Question 7.41:

The concentration of hydrogen ion in a sample of soft drink is 3.8 × 10^{-3} M. what is its pH?

Answer 7.41:

Given,

$\left[H^+\right] = 3.8 \times 10^{-3}$ M

$\therefore$ pH value of soft drink

$$= -\log\left[H^+\right]$$
$$= -\log\left(3.8 \times 10^{-3}\right)$$
$$= -\log 3.8 - \log 10^{-3}$$
$$= -\log 3.8 + 3$$
$$= -0.58 + 3$$
$$= 2.42$$

Question 7.42:

The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer 7.42:

Given, pH

= 3.76

It is known that,

$$\begin{aligned} \mathrm{pH} &= -\log\left[H^+\right] \\ \Rightarrow \log\left[H^+\right] &= -\mathrm{pH} \\ \Rightarrow \left[H^+\right] &= \text{antilog}(-\mathrm{pH}) \\ &= \text{antilog}(-3.76) \\ &= 1.74 \times 10^{-4}\ \mathrm{M} \end{aligned}$$

Hence, the concentration of hydrogen ion in the given sample of vinegar is 1.74 × 10^{-4} M.

Question 7.43:

The ionization constant of HF, HCOOH and HCN at 298K are 6.8×10^{-4}, 1.8×10^{-4} and 4.8×10^{-9} respectively. Calculate the ionization constants of the corresponding conjugate base.

Answer 7.43:

It is known that,

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

Given,
K_a of HF = 6.8×10^{-4}

Hence, K_b of its conjugate base F^-

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

$$= \frac{10^{-14}}{6.8 \times 10^{-4}}$$
$$= 1.5 \times 10^{-11}$$

Given,

K_a of HCOOH = 1.8×10^{-4}
Hence, K_b of its conjugate base $HCOO^-$

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

$$= \frac{10^{-14}}{1.8 \times 10^{-4}}$$
$$= 5.6 \times 10^{-11}$$

Given,

K_a of HCN = 4.8×10^{-9}
Hence, K_b of its conjugate base CN^-

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

$$= \frac{10^{-14}}{4.8 \times 10^{-9}}$$
$$= 2.08 \times 10^{-6}$$

Question 7.44:

The ionization constant of phenol is 1.0 × 10^{-10}. What is the concentration of phenolate ion in 0.05 M solution of phenol? What will be its degree of ionization if the solution is also 0.01M in sodium phenolate?

Answer 7.44:

Ionization of phenol:

$$\begin{array}{lcccc} & C_6H_5OH + H_2O & \longleftrightarrow & C_6H_5O^- & + H_3O^+ \\ \text{Initial conc.} & 0.05 & & 0 & 0 \\ \text{At equilibrium} & 0.05 - x & & x & x \end{array}$$

$$K_a = \frac{\left[C_6H_5O^-\right]\left[H_3O^+\right]}{\left[C_6H_5OH\right]}$$

$$K_a = \frac{x \times x}{0.05 - x}$$

As the value of the ionization constant is very less, x will be very small. Thus, we can ignore x in the denominator.

$$\therefore x = \sqrt{1 \times 10^{-10} \times 0.05}$$

$$= \sqrt{5 \times 10^{-12}}$$

$$= 2.2 \times 10^{-6}\,\text{M} = \left[H_3O^+\right]$$

Since $\left[H_3O^+\right] = \left[C_6H_5O^-\right]$,

$$\left[C_6H_5O^-\right] = 2.2 \times 10^{-6}\,\text{M}.$$

Now, let ∝ be the degree of ionization of phenol in the presence of 0.01 M C_6H_5ONa.

$$\begin{array}{lccc} & C_6H_5ONa & \longrightarrow C_6H_5O^- & + Na^+ \\ \text{Conc.} & & 0.01 & \end{array}$$

Also,

$$\begin{array}{lcccccc} & C_6H_5OH & + & H_2O & \longleftrightarrow & C_6H_5O^- & + \; H_3O^+ \\ \text{Conc.} & 0.05 - 0.05\alpha & & & & 0.05\alpha & 0.05\alpha \end{array}$$

$$[C_6H_5OH] = 0.05 - 0.05\alpha \;;\; 0.05\text{ M}$$

$$[C_6H_5O^-] = 0.01 + 0.05\alpha \;;\; 0.01\text{M}$$

$$[H_3O^+] = 0.05\alpha$$

$$K_a = \frac{[C_6H_5O^-][H_3O^+]}{[C_6H_5OH]}$$

$$K_a = \frac{(0.01)(0.05\alpha)}{0.05}$$

$$1.0\times10^{-10} = .01\alpha$$

$$\alpha = 1\times10^{-8}$$

Question 7.45:

The first ionization constant of H_2S is 9.1×10^{-8}. Calculate the concentration of HS^- ion in its 0.1 M solution. How will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of H_2S is 1.2×10^{-13}, calculate the concentration of S^{2-} under both conditions.

Answer 7.45:

(i) To calculate the concentration of HS^- ion:

Case I (in the absence of HCl):

Let the concentration of HS^- be x M.

$$\begin{array}{llll} & H_2S \longleftrightarrow & H^+ + & HS^- \\ C_i & 0.1 & 0 & 0 \\ C_f & 0.1-x & x & x \end{array}$$

$$\text{Then, } K_{a_1} = \frac{\left[H^+\right]\left[HS^-\right]}{\left[H_2S\right]}$$

$$9.1\times10^{-8} = \frac{(x)(x)}{0.1-x}$$

$$\left(9.1\times10^{-8}\right)(0.1-x) = x^2$$

Taking $0.1-x\,\text{M} \;;\; 0.1\text{M}$, we have $\left(9.1\times10^{-8}\right)(0.1) = x^2$.

$$9.1\times10^{-9} = x^2$$

$$x = \sqrt{9.1\times10^{-9}}$$

$$= 9.54\times10^{-5}\,\text{M}$$

$$\Rightarrow \left[HS^-\right] = 9.54\times10^{-5}\,\text{M}$$

Case II (in the presence of HCl):

In the presence of 0.1 M of HCl, let $\left[HS^-\right]$ be y M

$$\begin{array}{llll} \text{Then,} & H_2S \longleftrightarrow & HS^- + & H^+ \\ C_i & 0.1 & 0 & 0 \\ C_f & 0.1-y & y & y \end{array}$$

$$\text{Now, } K_{a_1} = \frac{\left[HS^-\right]\left[H^+\right]}{\left[H_2S\right]}$$

$$K_{a_1} = \frac{[y](0.1+y)}{(0.1-y)}$$

$$9.1\times10^{-8} = \frac{y\times0.1}{0.1} \qquad (\because 0.1-y \;;\; 0.1\text{M})$$

$$(\text{and } 0.1+y \;;\; 0.1\text{M})$$

$$9.1\times10^{-8} = y$$

$$\Rightarrow \left[HS^-\right] = 9.1\times10^{-8}$$

To calculate the concentration of $\left[S^{2-}\right]$

Case I (in the absence of 0.1 M HCl):

$HS^{-} \longleftrightarrow H^{+} + S^{2-}$

$\left[HS^{-}\right] = 9.54\times10^{-5}\,M$ (From first ionization, case I)

Let $\left[S^{2-}\right]$ be X.

Also, $\left[H^{+}\right] = 9.54\times10^{-5}\,M$ (From first ionization, case I)

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

$$K_{a_2} = \frac{\left(9.54\times10^{-5}\right)(X)}{9.54\times10^{-5}}$$

$$1.2\times10^{-13} = X = \left[S^{2-}\right]$$

Case II (in the presence of 0.1 M HCl):

Again, let the concentration of HS^{-} be X' M.

$\left[HS^{-}\right] = 9.1\times10^{-8}\,M$ (From first ionization, case II)

$\left[H^{+}\right] = 0.1\,M$ (From HCl, case II)

$$\left[S^{2-}\right] = X'$$

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

$$1.2\times10^{-13} = \frac{(0.1)(X')}{9.1\times10^{-8}}$$

$$10.92\times10^{-21} = 0.1X'$$

$$\frac{10.92\times10^{-21}}{0.1} = X'$$

$$X' = \frac{1.092\times10^{-20}}{0.1}$$

$$= 1.092\times10^{-19}\text{ M}$$

$$\Rightarrow K_{a_1} = 1.74\times10^{-5}$$

Question 7.46:

The ionization constant of acetic acid is 1.74×10^{-5}. Calculate the degree of dissociation of acetic acid in its 0.05 M solution. Calculate the concentration of acetate ion in the solution and its pH.

Answer 7.46:

Method 1

1) $CH_3COOH \longleftrightarrow CH_3COO^{-} + H^{+}$ $K_a = 1.74\times10^{-5}$

2) $H_2O + H_2O \longleftrightarrow H_3O^{+} + OH^{-}$ $K_w = 1.0\times10^{-14}$

Since $K_a >> K_{w,}$:

$$\begin{array}{lllll} & CH_3COOH + H_2O & \longleftrightarrow & CH_3COO^{-} & + H_3O^{+} \\ C_i = & 0.05 & & 0 & 0 \\ & 0.05 - .05\alpha & & 0.05\alpha & 0.05\alpha \end{array}$$

$$K_a = \frac{(.05\alpha)(.05\alpha)}{(.05 - 0.05\alpha)}$$

$$= \frac{(.05\alpha)(0.05\alpha)}{.05(1-\alpha)}$$

$$= \frac{.05\alpha^2}{1-\alpha}$$

$$1.74\times10^{-5}=\frac{0.05\alpha^2}{1-\alpha}$$

$$1.74\times10^{-5}-1.74\times10^{-5}\alpha=0.05\alpha^2$$

$$0.05\alpha^2+1.74\times10^{-5}\alpha-1.74\times10^{-5}$$

$$D=b^2-4ac$$

$$=\left(1.74\times10^{-5}\right)^2-4(.05)\left(1.74\times10^{-5}\right)$$

$$=3.02\times10^{-25}+.348\times10^{-5}$$

$$\alpha=\sqrt{\frac{K_a}{c}}$$

$$\alpha=\sqrt{\frac{1.74\times10^{-5}}{.05}}$$

$$=\sqrt{\frac{34.8\times10^{-5}\times10}{10}}$$

$$=\sqrt{3.48\times10^{-6}}$$

$$=CH_3COOH\longleftrightarrow CH_3COO^-+H^+$$

$$\underline{\alpha1.86\times10^{-3}}$$

$$\left[CH_3COO^-\right]=0.05\times1.86\times10^{-3}$$

$$=\frac{0.93\times10^{-3}}{1000}$$

$$=\underline{.000093}$$

Method 2

Degree of dissociation,

$$\alpha=\sqrt{\frac{K_a}{c}}$$

c = 0.05 M

K_a = 1.74 × 10^{-5}

$$\text{Then, } \alpha = \sqrt{\frac{1.74\times10^{-5}}{.05}}$$

$$\alpha = \sqrt{34.8\times10^{-5}}$$

$$\alpha = \sqrt{3.48}\times10^{-4}$$

$$\alpha = 1.8610^{-2}$$

$$CH_3COOH \longleftrightarrow CH_3COO^- + H^+$$

Thus, concentration of CH_3COO^- = c.α

$$= .05\times1.86\times10^{-2}$$

$$= .093\times10^{-2}$$

$$= .00093\,M$$

$$\text{Since } \left[oAc^-\right] = \left[H^+\right],$$

$$\left[H^+\right] = .00093 = .093\times10^{-2}.$$

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

$$= -\log\left(.093\times10^{-2}\right)$$

$$\therefore pH = 3.03$$

Hence, the concentration of acetate ion in the solution is 0.00093 M and its Ph is 3.03.

Question 7.47:

It has been found that the pH of a 0.01M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid and its pK_a.

Answer 7.47:

Let the organic acid be HA.

$$\Rightarrow HA \longleftrightarrow H^+ + A^-$$

Concentration of HA = 0.01 M pH

= 4.15

$-\log\left[H^+\right]=4.15$

$\left[H^+\right]=7.08\times10^{-5}$

Now, $K_a=\dfrac{\left[H^+\right]\left[A^-\right]}{[HA]}$

$\left[H^+\right]=\left[A^-\right]=7.08\times10^{-5}$

$[HA]=0.01$

Then,

$$K_a=\frac{\left(7.08\times10^{-5}\right)\left(7.08\times10^{-5}\right)}{0.01}$$

$K_a=5.01\times10^{-7}$

$pK_a=-\log K_a$

$=-\log\left(5.01\times10^{-7}\right)$

$pK_a=6.3001$

Question 7.48:

Assuming complete dissociation, calculate the pH of the following solutions:

(i) 0.003 M HCl

(ii) 0.005 M NaOH

(iii) 0.002 M HBr

(iv) 0.002 M KOH

Answer 7.48:

(i) 0.003MHCl:

$H_2O+HCl\longleftrightarrow H_3O^++Cl^-$

Since HCl is completely ionized,

$\left[H_3O^+\right]=[HCl].$

$\Rightarrow\left[H_3O^+\right]=0.003$

Now,

$$\mathrm{pH} = -\log\left[\mathrm{H_3O^+}\right] = -\log(.003)$$
$$= 2.52$$

Hence, the pH of the solution is 2.52**.**

(ii) 0.005MNaOH:

$$\mathrm{NaOH}_{(aq)} \longleftrightarrow \mathrm{Na}^+_{(aq)} + \mathrm{HO}^-_{(aq)}$$
$$\left[\mathrm{HO^-}\right] = [\mathrm{NaOH}]$$
$$\Rightarrow \left[\mathrm{HO^-}\right] = .005$$
$$\mathrm{pOH} = -\log\left[\mathrm{HO^-}\right] = -\log(.005)$$
$$\mathrm{pOH} = 2.30$$
$$\therefore \mathrm{pH} = 14 - 2.30$$
$$= 11.70$$

Hence, the pH of the solution is 11.70.

(iii) 0.002 HBr:

$$\mathrm{HBr} + \mathrm{H_2O} \longleftrightarrow \mathrm{H_3O^+} + \mathrm{Br^-}$$
$$\left[\mathrm{H_3O^+}\right] = [\mathrm{HBr}]$$
$$\Rightarrow \left[\mathrm{H_3O^+}\right] = .002$$
$$\therefore \mathrm{pH} = -\log\left[\mathrm{H_3O^+}\right]$$
$$= -\log(0.002)$$
$$= 2.69$$

Hence, the pH of the solution is 2.69.

(iv) 0.002 M KOH:

$$\mathrm{KOH}_{(aq)} \longleftrightarrow \mathrm{K}^+_{(aq)} + \mathrm{OH}^-_{(aq)}$$
$$\left[\mathrm{OH^-}\right] = [\mathrm{KOH}]$$
$$\Rightarrow \left[\mathrm{OH^-}\right] = .002$$
$$\text{Now, } \mathrm{pOH} = -\log\left[\mathrm{OH^-}\right]$$
$$= 2.69$$
$$\therefore \mathrm{pH} = 14 - 2.69$$
$$= 11.31$$

Hence, the pH of the solution is 11.31.

Question 7.49:

Calculate the pH of the following solutions:

a) 2 g of TlOH dissolved in water to give 2 litre of solution.

b) 0.3 g of $Ca(OH)_2$ dissolved in water to give 500 mL of solution.

c) 0.3 g of NaOH dissolved in water to give 200 mL of solution.
d) 1mL of 13.6 M HCl is diluted with water to give 1 litre of solution.

Answer 7.49:

(a) For 2g of TlOH dissolved in water to give 2 L of solution:

$$\left[\mathrm{TlOH}_{(aq)}\right]=\frac{2}{2}\,\mathrm{g/L}$$
$$=\frac{2}{2}\times\frac{1}{221}\,\mathrm{M}$$
$$=\frac{1}{221}\,\mathrm{M}$$
$$\mathrm{TlOH}_{(aq)}\longrightarrow \mathrm{Tl}^{+}_{(aq)}+\mathrm{OH}^{-}_{(aq)}$$
$$\left[\mathrm{OH}^{-}_{(aq)}\right]=\left[\mathrm{TlOH}_{(aq)}\right]=\frac{1}{221}\,\mathrm{M}$$
$$K_w=\left[\mathrm{H}^{+}\right]\left[\mathrm{OH}^{-}\right]$$
$$10^{-14}=\left[\mathrm{H}^{+}\right]\left(\frac{1}{221}\right)$$
$$221\times10^{-14}=\left[\mathrm{H}^{+}\right]$$
$$\Rightarrow \mathrm{pH}=-\log\left[\mathrm{H}^{+}\right]=-\log\left(221\times10^{-14}\right)$$
$$=-\log\left(2.21\times10^{-12}\right)$$
$$=11.65$$

(b) For 0.3 g of $Ca(OH)_2$ dissolved in water to give 500 mL of solution:

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

$$\left[Ca(OH)_2\right] = 0.3 \times \frac{1000}{500} = 0.6M$$

$$\left[OH^-_{aq}\right] = 2 \times \left[Ca(OH)_{2aq}\right] = 2 \times 0.6$$

$$= 1.2M$$

$$\left[H^+\right] = \frac{K_w}{\left[OH^-_{aq}\right]}$$

$$= \frac{10-14}{1.2}M$$

$$= 0.833 \times 10^{-14}$$

$$pH = -\log\left(0.833 \times 10^{-14}\right)$$

$$= -\log\left(8.33 \times 10^{-13}\right)$$

$$= (-0.902 + 13)$$

$$= 12.098$$

(c) For 0.3 g of NaOH dissolved in water to give 200 mL of solution:

$$NaOH \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$\left[NaOH\right] = 0.3 \times \frac{1000}{200} = 1.5M$$

$$\left[OH^-_{aq}\right] = 1.5M$$

$$\text{Then,} \left[H^+\right] = \frac{10^{-14}}{1.5}$$

$$= 6.66 \times 10^{-13}$$

$$pH = -\log\left(6.66 \times 10^{-13}\right)$$

$$= 12.18$$

(d) For 1mL of 13.6 M HCl diluted with water to give 1 L of solution:

13.6 × 1 mL = M_2 × 1000 mL

(Before dilution) (After dilution)

$13.6 \times 10^{-3} = M_2 \times 1L$ M_2

$= 1.36 \times 10^{-2}$ $[H^+] = 1.36 \times 10^{-2}$ pH = – log (1.36×10^{-2})

= (– 0.1335 + 2) = 1.866 … 1.87

Question 7.50:

The degree of ionization of a 0.1M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pK_a of bromoacetic acid.

Answer 7.50:

Degree of ionization, α = 0.132

Concentration, *c* = 0.1 M

Thus, the concentration of H_3O^+ = $c.\alpha$
= 0.1 × 0.132

= 0.0132

$$\begin{aligned}\mathrm{pH} &= -\log\left[\mathrm{H}^+\right]\\ &= -\log(0.0132)\\ &= 1.879 : 1.88\end{aligned}$$

Now,

$$\begin{aligned}K_a &= C\alpha^2\\ &= 0.1\times(0.132)^2\\ K_a &= .0017\\ pK_a &= 2.75\end{aligned}$$

Question 7.51:

The degree of ionization of a 0.1M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pK_a of bromoacetic acid.

Answer 7.51:

Degree of ionization, α = 0.132

Concentration, *c* = 0.1 M

Thus, the concentration of H_3O^+ = $c.\alpha$

= 0.1 × 0.132

= 0.0132

$$\begin{aligned}\mathrm{pH} &= -\log\left[\mathrm{H}^+\right]\\ &= -\log(0.0132)\\ &= 1.879 : 1.88\end{aligned}$$

Now,

$$K_a = C\alpha^2$$
$$= 0.1 \times (0.132)^2$$
$$K_a = .0017$$
$$pK_a = 2.75$$

Question 7.52:

What is the pH of 0.001 M aniline solution? The ionization constant of aniline can be taken from Table 7.7. Calculate the degree of ionization of aniline in the solution. Also calculate the ionization constant of the conjugate acid of aniline.

Answer 7.52:

$K_b = 4.27 \times 10^{-10}$

c = 0.001M pH

=?

α =?

$$k_b = c\alpha^2$$
$$4.27 \times 10^{-10} = 0.001 \times \alpha^2$$
$$4270 \times 10^{-10} = \alpha^2$$
$$65.34 \times 10^{-5} = \alpha = 6.53 \times 10^{-4}$$
$$\text{Then, } [\text{anion}] = c\alpha = .001 \times 65.34 \times 10^{-5}$$
$$= .065 \times 10^{-5}$$
$$\text{pOH} = -\log(.065 \times 10^{-5})$$
$$= 6.187$$
$$\text{pH} = 7.813$$

Now,

$$K_a \times K_b = K_w$$
$$\therefore 4.27 \times 10^{-10} \times K_a = K_w$$
$$K_a = \frac{10^{-14}}{4.27 \times 10^{-10}}$$
$$= 2.34 \times 10^{-5}$$

Thus, the ionization constant of the conjugate acid of aniline is 2.34×10^{-5}.

Question 7.53:

Calculate the degree of ionization of 0.05M acetic acid if its pK_a value is 4.74.

How is the degree of dissociation affected when its solution also contains (a) 0.01 M (b) 0.1 M in HCl?

Answer 7.53:

$$c = 0.05\ \text{M}$$
$$pK_a = 4.74$$
$$pK_a = -\log(K_a)$$
$$K_a = 1.82 \times 10^{-5}$$
$$K_a = c\alpha^2 \qquad \alpha = \sqrt{\frac{K_a}{c}}$$
$$\alpha = \sqrt{\frac{1.82 \times 10^{-5}}{5 \times 10^{-2}}} = 1.908 \times 10^{-2}$$

When HCl is added to the solution, the concentration of H^+ ions will increase. Therefore, the equilibrium will shift in the backward direction i.e., dissociation of acetic acid will decrease.

Case I: When 0.01 M HCl is taken.

Let x be the amount of acetic acid dissociated after the addition of HCl.

	CH_3COOH	$\longleftrightarrow$	H^+	+	CH_3COO^-
Initial conc.	0.05M		0		0
After dissociation	$0.05 - x$		$0.01 + x$		x

As the dissociation of a very small amount of acetic acid will take place, the values i.e., $0.05 - x$ and $0.01 + x$ can be taken as 0.05 and 0.01 respectively.

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$\therefore K_a = \frac{(0.01)x}{0.05}$$

$$x = \frac{1.82\times10^{-5}\times0.05}{0.01}$$

$$x = 1.82\times10^{-3}\times0.05\,M$$

Now,

$$\alpha = \frac{\text{Amount of acid dissociated}}{\text{Amount of acid taken}}$$

$$= \frac{1.82\times10^{-3}\times0.05}{0.05}$$

$$= 1.82\times10^{-3}$$

Case II: When 0.1 M HCl is taken.

Let the amount of acetic acid dissociated in this case be *X*. As we have done in the first case, the concentrations of various species involved in the reaction are:

$$[CH_3COOH] = 0.05 - X \; ; \; 0.05\,M$$

$$[CH_3COO^-] = X$$

$$[H^+] = 0.1 + X \; ; \; 0.1\,M$$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$\therefore K_a = \frac{(0.1)X}{0.05}$$

$$x = \frac{1.82\times10^{-5}\times0.05}{0.1}$$

$$x = 1.82\times10^{-4}\times0.05\,M$$

Now,

$$\alpha = \frac{\text{Amount of acid dissociated}}{\text{Amount of acid taken}}$$

$$= \frac{1.82\times10^{-4}\times0.05}{0.05}$$

$$= 1.82\times10^{-4}$$

Question 7.54:

The ionization constant of dimethylamine is 5.4 × 10^{-4}. Calculate its degree of ionization in its 0.02 M solution. What percentage of dimethylamine is ionized if the solution is also 0.1 M in NaOH?

Answer 7.54:

$$K_b = 5.4 \times 10^{-4}$$
$$c = 0.02 \text{ M}$$
$$\text{Then, } \alpha = \sqrt{\frac{K_b}{c}}$$
$$= \sqrt{\frac{5.4 \times 10^{-4}}{0.02}}$$
$$= 0.1643$$

Now, if 0.1 M of NaOH is added to the solution, then NaOH (being a strong base) undergoes complete ionization.

$$\begin{array}{ccccc} NaOH_{(aq)} & \longleftrightarrow & Na^{+}_{(aq)} & + & OH^{-}_{(aq)} \\ & & 0.1\text{ M} & & 0.1\text{ M} \end{array}$$

And,

$$\begin{array}{ccccccc} (CH_3)_2NH & + & H_2O & \longleftrightarrow & (CH_3)_2NH_2^{+} & + & \overset{-}{O}H \\ (0.02-x) & & & & x & & x \\ ;\ 0.02\text{M} & & & & & & ;\ 0.1\text{M} \end{array}$$

$$\text{Then, } \left[(CH_3)_2NH_2^{+}\right] = x$$
$$\left[OH^{-}\right] = x + 0.1 \ ; \ 0.1$$
$$\Rightarrow K_b = \frac{\left[(CH_3)_2NH_2^{+}\right]\left[OH^{-}\right]}{\left[(CH_3)_2NH\right]}$$
$$5.4 \times 10^{-4} = \frac{x \times 0.1}{0.02}$$
$$x = 0.0054$$

It means that in the presence of 0.1 M NaOH, 0.54% of dimethylamine will get dissociated.

Question 7.55:

Calculate the hydrogen ion concentration in the following biological fluids whose pH are given below:

(a) Human muscle-fluid, 6.83

(b) Human stomach fluid, 1.2

(c) Human blood, 7.38

(d) Human saliva, 6.4.

Answer 7.55:

(a) Human muscle fluid 6.83:

pH = 6.83 pH = – log [H^+]

∴6.83 = – log [H^+]

[H^+] =1.48 × 10^{-7} M

(b) Human stomach fluid, 1.2:

pH =1.2

1.2 = – log [H^+]

∴[H^+] = 0.063

(c) Human blood, 7.38:

pH = 7.38 = – log [H^+]

∴ [H^+] = 4.17 × 10^{-8} M

(d) Human saliva, 6.4:

pH = 6.4

6.4 = – log [H^+]

[H^+] = 3.98 × 10^{-7}

Question 7.56:

The pH of milk, black coffee, tomato juice, lemon juice and egg white are 6.8, 5.0, 4.2, 2.2 and 7.8 respectively. Calculate corresponding hydrogen ion concentration in each.

Answer 7.56:

The hydrogen ion concentration in the given substances can be calculated by using the given relation: $\text{pH} = -\log [H^+]$

(i) pH of milk = 6.8

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

$6.8 = -\log [H^+] \log$

$[H^+] = -6.8$

$[H^+] = \text{anitlog}(-6.8)$

$= 1.5 \times 19^{-7}$ M

(ii) pH of black coffee = 5.0

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

$5.0 = -\log [H^+] \log$

$[H^+] = -5.0$

$[H^+] = \text{anitlog}(-5.0)$

$= 10^{-5}$ M

(iii) pH of tomato juice = 4.2

Since, $\text{pH} = -\log [H^+]$
$4.2 = -\log [H^+] \log$

$[H^+] = -4.2$

$[H^+] = \text{anitlog}(-4.2)$

$= 6.31 \times 10^{-5}$ M

(iv) pH of lemon juice = 2.2

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

$2.2 = -\log [H^+] \log$

$[H^+] = -2.2$

$[H^+]$ = anitlog(−2.2)
= 6.31×10^{-3} M

(v) pH of egg white = 7.8

Since, pH = −log $[H^+]$

7.8 = −log $[H^+]$ log

$[H^+]$ = −7.8

$[H^+]$ = anitlog(−7.8)
=1.58×10^{-8} M

Question 7.57:

If 0.561 g of KOH is dissolved in water to give 200 mL of solution at 298 K. Calculate the concentrations of potassium, hydrogen and hydroxyl ions. What is its pH?

Answer 7.57:

$$\left[KOH_{aq}\right]=\frac{0.561}{\frac{1}{5}}g/L$$

$$=2.805\,g/L$$

$$=2.805\times\frac{1}{56.11}\text{ M}$$

$$=.05\text{ M}$$

$$KOH_{(aq)}\rightarrow K^+{}_{(aq)}+OH^-{}_{(aq)}$$

$$\left[OH^-\right]=.05\text{ M}=\left[K^+\right]$$

$$\left[H^+\right]\left[H^-\right]=K_w$$

$$\left[H^+\right]\frac{K_w}{\left[OH^-\right]}$$

$$=\frac{10^{-14}}{0.05}=2\times10^{-13}\text{ M}$$

$$\therefore\ pH=12.70$$

Question 7.58:

The solubility of $Sr(OH)_2$ at 298 K is 19.23 g/L of solution. Calculate the concentrations of strontium and hydroxyl ions and the pH of the solution.

Answer 7.58:

Solubility of $Sr(OH)_2$ = 19.23 g/L

Then, concentration of $Sr(OH)_2$

$$= \frac{19.23}{121.63}\ \text{M}$$

$$= 0.1581\ \text{M}$$

$$\mathrm{Sr(OH)}_{2(aq)} \longrightarrow \mathrm{Sr}^{2+}_{(aq)} + 2\left(\mathrm{OH}^-\right)_{(aq)}$$

$$\therefore \left[\mathrm{Sr}^{2+}\right] = 0.1581\ \text{M}$$

$$\left[\mathrm{OH}^-\right] = 2 \times 0.1581\ \text{M} = 0.3126\ \text{M}$$

Now,

$$K_w = \left[\mathrm{OH}^-\right]\left[\mathrm{H}^+\right]$$

$$\frac{10^{-14}}{0.3126} = \left[\mathrm{H}^+\right]$$

$$\Rightarrow \left[\mathrm{H}^+\right] = 3.2 \times 10^{-14}$$

$$\therefore \text{pH} = 13.495\ ;\ 13.50$$

Question 7.59:

The ionization constant of propanoic acid is 1.32×10^{-5}. Calculate the degree of ionization of the acid in its 0.05M solution and also its pH. What will be its degree of ionization if the solution is 0.01M in HCl also?

Answer 7.59:

Let the degree of ionization of propanoic acid be α.

Then, representing propionic acid as HA, we have:

$$\begin{array}{lllll} \mathrm{HA} & + \ \mathrm{H_2O} & \longleftrightarrow & \mathrm{H_3O^+} & + \ \mathrm{A^-} \\ (.05-0.0\alpha)\approx .05 & & & .05\alpha & .05\alpha \end{array}$$

$$K_a = \frac{\left[\mathrm{H_3O^+}\right]\left[\mathrm{A^-}\right]}{[\mathrm{HA}]}$$

$$= \frac{(.05\alpha)(.05\alpha)}{0.05} = .05\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{.05}} = 1.63\times10^{-2}$$

$$\text{Then,}\left[\ \mathrm{H_3O^+}\right] = .05\alpha = .05\times1.63\times10^{-2} = K_b.15\times10^{-4}\ \mathrm{M}$$

$$\therefore\ \mathrm{pH} = 3.09$$

In the presence of 0.1M of HCl, let α´ be the degree of ionization.

$$\text{Then,}\left[\ \mathrm{H_3O^+}\right] = 0.01$$

$$\left[\mathrm{A^-}\right] = 005\alpha'$$

$$[\mathrm{HA}] = .05$$

$$K_a = \frac{0.01\times.05\alpha'}{.05}$$

$$1.32\times10^{-5} = .01\times\alpha'$$

$$\alpha' = 1.32\times10^{-3}$$

Question 7.60:

The pH of 0.1M solution of cyanic acid (HCNO) is 2.34. Calculate the ionization constant of the acid and its degree of ionization in the solution.

Answer 7.60:

c = 0.1 M

pH = 2.34

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

$$-\log\left[H^+\right] = 2.34$$

$$\left[H^+\right] = 4.5\times10^{-3}$$

Also,

$$\left[H^+\right] = c\alpha$$

$$4.5\times10^{-3} = 0.1\times\alpha$$

$$\frac{4.5\times10^{-3}}{0.1} = \alpha$$

$$\alpha = 45\times10^{-3} = .045$$

Then,

$$K_a = c\alpha^2$$

$$= 0.1\times\left(45\times10^{-3}\right)^2$$

$$= 202.5\times10^{-6}$$

$$= 2.02\times10^{-4}$$