Ionic Equilibria

Important Formulae

1) Degree of dissociation ($\propto$) = $\frac{\text{Percent dissociation}}{100}$

2) Ostwald's dilution law : K = $\propto^2$ c ; $\propto = \sqrt{\frac{K}{c}}$

3) pH = $-log_{10}[H^+]$; pOH = $-log_{10}[OH^-]$

4) $K_w = [H^+] \times [OH^-]$

5) pH + pOH = 14

6) Solubility Product : $K_{sp} = x^x . y^y . S^{x+y}$

Ionic Equilibria Numericals

Q. 1 i) The pH of 10^{-8} M of HCl is ……

Solution :

10^{-8} M of HCl , $[H^{+}]$ = 10^{-8} M

pH = $-log_{10}[H^{+}]$

= $-log_{10}10^{-8}$

= $-(-8)log_{10}10$ $\quad\because log_{10}a^{m}$= m $log_{10}a$

= 8 $\quad\because log_{a}a = 1$

Q. 2 iv) Dissociation consist of acid is 1.8×10^{-5}.

Calculate percent dissociation of acetic acid in 0.01 M

Solution :

Given, Ka = 1.8×10^{-5}, c = 0.01 M

$$\propto = \sqrt{\frac{Ka}{c}}$$

$$= \sqrt{\frac{1.8 \times 10^{-5}}{0.01}}$$

$$= \sqrt{\frac{1.8 \times 10^{-5}}{10^{-2}}}$$

$$\propto = \sqrt{1.8 \times 10^{-3}}$$

$$= \sqrt{18 \times 10^{-4}}$$

$$= 4.242 \times 10^{-2}$$

$\therefore$ Percent dissociation = $\propto \times 100$

$$= 4.242 \times 10^{-2} \times 10^{2}$$

$$= 4.242 \%$$

Q. 2 vi) The pH of solution is 6.06. Calculate its H^+ ion concentration.

Solution :

$$\text{pH} = -log_{10}[H^+]$$

$$log_{10}[H^+] = -\text{ pH}$$

$$= -6.06$$

$$= -6 + (-0.06)$$

$$= -6 + (-0.06) + 1 - 1$$

$$= -7 + 0.94$$

$$log_{10}[H^+] = \bar{7}.94$$

$$[H^+] = \text{antilog}\ (\bar{7}.94)$$

$$= 8.710 \times 10^{-7}$$

Q. 3 vi) Acetic acid is 5% ionized in its decimolar solution.

Calculate the dissociation constant of acid.

Solution :

$Ka = \propto^2 c$

Here, $\propto = \frac{\text{Percent dissociation}}{100}$

$\propto = \frac{5}{100} = 5 \times 10^{-2}$

$c = 0.1 \text{ M} = 10^{-1} \text{ M}$

$Ka = \propto^2 c$

$Ka = (5 \times 10^{-2})^2 \times 10^{-1}$

$= 25 \times 10^{-4} \times 10^{-1}$

$= 2.5 \times 10^{-4}$

Q. 3 ix) pH of weak monobasic acid is 3.2 in its 0.02 M solution.

Calculate its dissociation constant.

Solution :

$$\text{pH} = -log_{10}[H^+]$$

$$log_{10}[H^+] = -\text{ pH}$$

$$= -3.2$$

$$= -3 + (-0.2)$$

$$= -3 + (-0.2) + 1 - 1$$

$$log_{10}[H^+] = -4 + 0.8 == \bar{4}.8$$

$$[H^+] = \text{antilog}\,(\bar{4}.8)$$

$$= 6.310 \times 10^{-4}$$

$$\propto = \frac{[H^+]}{c}$$

$$= \frac{6.310 \times 10^{-4}}{0.02}$$

$$= 3.155 \times 10^{-2}$$

$$\text{Ka} = \propto^2 c$$

$$\text{Ka} = (3.155 \times 10^{-2})^2 \times 2 \times 10^{-2}$$

$$= 9.954 \times 10^{-4} \times 2 \times 10^{-2}$$

$$= 19.908 \times 10^{-6}$$

$$= 1.91 \times 10^{-5}$$

Q. 3 x) In NaOH solution $[OH^-]$ is 2.87×10^{-4}

Calculate the pH of solution.

Solution :

$$pOH = -log_{10}[OH^-]$$

$$= -log_{10}(2.87 \times 10^{-4})$$

$$= -[\, log_{10}2.87 + log_{10}10^{-4} \,]$$

$$= -\, [\, 0.4579 - 4\, log_{10}10 \,]$$

$$= -\, [\, 0.4579 - 4 \,]$$

$$= 4 - 0.4579$$

$$pOH = 3.5421$$

$$pOH = 3.54$$

Now, $PH = 14 - pOH$

$$= 14 - 3.54$$

$$= 10.46$$

Q. 4 ix) The pH of rain water collected in a certain region of Maharashtra on particular day was 5.1 . Calculate its H^+ ion concentration of the rain water and its percent dissociation.

Solution :

$$pH = -log_{10}[H^+]$$

$$log_{10}[H^+] = -\ pH$$

$$= -\ 5.1$$

$$= -\ 5 + (-\ 0.1)$$

$$= -\ 5 + (-\ 0.1) + 1 - 1$$

$$= -\ 6 + 0.9$$

$$log_{10}[H^+] = \bar{6}.9$$

$$[H^+] = \text{antilog}\ (\bar{6}.9)$$

$$= 9.542 \times\ 10^{-6}$$

$$[H^+] = 9.54 \times\ 10^{-6}$$