

(D) 6

Q32: NTA Test 36 (Numerical)

The pH at the equivalent point for the titration of 0.10 M KH_2BO_3 with 0.1 M HCl is

$\left(K_a \text{ of } H_3BO_3 = 12.8\times10^{-10}\right)$

Report your answer by rounding it up to nearest whole number.

Q33: NTA Test 38 (Single Choice)

20 ml of 0.1 M NaOH is added to 30 ml0.1 M NaOH 30 ml 0.2 M CH_3COOH $\left(pK_a = 4.74\right)$, calculate the pH of the resulting solution? [given: $\log 2 = 0.3$]

(A) 3.44

(B) 4.01

(C) 4.44

(D) 4.71

Q34: NTA Test 39 (Numerical)

Calculate the pH of solution when 100 mL, 0.1M CH_3COOH and 100 mL, 0.1M HCOOH are mixed together. (Given: $K_a\,(CH_3COOH) = 2\times10^{-5}, K_a\,(HCOOH) = 6\times10^{-5}$)

Q35: NTA Test 40 (Numerical)

In 1 L saturated solution of AgCl $\left[K_{sp}\,(AgCl) = 1.6\times10^{-10}\right]$, 0.1 mol of CuCl $\left[K_{sp}\,(CuCl) = 1.0\times10^{-6}\right]$ is added. The resultant concentration of Ag^+ in the solution is 1.6×10^{-x} . The value of " x " is

Q36: NTA Test 41 (Single Choice)

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

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

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

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

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

Q37: NTA Test 42 (Single Choice)

The pH of 0.1 M solution of the following salts increases in the order

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

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

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

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

Q38: NTA Test 43 (Numerical)

30 ml of 0.2 M NaOH is added with 50 ml 0.2 M CH_3COOH solution. The extra volume of 0.2 M NaOH required to make the pH of the solution 5.00 is $\frac{10}{x}$ .The value of x is. The ionisation constant of $CH_3COOH = 2\times10^{-5}$.

Q39: NTA Test 44 (Single Choice)

Solid Na_2SO_4 is slowly added to a solution which is 0.020 M in $Ba(NO_3)_2$ and 0.020 M in $Pb(NO_3)_2$. Assume that there is no increase in volume on adding Na_2SO_4 . There preferential precipitation takes place. What is the concentration of Ba^{2+} when $PbSO_4$ starts to precipitate?

$[K_{sp}\,(BaSO_4) = 1.0\times10^{-10}$ and $K_{sp}\,(PbSO_4) = 1.6\times10^{-8}]$

(A) 5.0×10^{-9} M

(B) 8.0×10^{-7} M

(C) 1.25×10^{-4}M

(D) 1.95×10^{-8}M

Q40: NTA Test 45 (Single Choice)

For two weak acids A and B, the ratio of their percent ionization is 4:9. The ratio of their K_a would be-

(A) 4:9

(B) 2:3

(C) 16:81

(D) 3:2

Q41: NTA Test 47 (Single Choice)

Which one of the following mixture does not act as a buffer solution-

(A) Boric acid and borax

(B) Sodium Phosphate & disodium hydrogen phosphate

(C) Sodium propionate and propionic acid

(D) Sodium acetate and sodium propionate

Q42: NTA Test 47 (Numerical)

How many among the following species can be classified as Lewis acids?

$\overset{\oplus}{C}H_3, Cl^{\oplus}, CO_2, CCl_2, BCl_3, BI_3, Fe^{+2}, AlCl_3$

Q43: NTA Test 48 (Single Choice)

If pK_b for CN^- at $25^\circ C$ is 4.7, the pH of 0.5 M aqueous NaCN solution is

(A) 10

(B) 11.5

(C) 11

(D) 12

Answer Keys

Q1: (D)
Q2: (B)
Q3: (A)
Q4: 1.49
Q5: (A)
Q6: (D)
Q7: (D)
Q8: (A)
Q9: (B)
Q10: (D)
Q11: 36
Q12: (A)
Q13: (C)
Q14: (B)
Q15: 5.5
Q16: (A)
Q17: (C)
Q18: (C)
Q19: (C)
Q20: (C)
Q21: (A)
Q22: (B)
Q23: (C)
Q24: (B)
Q25: (C)
Q26: (B)
Q27: (A)
Q28: (A)
Q29: (D)
Q30: (D)
Q31: (C)
Q32: 5
Q33: (C)
Q34: 2.7
Q35: 7
Q36: (B)
Q37: (B)
Q38: 3
Q39: (C)
Q40: (C)
Q41: (D)
Q42: 8
Q43: (B)

Solutions

Q1: (D) The concentration of H^+ and HCO_3^- are approximately equal.

$$\begin{aligned} H_2CO_3 &\rightleftharpoons H^+ + HCO_3^-; \quad K_1 = 4.2 \times 10^{-7} \\ HCO_3^- &\rightleftharpoons H^+ + CO_3^{2-}; \quad K_2 = 4.8 \times 10^{-11} \end{aligned}$$

$\because K_1 \gg K_2$, so H_2CO_3 ionises more than HCO_3^- and hence, contribution of H^+ is mostly due to ionization of carbonic acid. Thus, the concentrations of H^+ and HCO_3^- are approximately equal.

Q2: (B) pH range is altered to 7 - 9

The pH range of the indicator is $pK_{In} \pm 1$. On increase in temperature pK_{In} decreases therefore pH range of indicator decreases.

Q3: (A) 3.7×10^4

$$K_{eq.} = \frac{[HCN]\left[CH_3COO^-\right]}{\left[CN^-\right][CH_3COOH]};$$

$$K_{CH_3COOH} = \frac{\left[CH_3COO^-\right]\left[H^-\right]}{[CH_3COOH]};$$

$$K_{HCN} = \frac{\left[H^+\right]\left[CN^-\right]}{[HCN]}$$

pK_b of $CN^- = 4.69 \therefore pK_{HCN} = 9.31$; $K_{HCN} = 4.9 \times 10^{-10}$

$pK_{CH_3COOH} = 4.75$; $K_{CH_3COOH} = 1.77 \times 10^{-5}$

$$\therefore K_{eq.} = \frac{K_{CH_3COOH}}{K_{HCN}} = \frac{1.77\times10^{-5}}{4.9\times10^{-10}} = 3.7 \times 10^4$$

Q4: 1.49

$$\begin{array}{lllll} & Br_2 + & SO_2 + & H_2O \rightarrow H_2SO_4 & + 2HBr \\ t=0 & & 0.01\,M & & \\ t=t & & & 0.01 & 0.02\,M \end{array}$$

$[H^+]$ from H_2SO_4 0.01 M

$[H^+]$ from HBr 0.02 M

0.03 M

$$\underset{(0.01\text{-}x)}{HSO_4^-}\,(aq) \rightleftharpoons \underset{(0.03+x)}{H^+}\,(aq) + \underset{x}{SO_4^{2+}};\ Ka_2 = 10^{-2}$$

$$Ka_2 = 10^{-2} = \frac{x(0.03+x)}{(0.01-x)}$$

$x = 2.36 \times 10^{-3} = 0.00236$

$\left[H^+\right]_{Total} = 0.03 + 0.00236$

$= 0.03236 = 3.24 + 0.00236$

$pH = -\log 3.240 \times 10^{-2}$

$= 2 - \log 3.24 = 1.49$

Q5: (A) 1.6×10^{-3} M

$$\left[Fe^{2+}\right]\left[S^{-2}\right] = 4 \times 10^{-19} \Rightarrow \left[S^{-2}\right] = \frac{4\times10^{-19}}{1\times10^{-2}} = 4 \times 10^{-17}\ M$$

In order to precipitate FeS, $\left[S^{-2}\right]$ required is 4×10^{-17} M from 0.01 M Fe^{2+} salts.

Now $\frac{\left[H^+\right]^2\left[4\times10^{-17}\right]}{0.1} = 1 \times 10^{-21}$

$\Rightarrow\ [H^+]^2 = 2.5 \times 10^{-6}$

$\Rightarrow\ [H^+] = 1.6 \times 10^{-3}$

Q6: (D) 2 mL

On mixing the given solutions, acidic buffer will be formed. Thus applying Henderson equation to find the pH-

$$pH = pK_a + \log_{10} \frac{[CN^-]}{[HCN]}$$

$$pH = -\log 5 \times 10^{-10} + \log \frac{5 \times V/V+10}{10 \times 2/V+10}$$

$$9 = -\log 5 \times 10^{-10} + \log \frac{V}{4}$$

$$9 = 10 - 0.7 + \log \frac{V}{4}$$

$$-0.3 = \log \frac{V}{4}$$

$$\therefore\ \log \frac{V}{4} =^{-} 0.3 = \log \frac{1}{2}$$

$$\frac{V}{4} = \frac{1}{2}$$

$\therefore\ V = 2$ mL.

Q7: (D) 1 .2 mol

$$AgCl_{(s)} + aq \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{aq} \quad ; \quad K_{sp}$$
$$Ag^{+}_{aq} + 2NH_{3_{aq}} \rightleftharpoons \left[Ag(NH_3)^{+}_{2_{aq}}\right] \quad ; \quad K_f$$

$$AgCl_{(s)} + 2NH_3 \rightleftharpoons \left[Ag(NH_3)^{+}_{2_{aq}}\right] + Cl^{-}_{aq} \qquad K = K_{sp} \times K_f$$

$(C - 0.2)M \qquad 0.1\ M \qquad 0.1\ M$ Equilibrium concentration

$$K = 10^{-10} \times 10^{8} = 10^{-2} = \frac{\left[Ag[NH_3)^{+}_{2}\right]\left[Cl^{-}\right]}{[NH_3]^2}$$

$$10^{-2} = \frac{10^{-1} \times 10^{-1}}{(M-0.2)^2}$$

$$\therefore\ M - 0.2 = 1$$

$$\therefore\ M = 1.2\ M$$

Q8: (A) 3.5×10^{-4}

For weak organic acid indicators

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

$$4.3 = pK_{In} + \log 7$$

$pK_{In} = 4.3 - 0.845 = 3.455$

$pK_{In} = -\log_{10} K_{In}$

$K_{In} = \text{Antilog}\left(-pK_{In}\right)$

$K_{In} = \text{Antilog}\ (-3.455)$

$= 3.5 \times 10^{-4}$

Q9: (B) 7.005

When a weak acid (acetic acid) and weak base $\left(NH_4OH\right)$ are mixed to form a solution. pH of the solution is then determined by following formula:

$$pH = 7 + \frac{1}{2}\left[pK_a - pK_b\right]$$

Given, $pK_a = 4.76$, $pK_b = 4.75$

$$pH = 7 + \frac{1}{2}\left[4.76 - 4.75\right]$$

$pH = 7 + 0.005$

$pH = 7.005$

Q10: (D) 3.1×10^{-10}

$$K_a = \frac{\left[H^+\right]\left[CN^-\right]}{[HCN]}$$

$$6.2 \times 10^{-10} = \frac{[H^+]\ [0.02]}{[0.01]}$$

$\left[NaCN\right] = \left[CN^-\right]$ due to common ion effect

$$\left[H^+\right] = \frac{6.2 \times 10^{10} \times 0.01}{0.02} = 3.1 \times 10^{-10}$$

Q11: 36

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

$$k_b = \frac{[NH_4^+]\ [OH^-]}{[NH_4OH]}$$

$$\Rightarrow\ 1.8 \times 10^{-5} = \frac{[NH_4^+]\ [0.01]}{[0.02]}$$

{As OH^- will mainly come from $Ca\left(OH_2\right)$ only}

$$\Rightarrow \left[NH_4^+\right] = 36 \times 10^{-6} = a \times 10^{-6}$$

$$\Rightarrow\ a = 36$$

Q12: (A) 5×10^{-6}

$$pK_a = -\log K_a = -\log\left(2 \times 10^{-5}\right)$$

$= 5 - \log 2 = 5 - 0.3 = 4.7$

$\mathrm{pH} = \frac{1}{2}\left(\mathrm{pK_w} + \mathrm{pK_a} + \log \mathrm{C}\right)$

$\mathrm{pH} = \frac{1}{2}\left[14 + 4.7 + \log(0.04)\right]$

$\mathrm{pH} = 8.63$

$\mathrm{pOH} = 14 - 8.63 \cong 5.3$

$\left[\mathrm{OH^-}\right] = 5 \times 10^{-6}$

Q13: (C) 10^{-3}

$\mathrm{k_a} = \frac{\alpha^2 \mathrm{C}}{1-\alpha}$, α is very small ;

$\therefore 1 - \alpha \approx 1$

$\mathrm{C}\alpha = 0.1$

$\mathrm{k_a} = (\mathrm{C}\alpha)\alpha$

$= 0.01 \times 0.1 = 10^{-3}$

Q14: (B) 5×10^{-7}

$= 0.05$ g/100 mL

$= 0.5$ g/L

$= \frac{0.5}{100} = 5 \times 10^{-3}$ mol/L

$\mathrm{K_{sp}} = 4\mathrm{s}^3 = 4 \times \left(5 \times 10^{-3}\right)^3 = 5 \times 10^{-7}$

Q15: 5.5

$$\begin{array}{ccccc} \mathrm{HA} & \rightleftharpoons & \mathrm{H^+} & + & \mathrm{A^-} \\ \mathrm{C_1} & & - & & - \\ \mathrm{C_1 - C_1\alpha_1} & & \mathrm{C_1\alpha_1} & & \mathrm{C_1\alpha_1} \end{array}$$

$\mathrm{pH} = 2 \Rightarrow [\mathrm{H^+}]_1 = 10^{-2} = \mathrm{c_1}\alpha_1 = 1 \times \alpha_1$

$\alpha_1 = 10^{-2}$

$\mathrm{pH} = 4 \Rightarrow [\mathrm{H^+}]_2 = 10^{-4} = \frac{1 \times 1000}{\mathrm{V}} \times \alpha_2$

$\mathrm{k_a} = 10^{-4} = \frac{\alpha_2^2 \mathrm{C_2}}{1-\alpha_2} = \frac{\alpha_2 \times \alpha_2 \times \mathrm{C_2}}{1-\alpha_2}$

$10^{-4} = \frac{\alpha_2 \times 10^{-4}}{1-\alpha_2} \Rightarrow \alpha_2 = 0\ .5$

$[\mathrm{H^+}]_2 = \frac{1 \times 1000}{\mathrm{V}} \times \alpha_2 = 10^{-4}$

$\frac{1 \times 1000}{\mathrm{V}} \times 0\ .5 = 10^{-4}$

$\mathrm{V} = 5 \times 10^{6}$ ml

$\Rightarrow \frac{\mathrm{x+y}}{2} = \frac{5+6}{2} = \frac{11}{2} = 5\ .5$

Q16: (A) $\mathrm{HNO_2}$;pka= 3.0

$\left[\mathrm{H^+}\right] = \sqrt{\mathrm{k_a} \times \mathrm{c}}$

$\mathrm{k_a} \uparrow \mathrm{pka} \downarrow \left[\mathrm{H^+}\right] \uparrow$

Q17: (C) 100 cc (N/100) HCl **is added to it**

If two solutions of same pH are mixed, no change in pH is observed. pH of water will be 7, that of 100 cc of 0.1 M HCl and 1 cc of 0.1 M HCl will be 1 (as H^+ concentration is 0.1)
0.01 N HCl has also pH = 2.
So adding this to a 1000 cc solution of pH 2 will not make any difference in pH

Q18: (C) 7

$$\underset{1\text{ m mol}}{CH_3COOH\ (aq)} + \underset{1\text{ m mol}}{NH_3\ (aq)} \longrightarrow \underset{1\text{ m mol}}{CH_3COONH_4\ (aq)}$$

$$\left[CH_3COONH_4\right] = \frac{1\text{ mM}}{20\text{ mL}} = 0.05\text{ M}$$

Salt of weak Acid + weak base.

$$pH = \frac{1}{2}\left(pK_w + pK_a - pK_b\right)$$
$$= \frac{1}{2}(14 + 5 - 5) = 7$$

Q19: (C) 7.01

$$BOH + HA \rightarrow B^+ + A^- + H_2O$$
$$pH = \frac{1}{2}pK_w + \frac{1}{2}pK_a - \frac{1}{2}pK_b$$
$$= \frac{1}{2}(14 + 4.80 - 4.78) = 7.01$$

Q20: (C) 392

$$K_a = C\alpha^2$$
$$25 \times 10^{-6} = (0.01)\,\alpha^2$$
$$\alpha = 5 \times 10^{-2}$$
$$\alpha = \frac{\Lambda_{eq}}{\Lambda^o_{eq}}$$
$$5 \times 10^{-2} = \frac{19.6}{\Lambda^o_{eq}}$$
$$\Lambda^o_{eq} = \frac{19.6}{5\times10^{-2}} = 392$$

Q21: (A) greater than 7

Na_2CO_3 is the salt of strong base and weak acid.

$$Na_2CO_3 + 2H_2O \rightarrow \underset{\text{Strong base}}{2NaOH} + \underset{\text{Weak acid}}{H_2CO_3}$$

Hence, the solution of Na_2CO_3 has pH greater than 7.

Q22: (B) 8.0×10^{-12}

$$\Delta G^o = -2.303\ RT\ \log K_{sp}$$
$$63.3 \times 1000 = -2.303 \times 8.314 \times 298 \log K_{sp}$$
$$\log K_{sp} = -11.09$$
$$K_{sp} = 10^{-11.09} = 8 \times 10^{-12}$$

Q23: (C) 13.2041

$$S = \frac{19.23}{192.3}\text{mol/litre} = 0.1\text{ M}$$

$$A(OH)_2 \rightleftharpoons A^{2+} + \underset{2S}{2OH^-}$$

$\therefore \left[OH^-\right] = \frac{0.1\times2\times80}{100} = 0.16$

$\therefore$ pOH $= 0.7958$

$\therefore$ pH $= 13.2041$

Q24: (B) 1.26

Sol −I	$V_1 = 50\,ml$	$C_1 = 10^{-1}M$
Sol −II	$V_2 = 50\,ml$	$C_2 = 10^{-2}M$

$V_f = 100$ ml

$C_f = \frac{10^{-1}\times50+10^{-2}\times50}{100} = 0.055M$

pH $= -\log(0.055) = 1.26$

Q25: (C) 1.0

$\alpha = \sqrt{\frac{10^{-3}}{10}} = 10^{-2}$

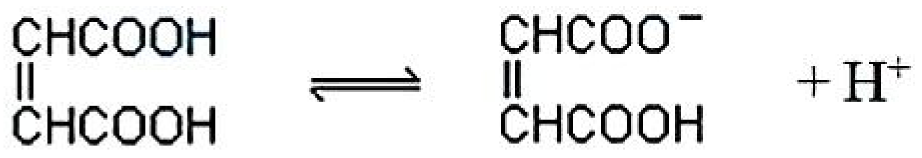
$\%\alpha = 10^{-2} \times 100 = 1\%$

Q26: (B) 2.3×10^{-4}

Degree of dissociation, $\alpha = \frac{14}{100} = 0.14$

Concentration of maleic acid, C $= 0.010$ M

The equilibrium reaction is



$$\begin{matrix} CHCOOH \\ \| \\ CHCOOH \end{matrix} \rightleftharpoons \begin{matrix} CHCOO^- \\ \| \\ CHCOOH \end{matrix} + H^+$$

$t = 0$	C	0	0
t_{eq}	$C(1-\alpha)$	$C\alpha$	$C\alpha$

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

$K_a = \frac{0.010\times0.14\times0.14}{0.86} = 2.3 \times 10^{-4}$.

Q27: (A) 10^{-4} M Ag^+ **and** 10^{-4} M Cl^-

After mixing,

$\left[Ag^+\right] = \frac{1}{2} \times 10^{-4} = 5 \times 10^{-5}$ M

$[Cl^-] = \frac{1}{2} \times 10^{-4} = 5 \times 10^{-5}$ M

$Q_{sp} = \left[Ag^{+}\right]\left[Cl^{-}\right] = \left(5 \times 10^{-5}\right)^{2} = 2.5 \times 10^{-9}$

Since, $Q_{sp} > K_{sp} \rightarrow$ precipition takes place

Q28: (A) 2:1

$$\begin{array}{llll} NH_4OH & + & HCl \rightleftharpoons & NH_4Cl + H_2O \\ 300 - x & & x\,mL & \\ (300 - x) - x & & x - x & x \end{array}$$

$pOH = 14 - pH = pK_b + \log \frac{x}{300-2x}$

$= 14 - 9.26 = -\log\left(10^{-4.74}\right) + \log \frac{x}{300-2x}$

$\Rightarrow \frac{x}{300-2x} = 1$

$\Rightarrow 300 = 3x \Rightarrow x = 100\text{ mL}$

$\frac{V_{NH_4OH}}{V_{HCl}} = 2$

Q29: (D) 8.95

$\underset{2}{NaOH} + \underset{2}{HA} \rightarrow Na\,A + H_2O$

At end point $\equiv 0.1 \times 20 = 2$

$\because$ 20 mL NaOH is required for the complete neutralization of HA

NaA is a salt of strong base and weak acid

Thus, will undergo hydrolysis and solution will becomes basic

$C = \left[NaA\right] = \frac{2}{20+20} = 0.05\text{ M}$

And $pK_a = -\log\left(6 \times 10^{-6}\right) = 5.2$

pH at the end point $= 7 + \frac{1}{2}\left(pK_a + \log C\right)$

$7 + \frac{1}{2}\left(5.2 + \log 0.05\right) = 8.95$

Q30: (D) 25

By $\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{10^{-4}}{1.2\times10^{-3}}} = \sqrt{\frac{1}{12}}$

$\%\,\alpha = 28.86\%$, as value of α is greater than 5% so we can not take assumptions and will have to use the exact formula.

So α will be given by:

$\alpha = \frac{-10^{-4}+\sqrt{10^{-8}+4\times10^{-3}\times1.2\times10^{-4}}}{2\times1.2\times10^{-3}}$

So the correct value of $\alpha = 25\,\%$

Q31: (C) 5

$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{0.1\times\left(10^{-2}\right)^2}{(1-0.01)} = 1.01 \times 10^{-5}$

$pK_a = -\log\left(1 \times 10^{-5}\right) = 5$

Q32: 5

First there will be an acid-base reaction between KH_2BO_3 and HCl .

	KH_2BO_3	$+ HCl \longrightarrow$	H_3BO_3	$+ KCl$
initially	0.1	0.1	0	0
after rxn	0	0	0.1	

conc. will become half after reaction of an equal volume of KH_2BO_3 and HCl

So $\frac{0.1}{5} = 0.05$

Now,

$$\begin{array}{llll} H_3BO_3 & +H_2O \longrightarrow H_4BO_4^- & +H^+ \\ 0.05(1-\alpha) & c\alpha & c\alpha \end{array}$$

$[H^+] = \sqrt{K_aC} = \sqrt{12.8 \times 10^{-10} \times 0.05} = 8 \times 10^{-6}$

$pH = -\log\left(8 \times 10^{-6}\right) \approx 5.0$

Q33: (C) 4.44

	$NaOH$	+	CH_3COOH	$\rightarrow$	CH_3COONa	+	H_2O
m mol	2		6				
	LR						
	–		4		2		

The resulting solution is an acidic buffer. so $pH = pK_a + \log\frac{[Salt]}{[Acid]}$

$= 4.71 + \log\frac{2}{4}$

$= 4.44$

Q34: 2.7

As the volume is doubled, concentration would be half. The total $\left[H^+\right]$ is obtained by addition of $\left[H^+\right]$ from both acids.

$[H^+] = \sqrt{(K_{a1}C_1 + K_{a2}C_2)}$

$= \sqrt{\left(2 \times 10^{-5} \times \frac{0.1}{2} + 6 \times 10^{-5} \times \frac{0.1}{2}\right)}$

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

$pH = 2.7$

Q35: 7

Let the solubility of AgCl is x mol litre^{-1}

$AgCl \rightleftharpoons \underset{x}{Ag^+} + \underset{x}{Cl^-}$

And that of CuCl is y mol litre^{-1}

$CuCl \rightleftharpoons \underset{y}{Cu^+} + \underset{y}{Cl^-}$

K_{sp} of AgCl $= \left[Ag^+\right]\left[Cl^{-1}\right]$

$1.6 \times 10^{-10} = x\,(x+y)$ ...(i)

Similarly K_{sp} of CuCl $= \left[Cu^+\right]\left[Cl^-\right]$

$1.6 \times 10^{-6} = y\,(x+y)$ ...(ii)

On solving (i) and (ii)

$\left[Ag^+\right] = 1.6 \times 10^{-7}$

$x = 7$

Q36: (B) $HCl < NH_4Cl < NaCl < NaCN$
pH pf $NaCl = 7$; pH of $NH_4Cl < 7$;

pH of $NaCN > 7$and pH of $0.1\ (M)HCl = 1$

Q37: (B) $HCl < NH_4Cl < NaCl < NaCN$
HCl: Strong acid (low pH)
NH_4Cl : Salt of weak base + strong acid → it gives acidic solution. (pH less than 7)
NaCl: Salt of strong base + strong base → it gives neutral solution with pH = 7.
NaCl: Salt of strong base + weak acid → it gives basic solution with pH > 7.
Thus the increasing order of pH is
$HCl < NH_4Cl < NaCl < NaCN$

Q38: 3

Milli.moles of $NaOH = 30 \times 0.2 = 6$

m.m of $CH_3COOH = 50 \times 0.2 = 10$

Now:

$$\begin{array}{cccc} CH_3COOH & + \quad NaOH & \longrightarrow CH_3COONa + H_2O \\ 10 & 6 & 0 \\ 10-6=4 & 6-6=0 & 6 \end{array}$$

$pH = -\log(2 \times 10^{-5}) + \log \frac{6}{4}$

$= 4.87$

Suppose 'v'mL of NaOH is added then

m.mol of $CH_3COONa = (6 + v \times 0.2)$

m.mol of $CH_3COOH = (4 - v \times 0.2)$

$pH = pK_a + \log\left[\frac{\text{Salt}}{\text{Acid}}\right]$
$5 = -\log(2 \times 10^{-5}) + \log\left(\frac{6+0.2v}{4-0.2v}\right)$

$0.3010 = \log\left(\frac{6+0.2v}{4-0.2v}\right)$

So $\left(\frac{6+0.2v}{4-0.2v}\right) = 2 \Rightarrow v = 3.33\ mL = \frac{10}{3} mL$

So $x = 3$

Q39: (C) $1.25 \times 10^{-4} M$

When $BaSO_4$ begins to precipitate

$\left[SO_4^{2-}\right] = \frac{K_{sp}[BaSO_4]}{[Ba^{2+}]} = \frac{1.0\times10^{-10}}{0.020}$

$= 5.0 \times 10^{-9}\ M$

When $PbSO_4$ begins to precipitate

$\left[SO_4^{2-}\right] = \frac{K_{sp}(PbSO_4)}{[Pb^{2+}]} = \frac{1.6\times10^{-8}}{0.020}$

$= 8.0 \times 10^{-7}\ M$

$\left[SO_4^{2-}\right]$ is less for $BaSO_4$ precipitation so $BaSO_4$ precipitates first when $PbSO_4$ begins to precipitate $\left[SO_4^{2-}\right] = 8.0 \times 10^{-7} M$ at that point concentration of $Ba^{2+} = \frac{K_{sp}(BaSO_4)}{[SO_4^{2-}]}$

$= \frac{1.0\times10^{-10}}{8.0\times10^{-7}} = 1.25 \times 10^{-4}\ M$

Q40: (C) 16:81

$K_a \propto c\,\alpha^2$

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

$= \frac{4^2}{9^2} = \frac{16}{81}$

Q41: (D) Sodium acetate and sodium propionate

Sodium acetate and sodium propionate are not conjugate acid base pairs

Q42: 8

$\overset{\oplus}{C}H_3, Cl^{\oplus}, CO_2, CCl_2, BCl_3, BI_3, Fe^{+2}, AlCl_3$

Q43: (B) 11.5

pk_b of $CN^- = 4.7$

pk_a of $HCN = 4.7$

NaCN is salt of weak acid (HCN)and strong base (NaOH) So

$pH = 7 + \frac{1}{2}pk_a + \frac{1}{2}\log C$

$= 7 + \frac{1}{2} \times 9.3 + \frac{1}{2}\log(0.5)$

$\therefore pH = 11.5$