ALLEN
CAREER INSTITUTE
KOTA (RAJASTHAN)
Path to Success

# IONIC EQUILIBRIUM

## NEET SYLLABUS

***IONIC EQUILIBRIUM :*** *ionization of acids and bases, strong and weak electrolytes, degree of ionization, ionization of polybasic acids, acid strength, concept of pH, Hydrolysis of salts (elementary idea)., buffer solutions, Henderson equation, solubility product, common ion effect (with illustrative examples).*

## *OBJECTIVES*

***After studying this unit, you will be able to :***

- *Classify acids and bases as weak or strong in terms of their ionization constants;*
- *Explain the dependence of degree of ionization on concentration of the electrolyte and that of the common ion;*
- *Describe pH scale for representing hydrogen ion concentration;*
- *Explain ionisation of water and its dual role as acid and base;*
- *Describe ionic product (Kw) and pKw for water;*
- *Appreciate use of buffer solutions;*
- *Calculate solubility product constant.*
- *Classify substances as acids or bases according to Arrhenius, Bronsted-Lowry and Lewis concepts;*

*"If you're not part of the solution, you're part of the precipitate."*

**Henry J. Tillman**

# IONIC EQUILIBRIUM

## 4.0 INTRODUCTION

**Formulae :**

(i) $\ln x = \log_e x = 2.303 \log_{10} x = 2.303 \log x$

(ii) $\log (x \times y) = \log x + \log y$

(iii) $\log \left(\frac{x}{y}\right) = \log x - \log y$

(iv) $\log x^y = y \log x$

**Ex.** (i) $\log 6 = \log (2 \times 3)$

$= \log 2 + \log 3$

$= 0.3010 + 0.4771 = 0.7781$

(ii) $\log 30 = \log (3 \times 10)$

$= \log 3 + \log 10$

$= \log 0.4771 + 1 = 1.4771$

(iii) $\log 1000 = \log 10^3$

$= 3 \log 10 \quad = 3 \times 1 = 3$

**A. Some values of Log :**

$\log 1 = 0$
$\log 2 = 0.3010$
$\log 3 = 0.4771$
$\log 4 = 0.6020$
$\log 5 = 0.699$
$\log 6 = 0.7781$
$\log 7 = 0.8451$
$\log 8 = 0.9030$
$\log 9 = 0.9542$
$\log 10 = 1$
$\log 11 = 1.04$
$\log 100 = 2$
$\log 1000 = 3$

**Some terms which are used in ionic equilibrium :** **Antilog** : Antilog (x) = $10^x$

**Ex.** Antilog (2) = $10^2 = 100$

Antilog (0.3010) = $10^{.3010} = 2$

Antilog [log (2)] = Antilog (0.3010) = 2

**pH - Scale :** Given by – **Sorenson**

pH Scale is called Sorenson scale.

pH scale is a measuring scale used to measure strength of acid and base and its value is equal to $-\log[H^+]$

i.e. $$pH = -\log[H^+] = \log \frac{1}{\left[H^+\right]}$$

**Ex.** $[H^+] = 10^{-3}$

$pH = -\log 10^{-3} = +3 \log 10 = 3$

**Conclusion :**

If $pH = x$ then $[H^+] = 10^{-x}$ or Vice versa

i.e. If $[H^+] = 10^{-x}$ then $pH = x$

**pOH** → It is equal to $-\log [OH^-]$ i.e. $pOH = -\log [OH^-] = \log \frac{1}{[OH^-]}$

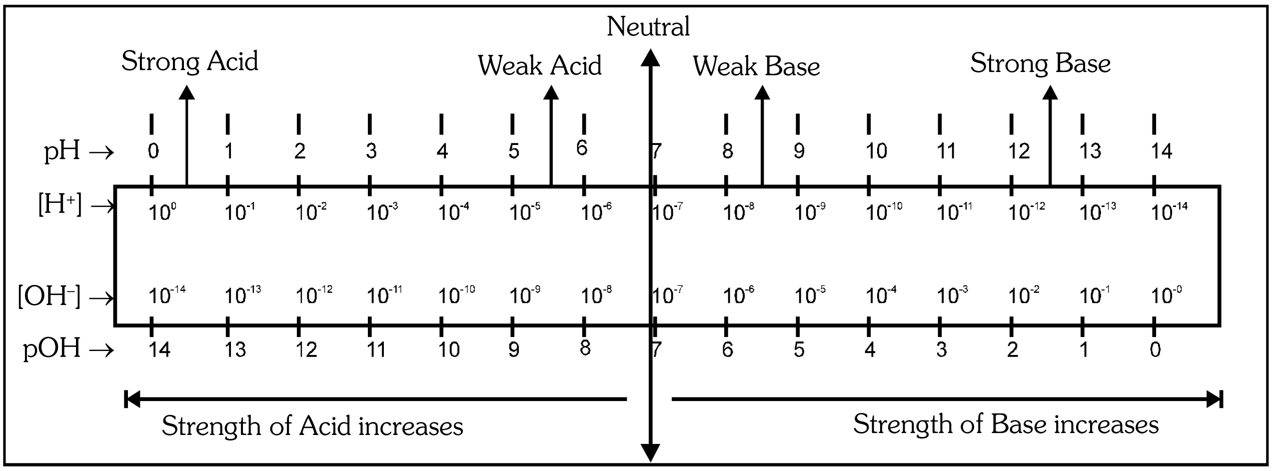
**Ex.** If $X = \frac{a}{b} \times 10^{-c}$, then find pX ?

**Sol.** $$pX = -\log X = -\log \left(\frac{a}{b} \times 10^{-c}\right)$$

$$pX = -\left[\log \frac{a}{b} + \log 10^{-c}\right] = -[\log a - \log b - c]$$

$$\boxed{pX = c + \log b - \log a}$$

**pH scale at 25°C**



**According to conductivity substances are of two types:**

**(1)** **Non-Conductor -** Those substances which do not show the flow of current or electricity.

**Ex.** Non - metals, plastic, rubber, wood etc.

**Exception** – Graphite is a non-metal but show conductivity due to motion of free electrons.

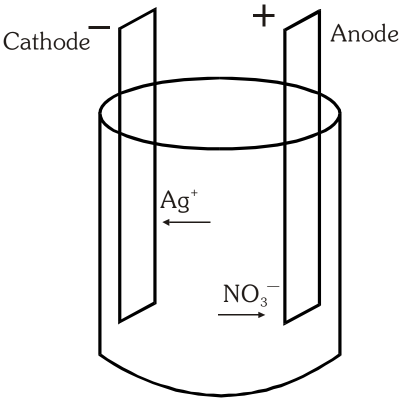
**(2)** **Conductors –** Those substances which show conductivity or flow of current are called conductors and these are of two types :

**(a)** **Metallic conductor –** Those conductor which show conductivity due to motion of free electrons.

**Ex.** All metals, Graphite

**(b)** **Ionic conductors –**

Those conductor which show conductivity due to movement of free ions. Ions are in free state in the solutions of ionic compounds. On passing electric current through the solution, ions move towards oppositely charged electrodes, i.e., the cation moves towards cathode (negative electrode) and the anion moves towards anode (positive electrode). Due to this reason, they are called cations and anions respectively. The current flows through the solution due to the movement of the ion Movement of ions through the solution of electrolyte ($AgNO_3$) towards oppositely charged electrodes.



**According to strength, ionic conductors are of two types –**

**(1)** **Strong electrolytes –** Those ionic conductors which are **completely ionized in aqueous solution** are called as strong electrolytes.

For strong electrolyte the value of degree of ionisation is 100% i.e. $\alpha = 1$

**Ex.** (a) Strong acid → $H_2SO_4$, $HCl$, $HNO_3$, $HClO_4$, $HBr$, $HI$

(b) Strong base → $KOH$, $NaOH$, $Ba(OH)_2$, $CsOH$, $RbOH$

(c) All Salts → $NaCl$, $KCl$, $CuSO_4$..........

**(2)** **Weak electrolytes –** Those electrolytes which are **partially ionized in aqueous solution** are called as weak electrolytes. For weak electrolytes the value of $\alpha$ is less than one.

**Ex.** (a) Weak acids → $HCN$, $CH_3COOH$, $HCOOH$, $H_2CO_3$, $H_3PO_3$, $H_3PO_2$, $B(OH)_3$, etc.

$$B(OH)_3 \downarrow H_3BO_3$$

(Boric acid)

(b) Weak bases → $NH_4OH$, $Cu(OH)_2$, $Zn(OH)_2$, $Fe(OH)_3$, $Al(OH)_3$ etc.

## 4.1 ARRHENIUS CONCEPT :

**(a)** According to Arrhenius when an electrolyte dissolves in water it splits up into two oppositely charged particles i.e. cation and anion.

**(b)** In an electrolytic solution (Aqueous solution of electrolyte), total +ve charge is equal to total –ve charge i.e. solution is electrically neutral.

⇒ In an electrolytic solution number of +ve ions may or may not be equal to the number of negative ions.

**(c)** Properties of an electrolytic solution are defined by its ions.

**Ex.** Blue colour of $CuSO_4$ aqueous solution is due to $Cu^{+2}$ ion (dark blue colour)

**(d)** When electric current is pass in aqueous solution of electrolyte then cation shows migration towards cathode where as anion shows migration towards anode.

**(e)** When a weak electolyte is dissolved in water an equilibrium is set up between unionized moles and ionized moles.

This condition of the reversible ionic reaction is known as ionic equilibrium. $AB\,(Aq.) \rightleftharpoons A^+(Aq.) + B^-(Aq.)$

According to L.M.A.

Ionisation constant $K = \dfrac{[A^+][B^-]}{[AB]}$

## Illustrations

**Illustration 1** If $[H^+] = \frac{5}{3} \times 10^{-4}$ then find pH ?

**Solution** $pH = -\log[H^+] = -\log\left(\frac{5}{3} \times 10^{-4}\right)$

$$= -\left[\log\frac{5}{3} + \log 10^{-4}\right] = -[\log 5 - \log 3 - 4]$$

$$= -[0.699 - 0.4771 - 4] = -[-3.778] = 3.778$$

**Illustration 2** ***Assertion :*** $H_2SO_4$ is a strong acid.

***Reason :*** $H_2SO_4$ undergoes almost completely ionised in aqueous solution.

(1) A (2) B (3) C (4) D

**Solution** **Ans. (1)**

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## 4.2 OSTWALD'S DILUTION LAW

⇒ Ostwald dilution law is applicable only for weak electrolytes.

**Statement :** According to Ostwald when solution of weak electrolyte is diluted then degree of ionisation of solution is increased, is called dilution law.

| Let | $AB\,(Aq.) \rightleftharpoons$ | $A^+(Aq.)$ + | $B^-(Aq.)$ |
|---|---|---|---|
| Initial concentration | C | 0 | 0 |
| Degree of ionisation (α) | $(C - C\alpha)$ | $(C\alpha)$ | $(C\alpha)$ |

According to L.M.A.

Ionisation constant $K = \dfrac{[A^+][B^-]}{[AB]}$ ; $K = \dfrac{C\alpha \times C\alpha}{C(1-\alpha)} = \dfrac{C\alpha^2}{(1-\alpha)}$

If $\alpha <<<< 1$ then $1 - \alpha \simeq 1$

$$K = C\alpha^2 \quad \text{or} \quad \alpha = \sqrt{\frac{K}{C}} \quad \text{(K = constant, At constant temperature)}$$

$$\alpha \propto \frac{1}{\sqrt{C}} \quad \left(C \propto \frac{1}{V}\right) \quad \alpha \propto \sqrt{V}$$

volume ≡ dilution $\alpha \propto \sqrt{\text{dilution}}$ by dilution law

dilution ↑ α ↑

At infinite dilution, $\alpha = 100\%$

Thus for weak electrolytes the degree of ionisation is directly proportional to square root of dilution or inversely proportional to square root of concentration. This law is known as Ostwald Dilution Law.

## GOLDEN KEY POINTS

- At infinite dilution the value of α becomes equal to one.
- The value of concentration should be in terms of normality for calculation of pH.

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**Application of Ostwald's Dilution Law :** $\boxed{K = C\alpha^2}$

**(a)** For mono basic weak acid (HA)

(i). Ionisation constant ($K_a$)

| | HA ⇌ | $H^+$ + | $A^-$ |
|---|---|---|---|
| Initial concentration | C | 0 | 0 |
| At equilibrium | C – Cα | Cα | Cα |

If degree of ionisation is α

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha \times C\alpha}{C - C\alpha}$$

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

∵ $\alpha <<<< 1$ ∴ $(1 - \alpha) \approx 1$

∴ $K_a = C\alpha^2$

Where

$K_a \rightarrow$ Ionisation constant of weak acid
$K_b \rightarrow$ Ionisation constant of weak base
} Both $K_a$ and $K_b$ depends only on temperature

(ii). $[H^+]$ (Concentration of $H^+$)

$[H^+] = C\alpha$ ........(1)

$K_a = C\alpha^2$ or $\alpha = \sqrt{\frac{K_a}{C}}$ .......(2)

from Eq. (1) and (2) $[H^+] = C \times \frac{\sqrt{K_a}}{\sqrt{C}}$

$[H^+] = \sqrt{K_a \times C}$

(iii). pH = –log $[H^+]$

put the value of $[H^+]$

$$pH = -\log(\sqrt{K_a \times C}) = -\log(K_a \times C)^{1/2}$$

$$pH = -\frac{1}{2}[\log K_a + \log C]$$

$$pH = -\frac{1}{2}\log K_a - \frac{1}{2}\log C$$

$$pH = \frac{1}{2}pK_a - \frac{1}{2}\log C$$

**(b)** For mono acidic weak base (BOH)

Ionisation constant ($K_b$)

| | BOH ⇌ | $B^+$ + | $OH^-$ |
|---|---|---|---|
| Initial concentration | C | 0 | 0 |
| At equilibrium | C – Cα | Cα | Cα |

If degree of ionisation is α

$$K_b = \frac{[B^+][OH^-]}{[BOH]} = \frac{C\alpha \times C\alpha}{C - C\alpha}$$

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

∵ $\alpha <<<< 1$ ∴ $(1 - \alpha) \approx 1$

∴ $K_b = C\alpha^2$

$[OH^-]$ (Concentration of $OH^-$)

$[OH^-] = C\alpha$ ........(1)

$K_b = C\alpha^2$ or $\alpha = \sqrt{\frac{K_b}{C}}$ .........(2)

from Eq. (1) and (2) $[OH^-] = C \times \frac{\sqrt{K_b}}{\sqrt{C}}$

$[OH^-] = \sqrt{K_b \times C}$

pOH = – log $[OH^-]$

put the value of $[OH^-]$

$$pOH = -\log(\sqrt{K_b \times C}) = -\log(K_b \times C)^{1/2}$$

$$pOH = -\frac{1}{2}[\log K_b + \log C]$$

$$pOH = -\frac{1}{2}\log K_b - \frac{1}{2}\log C$$

$$pOH = \frac{1}{2}pK_b - \frac{1}{2}\log C$$

**In summary :**

| | | | | |
|---|---|---|---|---|
| 1. | $K_a = C\alpha^2$ | | 1. | $K_b = C\alpha^2$ |
| 2. | $[H^+] = C\alpha = \sqrt{K_a \times C}$ | | 2. | $[OH^-] = C\alpha = \sqrt{K_b \times C}$ |
| 3. | $pH = -\log[H^+]$ | | 3. | $pOH = -\log[OH^-]$ |
| or | $pH = \frac{1}{2}pK_a - \frac{1}{2}\log C$ | | or | $pOH = \frac{1}{2}pK_b - \frac{1}{2}\log C$ |

**Limitation of Ostwald Dilution Law :**

It is not applicable for strong electrolytes.

**Factors affecting the Value of Degree of ionisation :**

(1) Temperature → On increasing temperature, ionization increases so $\alpha$ increases because dissociation is endothermic process.

(2) Dilution → $\alpha \propto \sqrt{V}$ so on dilution, $\alpha$ increases.

(3) Nature of electrolytes

| (i) Strong electrolytes | (ii) Weak electrolytes |
|---|---|
| $\alpha = 100\%$ | $\alpha < 100\%$ |

(4) Nature of solvent

If Dielectric constant $\mu$ of solvent increases then the value of $\alpha$ increases.

$H_2O \rightarrow \mu = 81$

$D_2O \rightarrow \mu = 79$

$C_6H_6 \rightarrow \mu = 2.5$

$CCl_4 \rightarrow \mu = 0$

**Mixing of ions :**

| **(a) Common ion effect** | **(b) Odd ion effect** |
|---|---|
| When a strong electrolyte having a common ion, is mixed with weak electrolyte then the degree of ionisation ($\propto$) of weak electrolyte is decreased. This effect is called common ion effect. | When a strong electrolyte having a odd ion, is mixed with weak electrolyte then the degree of ionisation ($\propto$) of weak electrolyte is increased. This effect is called odd ion effect. |

**Common ion :** On mixing $CH_3COONa$ with $CH_3COOH$ solution

Initially $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

(On mixing) $CH_3COONa \longrightarrow CH_3COO^- + Na^+$ ; $[CH_3COO^-]$ increases so $[H^+]$ decreases respectively.

**Odd ion :** On mixing, NaOH with $CH_3COOH$ solution

Initially $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

(On mixing) $NaOH \longrightarrow Na^+ + OH^-$

Due to odd ion effect $H^+ + OH^- \rightarrow$ Water

## Illustrations

**Illustration 3** For which, dilution law is applicable :

(1) NaCl (SASB salt) (2) HCl (SA)

(3) $CH_3COONa$ (WASB salt) (4) None

**Solution** **Ans. (4)**

**Illustration 4** Find out the value of $\alpha$ of $10^{-2}$ M HCN solution if $[H^+] = 10^{-3}$.

**Solution** $[H^+] = C\alpha$

$$10^{-3} = 10^{-2}\alpha \text{ or } \alpha = \frac{10^{-3}}{10^{-2}} = 10^{-1} \text{ or } \alpha = 0.1$$

$\alpha = 10\%$

**Illustration 5** For 10 M $CH_3COOH$ solution if $K_a = 10^{-5}$ then find out :

(i) $\alpha$ (ii) $[H^+]$ (iii) pH

**Solution** (i) $\alpha$

(Degree of ionisation) :– $K_a = C\alpha^2$

$$10^{-5} = 10\alpha^2 \text{ or } \alpha^2 = \frac{10^{-5}}{10} = 10^{-6} \text{ or } \alpha = 10^{-3}$$

(ii) $[H^+] = C\alpha = 10 \times 10^{-3} = 10^{-2}$

(iii) $pH = -\log [H^+] = -\log 10^{-2} = 2$

**Illustration 6** For $10^{-3}$ M $H_2CO_3$ if $\alpha = 10\%$ then find out the value of pH ?

**Solution** Normality $= M \times V.F. = 10^{-3} \times 2 = 2 \times 10^{-3}$

$[H^+] = C\,\alpha = 2 \times 10^{-3} \times 0.1 = 2 \times 10^{-4}$

$pH = -\log [H^+] = -\log (2 \times 10^{-4}) = -\log 2 - \log 10^{-4} = -0.3010 + 4 = 3.7$

**Illustration 7** ***Assertion*** **:** NaCl + HCl does not experience common ion effect.

***Reason*** **:** Both NaCl and HCl are strong electrolytes.

(1) A (2) B (3) C (4) D

**Solution** **Ans. (1)**

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## 4.3 EXPLANATION OF WATER

**a.** Nature of water is neutral. $[H^+] = [OH^-]$ [$=10^{-7}$M (at 25°C)]

i.e. at 25°C pH = 7 and pOH = 7

$\therefore$ pH = pOH

**b.** No. of $H_2O$ moles in 1 litre water $= \frac{1000}{18} = 55.5$ moles

**c.** Molar concentration of $H_2O = 55.5$ mol $L^{-1}$

**d.** Number of $H_2O$ molecules in 1 litre water $= 55.5\ N_A$ ($N_A$= AVOGADRO'S NUMBER)

**e.** Concentration of $H^+$ and $OH^-$ ions in 1 litre water

$[H^+] = 10^{-7}$ mol $L^{-1}$ and $[OH^-] = 10^{-7}$ mol $L^{-1}$

**f.** No. of $H^+$ and $OH^-$ ions in 1 litre water

No. of $H^+$ ions $= 10^{-7}N_A$ and No. of $OH^-$ ions $= 10^{-7}\ N_A$

**g.** In water (Number of $H_2O$ molecules : Number of $H^+$ ions)

$= 55.5\ N_A : 10^{-7}\ N_A$

$= 55.5 \times 10^7 : 1$

i.e. one $H^+$ ion is obtained from $55.5 \times 10^7$ $H_2O$ molecules

**h. Degree of ionisation of water (α) :**

$$H_2O \rightleftharpoons H^+ + OH^-$$

According to Ostwald's dilution law

$[H^+] = c\,\alpha$

$$\alpha = \frac{10^{-7}}{55.5} = 1.8 \times 10^{-9}$$

$$\boxed{\alpha = 1.8 \times 10^{-7}\%}$$

Hence, water is a very weak electrolyte.

**i. K (Ionisation constant of water) :**

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$K = \frac{[H^+][OH^-]}{[H_2O]}$$

$$K = \frac{10^{-7} \times 10^{-7}}{55.5} \quad \text{or} \quad \boxed{K = 1.8 \times 10^{-16}}$$

**j. Ionic product of water $K_w$ :**

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$K = \frac{[H^+].[OH^-]}{[H_2O]}$$

$$K \times [H_2O] = [H^+][OH^-]$$

Since, dissiociation takes place to a very small extent, $[H_2O]$ may be regarded as constant.

Thus, the product $K.[H_2O]$ gives another constant which designated as $K_{w.}$ . So,

$$\boxed{K_W = [H^+][OH^-]}$$

At 25°C – $K_W = 10^{-7} \times 10^{-7} = 10^{-14}$

$K[H_2O] = K_W \Rightarrow \boxed{K_W > K}$ (always)

Various forms of $K_W$

(a) $K_W = [H^+][OH^-]$ for water $[H^+] = [OH^-]$

(b) $K_W = [H^+]^2$

(c) $K_W = [OH^-]^2$

(d) $K_W = [H_3O^+][OH^-]$

(e) $K_W = [H_3O^+]^2$ $\{[H_3O^+] = [H^+]\}$

Hydronium ion Proton

- **Relation in between pH and pOH :**

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

taking –log on both sides

$-\log K_W = -\log[H^+] - \log[OH^-]$

$$\boxed{pK_w = pH + pOH}$$

- Nature of water is neutral so,

$[pH = pOH]$

| | |
|---|---|
| $pK_W = pH + pH$ | $pK_W = pOH + pOH$ |
| $2pH = pK_W$ | $2\ pOH = pK_W$ |
| $pH = \frac{pK_W}{2}$ | $pOH = \frac{pK_W}{2}$ |

$$\boxed{pH = pOH = \frac{pK_W}{2}}$$

At 25°C, $K_W = 10^{-14}$ or $pK_W = 14$ $\therefore pH + pOH = 14$ or $pH = pOH = 7$

**k. Effect of temperature :-**

$$H_2O \rightleftharpoons H^+ + OH^-$$

Ionization of water is endothermic process so, on increasing temperature $\alpha$ increases or $[H^+]$ and $[OH^-]$ increases or $[H^+][OH^-]$ increases i.e. $K_w$ increases means pH decreases or pOH decreases.

$T\uparrow = \alpha\uparrow = [H^+]\ [OH^-]\uparrow = K_W\uparrow \Rightarrow pK_W\downarrow$

At 25° C, $K_W = 10^{-14}$

At 90° C, $K_W = 10^{-12}$

| Parameters | At 25°C | At 90°C |
|---|---|---|
| $K_W$ | $10^{-14}$ | $10^{-12}$ |
| $pK_W$ | 14 | 12 |
| $pH = \frac{pK_w}{2} = pOH$ | 7 | 6 |
| $[H^+] = [OH^-] = \sqrt{K_W}$ | $10^{-7}$ | $10^{-6}$ |
| $pH + pOH = pK_W$ | 14 | 12 |

## GOLDEN KEY POINTS

- On increasing temperature, both $[H^+]$ and $[OH^-]$ increases equally so water remains neutral but neutral pH changes from 7 to 6 at 90°C.

## Illustrations

**Illustration 8** Dissociation constant of water at 25° C is

(1) $10^{-14} \times (55.5)^{-1}$ (2) $10^{-7} \times (18)^{-1}$ (3) $10^{-14} \times (18)^{-1}$ (4) $10^{-7} \times (55.4)^{-1}$

**Solution** **Ans. (1)**

**Illustration 9** What should be the number of $H^+$ ions in 1 mL of distilled water, if the number of $H^+$ ions in 1 L is $6.023 \times 10^{16}$?

**Solution** Number of $H^+$ ions in 1 mL distilled water $= \frac{6.023\times10^{16}}{1000} = 6.023\times10^{13}$

## BEGINNER'S BOX-1

**1.** The pH of a 0.005 M $H_2SO_4$ solution is–

(1) 3.3 (2) 5.0 (3) 2.0 (4) 4.0

**2.** If pure water has $pK_w = 13.36$ at 50°C, the pH of pure water will be–

(1) 6.68 (2) 7.0 (3) 7.13 (4) 6.0

**3.** How many $H^+$ ions are present in 1 ml of a solution whose pH is 13 ?

(1) $10^{-16}$ (2) $6.022 \times 10^{13}$ (3) $6.022 \times 10^{7}$ (4) $6.022 \times 10^{23}$

**4.** The pH of solutions A, B, C and D are 9.5, 2.5, 3.5 and 5.5 respectively. The most acidic solution is–

(1) D (2) C (3) A (4) B

**5.** Calculate the concentration of the formate ion present in 0.100 M formic acid (HCOOH) solution at equilibrium ($K_a = 1.7 \times 10^{-4}$).

(1) $4.1 \times 10^{-3}$ M (2) $3.1 \times 10^{-3}$ M (3) $2.1 \times 10^{-3}$ M (4) $5.1 \times 10^{-3}$ M

**6.** Which of the following is the weakest acid ?

(1) Phenol ($K_a = 1.3 \times 10^{-10}$) (2) Hydrocyanic acid ($K_a = 4.9 \times 10^{-10}$)

(3) Acetic acid ($K_a = 1.8 \times 10^{-5}$) (4) Benzoic acid ($K_a = 6.5 \times 10^{-5}$)

**7.** The pH of 0.1 M monobasic acid is 4.50. The acidity constant ($K_a$) of the monobasic acid is–

(1) $1.0 \times 10^{-7}$ (2) $1.0 \times 10^{-5}$ (3) $1.0 \times 10^{-4}$ (4) $1.0 \times 10^{-8}$

**8.** Which of the following is the strongest base ?

(1) $C_6H_5NH_2$ ($pK_b = 9.42$) (2) $C_6H_5NHCH_3$ ($pK_b = 9.15$)

(3) $C_6H_5N(CH_3)_2$ ($pK_b = 8.94$) (4) $C_6H_5NHC_2H_5$ ($pK_b = 8.89$)

**9.** Value of dissociation constant of acetic acid is $10^{-6}$, where as dissociation constant of formic acid is $10^{-5}$. Which of the following will be the value of pKa (acetic acid) - pKa (formic acid)

(1) 10 (2) +1 (3) $10^{-1}$ (4)–1

**10.** A solution has pOH equal to 13 at 298 K. The solution will be

(1) highly acidic (2) highly basic (3) moderately basic (4) unpredictable

**11.** What would be $[H^+]$ of 0.006 M benzoic acid ($K = 6 \times 10^{-5}$)

(1) $0.6 \times 10^{-4}$ (2) $6 \times 10^{-4}$ (3) $6 \times 10^{-3}$ (4) $3.6 \times 10^{-4}$

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## 4.4 SALTS, TYPES OF SALT AND CONJUGATE THEORY

**(A) SALT :** Salts are regarded as compounds made up of positive and negative ions. The positive part comes from a base while negative part from an acid. Salts are ionic compounds.

i.e. A compound formed by the combination of acid and base is known as salt.

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}; \quad \Delta H = -ve$$

$$HCl + NaOH \rightarrow NaCl\ (\text{Salt}) + H_2O$$

### (B) TYPES OF SALT

**(a) Normal/general salts :–** The salts formed by the loss of all possible protons. (replaceable hydrogen ions as $H^+$) are called normal salts. Such a salt does not contain either a replaceable hydrogen ion or a hydroxyl ($OH^-$) group.

**Ex.:** NaCl, $NaNO_3$, $K_2SO_4$, $Ca_3(PO_4)_2$, $Na_2HPO_3$, $NaH_2PO_2$ etc.

**(b) Acid salts :–** Salts formed by incomplete neutralisation of poly basic acids are called acid salts. Such salts still contain one or more replaceable hydrogen ions ($H^+$).

**Ex.:** $NaHCO_3$, $NaHSO_4$, $NaH_2PO_4$, $Na_2HPO_4$ etc.

**(c) Basic salts :–** Salts are formed by incomplete neutralisation of poly acidic base are called basic salts. Such salts still contain one or more hydroxyl ion group.

**Ex.:** Zn(OH)Cl, Mg(OH)Cl, $Fe(OH)_2Cl$, $Bi(OH)_2Cl$ etc.

**(d)** **Double salts :–** The addition compounds formed by the combination of two simple salts are termed double salts. Such salts are stable in solid state only i.e. when dissolved in water they break into simple ions.

**Ex. :** $FeSO_4.(NH_4)_2SO_4.6H_2O$, $K_2SO_4.\ Al_2(SO_4)_3.24H_2O$ etc.

(Ferrous ammonium sulphate) (Potash alum)

**(e)** **Complex salts :–** These salts when dissolved in water give complex ions. Complex ions are stable in solid state as well as in solutions.

**Ex. :** $K_4[Fe(CN)_6]$, $[Co(NH_3)_6]SO_4$, $[Ag(NH_3)_2Cl$ etc.

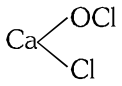
**(f)** **Mixed salts :–** The salts which furnishes more than one type of cations or more than one type of anions when dissolved in water are called mixed salts.

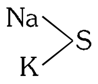
These are formed by the neutralisation of more than one acids and bases.

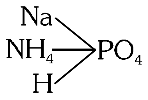
**Ex.** (i)

$$\underset{\text{acid}}{H_2SO_4} + \underset{\text{base}}{NaOH} \longrightarrow \underset{\text{acidic salt}}{NaHSO_4} + H_2O$$

$$\underset{\text{acidic salt}}{NaHSO_4} + \underset{\text{base}}{KOH} \longrightarrow \underset{\text{(mixed salt)}}{NaKSO_4} \longrightarrow \underset{\text{(more than two radicals)}}{Na^+\,K^+\,SO_4^{-2}}$$

(ii) 
$Ca(OCl)Cl$

(iii) 
$NaKS$

(iii) 
$NaNH_4HPO_4$

**(a) Types of general salts :**

(i) SASB (ii) SAWB (iii) WASB (iv) WAWB

**(C)** **Conjugate acid-base pair :**

(i) Conjugate acid-base pairs means difference between two species of only one $H^+$ ion.

(ii) When an acid loses a proton, the residual part of it has a tendency to regain a proton. Therefore it behaves as a base.

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

Weak acid Conjugate strong base

($K_a = 1.85 \times 10^{-5}$) (acceptor ion)

$$NH_4OH + H^+ \rightleftharpoons NH_4OH_2^+ \longrightarrow NH_4^+ + H_2O$$

Weak base Conjugate strong acid

($K_b = 1.85 \times 10^{-5}$) (acceptor ion)

$$HCl \longrightarrow H^+ + Cl^-$$

Strong acid Conjugate weak base (spectator ion)

Accepting tendency of $H^+ \simeq 0$

$$NaOH + H^+ \longrightarrow NaOH_2^+ \longrightarrow Na^+ + H_2O$$

Strong base Conjugate weak acid

(spectator ion)

**Note :** Strong acids have weak conjugate bases while weak acids have strong conjugate bases. Similarly, strong bases have weak conjugate acids while weak bases have strong conjugate acids.

**(D) Relation between conjugate acid-base pair :–**

**Example –** $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

Acid — conjugate base

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

Base — conjugate acid

| **$CH_3COOH$**<br>**acid** | **$CH_3COO^-$**<br>**conjugate base** |
|---|---|
| $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ | $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ |
| $K_a = \dfrac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$ ...... (i) | $K_b = \dfrac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$ ..... (ii) |

In both the reactions $H_2O$ in excess quantity so active mass of $H_2O$ is one.

Now multiply the equation (i) and (ii) $K_a \times K_b = [H^+][OH^-]$

we know $[H^+] \times [OH^-] = K_W$ (Ionic product of water) $\boxed{K_a \times K_b = K_w}$

Taking –log on both sides $\boxed{pK_a + pK_b = pK_w}$

we know that for water at 25°C, $K_W = 10^{-14}$ or $pK_W = 14$

So $\boxed{K_a \times K_b = 10^{-14}}$ or $\boxed{pK_a + pK_b = 14}$

Above relation is applicable only for conjugate acid-base pairs.

## Illustrations

**Illustration 10** Which salt is not an example of acidic salt :-

(1) HCOONa (2) $NaH_2PO_2$ (3) NaHS (4) (1) and (2) both

**Solution** **Ans. (4)**

**Illustration 11** Calculate the no. of ions in Mohr's salt $[FeSO_4.(NH_4)_2SO_4.6H_2O]$ or Ferrous Ammonium Sulphate

**Solution** $FeSO_4.(NH_4)_2SO_4.6H_2O \longrightarrow Fe^{+2} + SO_4^{-2} + 2NH_4^+ + SO_4^{-2} + 6H_2O$

Total number of ions = 5

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### 4.5. HYDROLYSIS OF SALTS

Salt hydrolysis is defined as the process in which water reacts with cation or anion or both of a salt to change the concentration of $H^+$ and $OH^-$ ions of water.

Salt hydrolysis is reverse process of neutralization.

Water + Salt $\rightleftharpoons$ Acid + Base ; $\Delta H = +ve$

**(A) Hydrolysis of strong acid and strong base [SA – SB] types of salt –**

**Ex.** NaCl, $BaCl_2$, $Na_2SO_4$, $KClO_4$, $BaSO_4$, $NaNO_3$, KBr, KCl etc.

$Na^+ + Cl^- + H_2O \longrightarrow Na^+ + OH^- + H^+ + Cl^-$

$H_2O \rightleftharpoons H^+ + OH^-$ (It is not salt hydrolysis)

(i) Hydrolysis of salt of [SA – SB] is not possible as both cation and anion are not reactive.

(ii) Aqueous solution of these type of salt is neutral in nature. (pH = pOH = 7)

(iii) pH of the solution is 7.

(iv) No effect on litmus paper.

## (B) Hydrolysis of strong acid and weak base [SA - WB] types of salt –

**Ex.** $CaSO_4$, $NH_4Cl$, $(NH_4)_2SO_4$, $Ca(NO_3)_2$, $ZnCl_2$, $CuCl_2$, $CaCl_2$, AgCl, AgI, $AgNO_3$ etc

$$NH_4Cl + H_2O \rightleftharpoons \underset{WB}{NH_4OH} + \underset{SA}{HCl}$$

$$NH_4^+ + Cl^- + H_2O \rightleftharpoons NH_4OH + H^+ + Cl^-$$

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

(i) In this type of salt hydrolysis, cation reacts with $H_2O$ therefore called as cationic hydrolysis. The cation of the salt which has come from weak base is reactive.

(ii) Solution is acidic in nature as $[H^+]$ is increased.

(iii) pH of the solution is less than 7.

(iv) Solution turns blue litmus paper red.

$K_h$ = Hydrolysis constant
$K_W$ = Ionic product of water
$K_a$ = Ionisation constant of acid
$K_b$ = Ionisation constant of base
h = Degree of hydrolysis
C = Concentration of salt (concentration of ions)

### (a) Relation between $K_h$, $K_W$ and $K_b$

$$B^+ + H_2O \rightleftharpoons BOH + H^+$$

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

Hydrolysis constant $[K_h]$

$$K_h = \frac{[NH_4OH][H^+]}{[NH_4^+]} \quad .......(1)$$

For weak Base $\quad NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$$K_b = \frac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_4OH\right]} \quad ........(2)$$

For water $\quad H_2O \rightleftharpoons H^+ + OH^-$

$$K_W = [H^+][OH^-] \quad ........(3)$$

Now mutiplying Eq. (1) and (2) = Eq. (3)

$$\frac{\left[NH_4OH\right]\left[H^+\right]}{\left[NH_4^+\right]} \times \frac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_4OH\right]} = [H^+][OH^-]$$

i.e. $\quad K_h \times K_b = K_W$

$$\boxed{K_h = \frac{K_W}{K_b}} \quad ........(4)$$

### (b) Degree of hydrolysis – Represented by h

$$\begin{array}{lccc} & NH_4^+ + H_2O \rightleftharpoons & NH_4OH & + \; H^+ \\ \text{Initial concentration of salt} & C & 0 & 0 \\ & C - Ch & Ch & Ch \end{array}$$

$$K_h = \frac{\left[NH_4OH\right]\left[H^+\right]}{\left[NH_4^+\right]} = \frac{Ch \times Ch}{C - Ch} = \frac{C^2h^2}{C(1-h)} = \frac{Ch^2}{(1-h)}$$

Since $\quad h <<<< 1 \quad$ then $\quad (1 - h) \approx 1$

$$\therefore \quad \boxed{K_h = Ch^2} \qquad \text{........(5)}$$

$$h^2 = \frac{K_h}{C} \quad \Rightarrow \quad h = \sqrt{\frac{K_h}{C}} \quad \text{........(6)}$$

$$\therefore \quad K_h = \frac{K_W}{K_b} \quad \Rightarrow \quad h = \sqrt{\frac{\frac{K_W}{K_b}}{C}}$$

$$h = \sqrt{\frac{K_W}{K_b \times C}} \qquad \text{........(7)}$$

**(c)** **pH of the solution :** pH = – log [H$^+$]

$$[H^+] = Ch = C\sqrt{\frac{K_W}{K_b \times C}} \quad \Rightarrow \quad [H^+] = \sqrt{\frac{K_W \times C}{K_b}} \quad \text{........(8)}$$

taking – log on both sides

$$-\log[H^+] = -\log\sqrt{\frac{K_W \times C}{K_b}} \quad \Rightarrow \quad pH = -\log\left(\frac{K_W \times C}{K_b}\right)^{½}$$

$$pH = -\frac{1}{2}\,[\log K_W + \log C - \log K_b]$$

$$pH = -\frac{1}{2}\log K_W - \frac{1}{2}\log C - \frac{1}{2}(-\log K_b)$$

$$pH = \frac{1}{2}pK_W - \frac{1}{2}\log C - \frac{1}{2}pK_b$$

$$\boxed{pH = 7 - \frac{1}{2}pK_b - \frac{1}{2}\log C} \qquad \text{........(9)}$$

**(C)** **Hydrolysis of weak acid and strong base [WA – SB] types of salt –**

**Ex.** $CH_3COONa$, $HCOONa$, $KCN$, $NaCN$, $K_2CO_3$, $BaCO_3$, $K_3PO_4$ etc.

$$CH_3COONa + H_2O \rightleftharpoons \underset{WA}{CH_3COOH} + \underset{SB}{NaOH}$$

$$CH_3COO^- + Na^+ + H_2O \rightleftharpoons CH_3COOH + Na^+ + OH^-$$

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$

(i) In this type of salt hydrolysis, anion reacts with water therefore called as anionic hydrolysis. The anion of the salt which has come from weak acid is reactive.

(ii) Solution is basic in nature as [OH$^-$] increases.

(iii) pH of the solution is greater than 7.

(iv) Solution turns red litmus paper blue.

**(a)** **Relation between $K_h$ , $K_W$ and $K_a$**

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$

$$K_h = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} \qquad \text{........(1)}$$

For weak acid

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

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \qquad \text{........(2)}$$

For water

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$K_W = [H^+][OH^-] \qquad \text{........(3)}$$

Now multiply eq. (1) × eq. (2) = eq. (3)

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

$$K_h \times K_a = K_W$$

$$\boxed{K_h = \frac{K_W}{K_a}} \qquad \text{........(4)}$$

**(b) Degree of hydrolysis (h) :**

| | $CH_3COO^-$ | + | $H_2O$ | $\rightleftharpoons$ | $CH_3COOH$ | + | $OH^-$ |
|---|---|---|---|---|---|---|---|
| Initial concentration of salt | C | | | | 0 | | 0 |
| | C – Ch | | | | Ch | | Ch |

$$K_h = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} = \frac{Ch \times Ch}{C - Ch} = \frac{C^2h^2}{C(1-h)}$$

$$K_h = \frac{Ch^2}{(1-h)}$$

Since h <<<< 1 then $(1 - h) \approx 1$

$$\therefore \quad \boxed{K_h = Ch^2} \qquad \text{........(5)}$$

$$h^2 = \frac{K_h}{C} \quad \text{or} \quad h = \sqrt{\frac{K_h}{C}} \qquad \text{........(6)}$$

$$\boxed{h = \sqrt{\frac{K_W}{K_a \times C}}} \qquad \text{........(7)}$$

**(c) pH of the solution**

$$[OH^-] = Ch$$

$$[OH^-] = C \times \sqrt{\frac{K_W}{K_a \times C}} \quad \text{or} \quad [OH^-] = \sqrt{\frac{K_W \times C}{K_a}} \qquad \text{........(8)}$$

taking – log on both sides

$$-\log[OH^+] = -\log\left(\frac{K_w.C}{K_a}\right)^{1/2}$$

$$pOH = -\frac{1}{2}[\log K_W + \log C - \log K_a]$$

$$pOH = \frac{1}{2}pK_W - \frac{1}{2}pK_a - \frac{1}{2}\log C \quad \text{or} \quad pOH = 7 - \frac{1}{2}pK_a - \frac{1}{2}\log C$$

$$\because \quad pH + pOH = 14$$

$$pH = 14 - pOH$$

$$\therefore \quad pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C \qquad \text{........(9)}$$

**(D) Hydrolysis of weak acid and weak base (WA - WB) type of salt :**

**Ex.** $CH_3COONH_4$, $AgCN$, $NH_4CN$, $CaCO_3$, $[NH_4]_2CO_3$, $ZnHPO_3$ etc.

$$CH_3COONH_4 + H_2O \rightleftharpoons \underset{WA}{CH_3COOH} + \underset{WB}{NH_4OH}$$

$$CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + NH_4OH$$

(i) Maximum hydrolysis occurs of the salt of (WA – WB) as both the cation and anion are reactive.

(ii) Solution is almost neutral but it may be acidic or basic depending upon the nature of acid base

(iii) pH of the solution is near to 7.

**For WA - WB types of salt :**

| Terms | $K_a > K_b$ | $K_b > K_a$ | $K_a = K_b$ |
|---|---|---|---|
| **1.** Hydrolysis | Cationic-anionic | Anionic-cationic | Neutral hydrolysis |
| **2.** Nature | Acidic | Basic | Neutral |
| **3.** pH | pH < 7 | pH > 7 | pH = 7 |

**(a) Relation between $K_h$, $K_W$, $K_a$ and $K_b$**

$$CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + NH_4OH$$

$$K_h = \frac{[CH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]} \quad ........(1)$$

For weak base

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

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} \quad ........(2)$$

For weak acid

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

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \quad ........(3)$$

For water

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$K_W = [H^+][OH^-] \quad ........(4)$$

Multiply Eq. (1) × Eq. (2) × Eq. (3) = Eq. (4)

$$\frac{[CH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]} \times \frac{[NH_4^+][OH^-]}{[NH_4OH]} \times \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = [H^+][OH^-]$$

$$K_h \times K_b \times K_a = K_W$$

$$\boxed{K_h = \frac{K_W}{K_a \times K_b}} \quad ........(5)$$

**(b) Degree of hydrolysis (h) –**

| | $CH_3COO^-$ | + $NH_4^+ + H_2O$ | $\rightleftharpoons$ | $CH_3COOH$ | + $NH_4OH$ |
|---|---|---|---|---|---|
| Initial concentration of salt | C | C | | 0 | 0 |
| | C – Ch | C – Ch | | Ch | Ch |

$$K_h = \frac{[CH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]} = \frac{Ch \times Ch}{(C-Ch)(C-Ch)} = \frac{C^2h^2}{C(1-h)\times C(1-h)}$$

Since $h <<<< 1$ then $(1 - h) \approx 1$

$$\therefore \quad \boxed{K_h = h^2} \qquad .......(6)$$

$$\text{or} \quad h^2 = \frac{K_W}{K_a \times K_b} \quad \text{or} \quad \boxed{h = \sqrt{\frac{K_W}{K_a \times K_b}}} \qquad .......(7)$$

**(c) pH of the solution**

from equation (3)

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

$$[H^+] = \frac{K_a \times [CH_3COOH]}{[CH_3COO^-]} = \frac{K_a \times Ch}{C - Ch} = \frac{K_a \times h}{1-h}$$

Since $h <<<< 1$ then $(1 - h) \approx 1$

$[H^+] = K_a \times h$ [Now put the value of h from eq. (5)]

$$= K_a \times \sqrt{\frac{K_W}{K_a \times K_b}}$$

$$\boxed{[H^+] = \sqrt{\frac{K_W \times K_a}{K_b}}} \qquad .......(8)$$

taking – log on both sides $-\log [H^+] = -\log \left(\frac{K_W \times K_a}{K_b}\right)^{1/2}$

$$pH = -\frac{1}{2}\,[\,\log (K_w \times K_a) - \log K_b\,]$$

$$pH = -\frac{1}{2}\,[\,\log K_W + \log K_a - \log K_b]$$

$$pH = -\frac{1}{2}\,[\,\log K_W] - \frac{1}{2}\,[\log K_a] - \frac{1}{2}\,[-\log K_b]$$

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

$$\boxed{pH = 7 + \frac{1}{2}pK_a - \frac{1}{2}pK_b} \qquad .......(9)$$

## GOLDEN KEY POINTS

- Degree of hydrolysis, pH of weak acid and weak base [WA – WB] type salts do not depend on the concentration of salt.
-

| | SA SB salts | SA WB salts | WA SB salts | WA WB salts |
|---|---|---|---|---|
| **1.** | Neutral solution | Acidic solution | Basic solution | Almost neutral solution |
| **2.** | No hydrolysis | Cationic hydrolysis | Anionic hydrolysis | Hydrolysis by both cation and anion |
| **3.** | – | $K_h = \frac{K_W}{K_b}$ | $K_h = \frac{K_W}{K_a}$ | $K_h = \frac{K_W}{K_a \times K_b}$ |
| **4.** | – | $h = \sqrt{\frac{K_W}{K_b.C}}$ | $h = \sqrt{\frac{K_W}{K_a.C}}$ | $h = \sqrt{\frac{K_W}{K_a.K_b}}$ |
| **5.** | – | $\left[H^+\right] = C.h = \sqrt{\frac{K_W.C}{K_b}}$ | $\left[OH^-\right] = C.h = \sqrt{\frac{K_W.C}{K_a}}$ | $\left[H^+\right] = K_a.h = \sqrt{\frac{K_W.K_a}{K_b}}$ |
| **6.** | – | $pH = 7 - \frac{1}{2}\ pK_b - \frac{1}{2}\log C$ | $pH = 7 + \frac{1}{2}\ pK_a + \frac{1}{2}\log C$ | $pH = 7 + \frac{1}{2}\ pK_a - \frac{1}{2}\ pK_b$ |

## Illustrations

**Illustration 12** What is the pH of 1 M $CH_3COONa$ solution? Ka of acetic acid = $1.8 \times 10^{-5}$, $K_w = 10^{-14}$ mol$^2$ L$^{-2}$

(1) 2.4 (2) 3.6 (3) 4.8 (4) 9.4

**Solution** **Ans. (4)**

**Illustration 13** Calculate the degree of hydrolysis of a mixture containing 0.1N $NH_4OH$ and 0.1N HCN

If $K_a = 10^{-5}$ and $K_b = 10^{-5}$

**Solution** Salt is [WA – WB]

$$h = \sqrt{\frac{K_W}{K_a \times K_b}} = \sqrt{\frac{10^{-14}}{10^{-5} \times 10^{-5}}} = \sqrt{10^{-14} \times 10^{+10}} = \sqrt{10^{-4}} = 10^{-2}$$

**Illustration 14** ***Assertion :*** An aqueous solution of $NH_4NO_3$ is acidic in characters.

***Reason :*** $NH_4NO_3$ in an aqueous solution undergoes anionic hydrolysis.

(1) A (2) B (3) C (4) D

**Solution** **Ans. (3)**

## BEGINNER'S BOX-2

**1.** When sodium acetate ($CH_3COONa$) is added to aqueous solution of acetic acid ($CH_3COOH$), the–

(1) the pH value becomes zero
(2) pH value remains unchanged
(3) pH value decreases
(4) pH value increases

**2.** Which of the following cations is not hydrolyzed in aqueous solution ?

(i) $Mg^{2+}$ (ii) $Ca^{2+}$ (iii) $Na^+$ (iv) $K^+$

(1) (i), (ii) (2) (iii), (iv) (3) (i), (ii), (iii), (iv) (4) (i), (ii), (iii)

**3.** Which of the anions is not hydrolyzed in aqueous solution ?

| | | | |
|---|---|---|---|
| (i) $Cl^-$ | (ii) $NO_3^-$ | (iii) $Br^-$ | (iv) $ClO_4^-$ |
| (1) (i), (ii), (iii), (iv) | (2) (ii), (iii), (iv) | (3) (i), (ii), (iii) | (4) (ii), (iv) |

**4.** Which of the following salts does not undergo hydrolysis ?

(1) KCN (2) KCl (3) $NH_4NO_3$ (4) $FeCl_3.6H_2O$

**5.** Which of the following salts undergoes anionic hydrolysis ?

(1) $AlCl_3$ (2) $CuSO_4$ (3) $Na_2CO_3$ (4) $NH_4Cl$

**6.** For cationic hydrolysis, pH is given by–

(1) $pH = \frac{1}{2}pK_w + \frac{1}{2}pK_a + \frac{1}{2}\log C$ (2) $pH = \frac{1}{2}pK_w - \frac{1}{2}pK_b - \frac{1}{2}\log C$

(3) $pH = \frac{1}{2}pK_w + \frac{1}{2}pK_a - \frac{1}{2}pK_b$ (4) $pH = \frac{1}{2}pK_w + \frac{1}{2}pK_b + \frac{1}{2}\log C$

**7.** Which of the following salts is neutral in water ?

(1) KCl (2) $NH_4NO_3$ (3) $NH_4CN$ (4) $NH_4OH$

---

## 4.6 SOLUBILITY AND SOLUBILITY PRODUCT ($K_{sp}$)

### (A) SOLUBILITY

**(a) Definition** At constant temperature the maximum number of moles of solute which can be dissolved in a solvent to obtain 1 litre of solution (i.e. saturated solution) is called solubility.

$$S(M) = \frac{\text{Number of moles of solute}}{\text{Volume of solution (L)}} \qquad S = \frac{x}{M_w \times V_L}\ \text{mol L}^{-1}$$

$$S_{g/\ell} = S_M \times \text{molar mass}$$

### GOLDEN KEY POINTS

- Solubility does not depend on amount of substances and volume of solution where as depends on the following–
  - (i) Temperature
  - (ii) Presence of common ion
  - (iii) Nature of solvent (Molecular wt. of AgCl = 143.5)
    (Molecular wt. of $BaSO_4$ = 233)
- Solubility can be expressed in terms of molarity.

### (B) SOLUBILITY PRODUCT($K_{sp}$) :

When a sparingly soluble salt such as AgCl is put into water, a very small amount of AgCl dissolves in water and most of the salt remains undissolved in its saturated solution.

- The salt AgCl is an electrolyte, its dissociation occurs in solution. Hence, the quantity of AgCl that dissolves in water dissociates into $Ag^+$ and $Cl^-$ ions. Thus, in the saturated solution of AgCl an equilibrium exists between undissolved solid AgCl and its ions, $Ag^+$ and $Cl^-$ ions.

$$AgCl_{(s)} \underset{\text{Precipitation}}{\overset{\text{Dissolution}}{\rightleftharpoons}} Ag^+_{(aq)} + Cl^-_{(aq)}$$

according to law of mass action $K = \frac{[Ag^+].[Cl^-]}{[AgCl]}$

Since, the concentration of undissolved solid AgCl is constant. Thus, the product K.[AgCl] gives another constant which is designated as Ksp.

So, $K.[AgCl] = [Ag^+].[Cl^-]$ $\therefore$ $K_{sp} = [Ag^+] . [Cl^-]$

At constant temperature product of concentrations of ions in a saturated solution of substance is called solubility product of that substance. (Saturated solution is that solution in which solid solute and the ions in solution remain in equilibrium with each other.

- **$K_{sp}$ for $CaCl_2$** $CaCl_2(s) \rightleftharpoons Ca^{+2}(aq) + 2Cl^-(aq)$

  Solubility product in terms of concentration of ions $K_{sp} = [Ca^{+2}]\,[Cl^-]^2$
- **$K_{sp}$ for $AlCl_3$** $AlCl_3(s) \rightleftharpoons Al^{+3}(aq) + 3Cl^-(aq)$

  Solubility product in terms of concentration of ions $K_{sp} = [Al^{+3}]\,[Cl^-]^3$
- **General form** $AxBy(s) \rightleftharpoons xA^{+y}(aq) + yB^{-x}(aq)$

  $K_{sp} = [A^{+y}]^x\,[B^{-x}]^y$

## 4.7 APPLICATION OF SOLUBILITY PRODUCT ($K_{sp}$)

### (A) To find out the solubility (S) :

**(i) $K_{sp}$ of AB (Mono-mono, di-di, tri-tri valency) type salt –**

**Ex.** NaCl, $BaSO_4$, $CH_3COONa$, $CaCO_3$, NaCN, KCN, $NH_4CN$, $NH_4Cl$ etc.

| $AB(s) \rightleftharpoons$ | $A^+(aq)$ + | $B^-(aq)$ |
|---|---|---|
| a | 0 | 0 |
| (a–s) | s | s |

$K_{sp} = [A^+]\,[B^-]$

$K_{sp} = s^2$ or $s = \sqrt{K_{sp}}$

**(ii) $K_{sp}$ of $AB_2$ or $A_2B$ (Mono-di or di-mono valency) type salt –**

**Ex.** $CaCl_2$, $CaBr_2$, $K_2S$, $(NH_4)_2SO_4$, $K_2SO_4$, $K_2CO_3$ etc.

| $AB_2(s) \rightleftharpoons$ | $A^{+2}(aq)$ + | $2B^-(aq)$ |
|---|---|---|
| a | 0 | 0 |
| a – s | s | 2s |

$K_{sp} = [A^{+2}]\,[B^-]^2$

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

$$s = \left(\frac{K_{sp}}{4}\right)^{1/3}$$

**(iii) $K_{sp}$ of $AB_3$ or $A_3B$ (Mono-tri or tri-mono valency) type salt –**

**Ex.** $FeCl_3$, $AlCl_3$, $K_3PO_4$ etc.

| $AB_3(s) \rightleftharpoons$ | $A^{+3}(aq)$ + | $3B^-(aq)$ |
|---|---|---|
| a | 0 | 0 |
| a–s | s | 3s |

$$K_{sp} = \left[A^{+3}\right]\left[B^-\right]^3 = s \times (3s)^3 = 27s^4$$

$$s = \left(\frac{K_{sp}}{27}\right)^{1/4}$$

**(iv) $K_{sp}$ of $A_2B_3$ or $A_3B_2$ (Di-tri or tri-di valency) type salt**

**Ex.** $Al_2(SO_4)_3$, $Ba_3(PO_4)_2$ etc.

$$\begin{array}{llll} A_2B_3(s) & \rightleftharpoons 2A^{+3}(aq) & + & 3B^{-2}(aq) \\ a & 0 & & 0 \\ a-s & 2s & & 3s \end{array}$$

$$K_{sp} = \left[A^{+3}\right]^2\left[B^{-2}\right]^3 = 2s\times2s\times3s\times3s\times3s = 108s^5$$

$$s = \left(\frac{K_{sp}}{108}\right)^{1/5}$$

**(v) General form :**

$$\begin{array}{llll} A_xB_y(aq) & \rightleftharpoons xA^{+y}(aq) & + & yB^{-x}(aq) \\ a & 0 & & 0 \\ a-s & xs & & ys \end{array}$$

$K_{sp} = [A^{+y}]^x \,.\, [B^{-x}]^y$

$K_{sp} = (xs)^x.\,(ys)^y$ $\boxed{K_{sp} = x^x.y^y.s^{(x+y)}}$

## Illustrations

**Illustration 15** What will be the solubility product of following in terms of solubility (mol $L^{-1}$)

(i) $Al_2(SO_4)_3(s)$ (ii) $Na_2KPO_4(s)$ (iii) $NaKRbPO_4(s)$

**Solution** (i) $Al_2(SO_4)_3(s) \rightleftharpoons 2Al^{+3}(aq) + 3SO_4^{-2}(aq)$

$Ksp = 2^2 \times 3^3 \times (S)^{2+3} = 4 \times 27 \times S^5 = 108\,S^5$

(ii) $Na_2KPO_4(s) \rightleftharpoons 2Na^+(aq) + K^+(aq) + PO_4^{-3}(aq)$

$Ksp = 2^2 \times 1^1 \times 1^1(S)^{2+1+1} = 4S^4$

(iii) $NaKRbPO_4(s) \rightleftharpoons Na^+(aq) + K^+(aq) + Rb^+(aq) + PO_4^{-3}(aq)$

$Ksp = 1^1 \times 1^1 \times 1^1 \times 1^1 \times (S)^{1+1+1+1} = S^4$

**Illustration 16** If solubility product of the base $M(OH)_3$ is $2.7 \times 10^{-11}$, the concentration of $OH^{-1}$ will be

(1) $3 \times 10^{-3}$ (2) $3 \times 10^{-4}$ (3) $10^{-3}$ (4) $10^{-11}$

**Solution** **Ans. (1)**

**Illustration 17** ***Assertion :*** For a sparingly soluble salt, S is the maximum dissolved value of solute in a solution.

***Reason :*** $K_{sp}$ corresponds to the ionic product of the salt in a saturated solution.

(1) A (2) B (3) C (4) D

**Solution** **Ans. (2)**

**Illustration 18** The solubility of $BaSO_4$ in water is $1.07 \times 10^{-5}$ mol $dm^{-3}$. Estimate its solubility product.

**Solution** Solubility equilibrium for $BaSO_4$ is $BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO^{2-}_{(aq)}$

The solubility product is, $K_{sp} = [Ba^{2+}]\,[SO_4^{2-}]$

If S is the molar solubility of $BaSO_4$. then $K_{sp} = S^2$ because x = 1 = y.

Now, $S = 1.07 \times 10^{-5}$ M.

Hence, $K_{sp} = (1.07 \times 10^{-5})^2 = 1.145 \times 10^{-10}$

**Illustration 19** The solubility product of AgBr is $5.2 \times 10^{-13}$. Calculate its solubility in mol $dm^{-3}$ and g $dm^{-3}$. (Molar mass of AgBr. = 187.8 g $mol^{-1}$)

**Solution** The solubility equilibrium of AgBr is $AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

and $K_{sp} = [Ag^+]\,[Br^-] = S^2$ because x = 1 = y.

The molar solubility S of AgBr is given by $S = \sqrt{K_{sp}} = \sqrt{5.2\times10^{-13}} = 7.2 \times 10^{-7}$ mol $dm^{-3}$

The solubility in g $dm^{-3}$ = molar solubility (mol $dm^{-3}$) $\times$ molar mass (g $mol^{-1}$)

Hence, solubility = $7.2 \times 10^{-7}$ (mol $dm^{-3}$) $\times 187.8 = 1.35 \times 10^{-4}$ g $dm^{-3}$.

### (B) Condition of precipitation /Ionic product (IP or $Q_{sp}$) :

- Ionic product (IP) of an electrolyte is defined in the same way as $K_{sp}$. The only difference is that ionic product expression contains the initial concentration of ions or the concentration at any time whereas the expression of $K_{sp}$ contains only equilibrium concentration. Thus, for AgCl.
  $Q_{sp} = IP = [Ag^+]_i\,[Cl^-]_i$ and $K_{sp} = [Ag^+]_{eq}.[Cl^-]_{eq}$
- Ionic product changes with concentration but $K_{sp}$ does not. $K_{sp}$ is applicable for saturated solution of the sparingly soluble electrolyte.
- To decide whether an ionic compound will precipitate, its $K_{sp}$ is compared with the value of ionic product. The following three cases arise :
  (i) $Q_{sp} < K_{sp}$ : The solution is unsaturated and precipitation will not occur.
  (ii) $Q_{sp} = K_{sp}$ : The solution is saturated and solubility equilibrium exists.
  (iii) $Q_{sp} > K_{sp}$ : The solution is supersaturated and hence precipitation of the compound will occur.

  Thus, a salt is precipitated when its ionic product exceeds the solubility product of the salt.

### (C) Common ion effect on solubility :

**Important point :-** Solubility of substances always decreases in the presence of common ion. According to Le-Chatelier's principle on increasing common ion concentration equilibrium shifts in backward direction until the equilibrium is re-established so, the solubility of substances decreases but $K_{sp}$ remains same because it is an equilibrium constant which depends only on temperature.

## Illustrations

**Illustration 20** Find out the solubility of AgCl in the presence of C molar NaCl solution ?

**Solution.**

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

(Let solubility of AgCl is S mol $L^{-1}$)

$$Ksp = [Ag^+]\,[Cl^-]$$
$$Ksp = S^2$$

In NaCl solution

$$\underset{C}{NaCl} \longrightarrow \underset{C}{Na^+} + \underset{C}{Cl^-}$$

Let solubility of AgCl in the presence of NaCl solution is S' mol $L^{-1}$.

$$\underset{S'}{AgCl} \rightleftharpoons \underset{S'}{Ag^+} + \underset{S'+C}{Cl^-}$$

According to L.M.A.

$Ksp = [Ag^+]'\,[Cl^-]'$

$Ksp = S'\,(S' + C) = S'^2 + S'C$ (Neglecting the higher power terms of S')

$Ksp = S'\,C$ $\quad \boxed{S' = \dfrac{Ksp}{C}}$

**Illustration 21** Find out the solubility of $CaCl_2$ solution in the presence of C NaCl solution ?

**Solution**

$$\underset{S}{CaCl_2} \rightleftharpoons \underset{S}{Ca^{+2}} + \underset{2S}{2Cl^-}$$

[Let solubility of $CaCl_2$ is S mol $L^{-1}$]

$Ksp = [Ca^{+2}]\,[Cl^-]^2 = 4S^3$

For NaCl solution

$$\underset{C}{NaCl} \longrightarrow \underset{C}{Na^+} + \underset{C}{Cl^-}$$

Let solubility of $CaCl_2$ in the presence of NaCl solution is S' mol $L^{-1}$.

$$\underset{S'}{CaCl_2} \rightleftharpoons \underset{S'}{Ca^{+2}} + \underset{2S'+C}{2Cl^-}$$

According to L.M.A.

$Ksp = [Ca^{+2}]'\,[Cl^-]'^2$

$Ksp = S'\,(2S' + C)^2 = S'\,(4S'^2 + 4S'C + C^2)$

$Ksp = 4S'^3 + 4S'^2C + S'C^2$ (Neglecting the higher power terms of S')

$$\boxed{S' = \frac{Ksp}{C^2}}$$

**Illustration 22** Find out the solubility of NaCl in the presence of C $CaCl_2$ solution ?

**Solution**

$$\underset{C}{CaCl_2} \longrightarrow \underset{C}{Ca^{+2}} + \underset{2C}{2Cl^-}$$

Let solubility of NaCl in the presence of $CaCl_2$ solution is S' mol $L^{-1}$.

$$\underset{S'}{NaCl} \rightleftharpoons \underset{S'}{Na^+} + \underset{S'+2C}{Cl^-}$$

According to L.M.A.

$Ksp = [Na^+]' [Cl^-]'$

$Ksp = S' (S' + 2C) = S'^2 + 2S'C$ (Neglecting the higher power terms of S')

$$\boxed{S' = \frac{Ksp}{2C}}$$

## GOLDEN KEY POINTS

*● **Group precipitation of Salt**

| Group | Radicals | Condition for precipitation (Group reagent) | Forms of precipitation |
|---|---|---|---|
| Zero | $NH_4^+$ | ........... | ........... |
| First | $Pb^{+2}, Hg^{+1}, \left(Hg_2^{+2}\right), Ag^+$ | By mixing of dilute HCl | Chloride<br>$AgCl, Hg_2Cl_2, PbCl_2$ |
| Second II(A) | $Pb^{+2}, Cu^{+2} Hg^{+2}, Cd^{+2}, Bi^{+3}$ | $H_2S$ gas passed in the presence of acidic medium | Sulphide |
| Second II(B) | $As^{+3}, Sb^{+3}, Sn^{+2}, Sn^{+4}$ | | $PbS, HgS, CuS, CdS,$<br>$SnS, SnS_2, As_2S_3,$<br>$Sb_2S_3, Bi_2S_3$ |
| Third | $Al^{+3}, Cr^{+3}, Fe^{+3}$ | $NH_4OH$ mixed in the presence of $NH_4Cl$ | Hydroxide<br>$Al(OH)_3, Fe(OH)_3, Cr(OH)_3$ |
| Fourth | $Zn^{+2}, Ni^{+2}, Mn^{+2}, Co^{+2}$ | $H_2S$ gas passed in presence of basic medium | Sulphide<br>$MnS, CoS, NiS, ZnS$ |
| Fifth | $Ba^{+2}, Sr^{+2}, Ca^{+2}$ | $(NH_4)_2CO_3$ mixed in the presence of $NH_4Cl$ | Corbonate<br>$BaCO_3, SrCO_3, CaCO_3$ |
| Sixth | $Na^+, K^+, Mg^{+2}$ | By mixing of $Na_2HPO_4$ | Hydrogen phosphate ($MgHPO_4$) |

- $Hg^{+1}$ (us) → Unstable in aqueous solution.

  [Stable in dimer form ($Hg_2^{+2}$) ⇒ $Hg_2Cl_2$]

  $$\underset{S}{Hg_2Cl_2} \rightleftharpoons \underset{S}{Hg_2^{+2}} + \underset{2S}{2Cl^-}$$

  $K_{sp} = [Hg_2^{+2}] [Cl^-]^2$

  $= (S) (2S)^2$

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

- $Hg^{+2}$ (ic) → $HgCl_2$ → stable in aqueous solution

  $$\underset{S}{HgCl_2} \rightleftharpoons \underset{S}{Hg^{+2}} + \underset{2S}{2Cl^-}$$

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

- For precipitation of II group, $H_2S$ gas is passed in acidic medium to decrease $S^{-2}$ concentration by common ion effect. So that ionic product of cations of group II and $S^{-2}$ ions exceed the solubility product of their corresponding metal sulphides. Therefore only the cations of group II gets precipitated. But the sulphides of the cation of the other groupes (III, IV & V) are not precipitated under these conditions because their solubility products are quite high.

  When $H_2S$ gas is directly passed through solution then IVth group is also precipitated with IInd group.

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

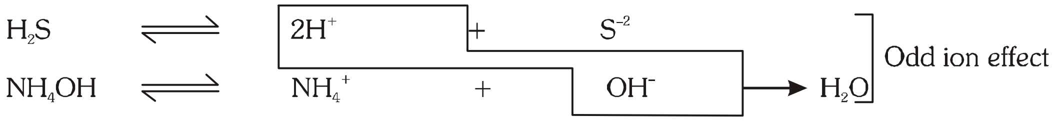
  $Ksp_{II} < Ksp_{IV} < [\text{radicals of II}^{nd} \text{ and IV group}]\,[S^{-2}]$

  When $H_2S$ gas pass in acidic medium $[H^+]$ then only IInd group radicals are precipitated.

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

  $Ksp_{II} < [\text{radicals of II}^{nd} \text{ and IV group}]\,[S^{-2}] < Ksp_{IV}$

- For precipitation of group III, $NH_4OH$ is added in the presence of $NH_4Cl$ to decrease the concentration of $OH^-$ by common ion effect so that ionic product of cations of group III and hydroxides ions exceed the solubility product of their corresponding metal hydroxides and only group III cations will precipitated. On the other hand cations of group IV, V and VI, which require large concentration of $OH^-$ due to their high solubility product, will not be precipitated.
- In IV group, $H_2S$ gas is passed in basic medium to increase $S^{-2}$ ion concentration by odd ion effect, so that the ionic product of cations of group IV and $S^{-2}$ ions exceed the solubility product of their corresponding metal sulphide and hence gets precipitated.



So $[S^{-2}]\uparrow$

$Ksp_{IV} < [\text{Radicals of group IV}]\,[S^{-2}]$

---

### (D) Precipitation of soap and salt (NaCl)

#### (i) Precipitation of NaCl (common salt) :

For the precipitation of NaCl, HCl gas passed in saturated solution of NaCl to increase the concentration of $Cl^-$ ion by common ion, so ionic product of concentration of $Na^+$ and $Cl^-$ ion becomes more than Ksp of NaCl and NaCl easily precipitated.

Saturated solution of NaCl $\quad NaCl \rightleftharpoons Na^+ + Cl^-$

$Ksp = [Na^+]\,[Cl^-]$

By passing HCl gas $\quad HCl \longrightarrow H^+ + Cl^-$

due to common ion $\quad Ksp < [Na^+]\,[Cl^-]\uparrow$

#### (ii) Precipitation of soap :

For precipitation of soap, NaCl mixed in saturated solution of soap to increase the concentration of $Na^+$ ion by common ion. So ionic product of concentration of $Na^+$ and strearate ions is more than Ksp of soap and soap is easily precipitated.

**Ex.** $C_{17}H_{35}COONa$ (Sodium stearate)

Saturated solution of $C_{17}H_{35}COONa \Rightarrow C_{17}H_{35}COONa^+ \rightleftharpoons C_{17}H_{35}COO^- + Na^+$

$Ksp = [C_{17}H_{35}COO^-]\,[Na^+]$

By mixing NaCl solution $\quad NaCl \longrightarrow Na^+ + Cl^-$

due to common ion $\quad Ksp < [C_{17}H_{35}COO^-]\,[Na^+]\uparrow$

## Illustrations

**Illustration 23** Solubility products of $M(OH)_3$ and $M(OH)_2$ are $10^{-23}$ and $10^{-14}$ respectively. Which will be precipitated first on adding $NH_4OH$, if $M^{+2}$ and $M^{+3}$ both the ions are in solution ?

(1) $M^{+2}$ (2) $M^{+3}$

(3) Both $M^{+2}$ and $M^{+3}$ together (4) Precipitation will not take place.

**Solution** **Ans. (2)**

**Illustration 24** Solubility of $CaCl_2$ is $4 \times 10^{-8}$, then find out its $K_{sp}$ and its new solubility in the presence of $10^{-2}$ M $Ca(OH)_2$.

**Solution**

$$CaCl_2 \rightleftharpoons Ca^{+2} + 2Cl^-$$

$$\begin{array}{lll} 1 & s & 2s \end{array} \quad \text{(Initial solubility = s)}$$

$$K_{sp} = 4s^3 = 4(4 \times 10^{-8})^3 = 256 \times 10^{-24} \qquad \text{.......(i)}$$

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

$$\begin{array}{lll} C & 0 & 0 \\ 0 & C & 2C \end{array}$$

$$CaCl_2 \rightleftharpoons Ca^{+2} + 2Cl^-$$

$$\begin{array}{lll} 1 & s' + C & 2s' \end{array} \quad \text{(New solubility = s')}$$

$$K_{sp} = [Ca^{+2}]'\,[Cl^-]^{2\,'}$$

$$= [s' + C]\,[2s']^2 = 4s'^3 + 4s'^2C \qquad (s'^3 = \text{negligible})$$

$$K_{sp} = 4s'^2C \qquad \text{.......(ii)}$$

From equation (i) and (ii) $s'^2 = \dfrac{Ksp}{4C} = \dfrac{256 \times 10^{-24}}{4 \times 10^{-2}} = 64 \times 10^{-22}$

$$s' = 8 \times 10^{-11} \text{ mol } L^{-1}$$

## 4.8 FEW IMPORTANT POINTS

**(A) Isohydric solution –** If different–different solution has same pH are called isohydric solution.

**Sp. point :–** Isohydric conditions for two given weak acids $HA_1$ and $HA_2$ which has concentration $C_1$ and $C_2$ and ionisation constants are $Ka_1$ and $Ka_2$ respectively.

Then $[H^+]_1 = \sqrt{Ka_1\,C_1}$ and $[H^+]_2 = \sqrt{Ka_2\,C_2}$

According to given condition for isohydric solution pH of different solution is same.

Hence $[H^+]_1 = [H^+]_2$

$$\sqrt{Ka_1\,C_1} = \sqrt{Ka_2\,C_2}$$

$$\boxed{Ka_1C_1 = Ka_2C_2} \text{ or If n = 1 mole then } \boxed{\frac{Ka_1}{V_1} = \frac{Ka_2}{V_2}} \because \boxed{C \propto \frac{1}{V}}$$

**(B) Relative strength of Acids :–** It indicates that how many times an acid is stronger than the other acid.

We know $\boxed{\text{Strength of acid} \propto [H^+]}$

If there are two weak acids $HA_1$ and $HA_2$ which has concentration $C_1$ and $C_2$, degree of ionisation $\alpha_1$ and $\alpha_2$ and ionisation constants $Ka_1$ and $Ka_2$ respectively then ratio of their strength of acids.

We know $[H^+] = C\alpha = \sqrt{KaC}$

$$\text{So } \frac{\text{Strength of weak acid } HA_1}{\text{Strength of weak acid } HA_2} = \frac{[H^+]_1}{[H^+]_2} = \frac{C_1\alpha_1}{C_2\alpha_2} = \frac{\sqrt{Ka_1C_1}}{\sqrt{Ka_2C_2}}$$

If $C_1 = C_2$

$$\boxed{\text{R.S.} = \frac{\text{Strength of weak acid } HA_1}{\text{Strength of weak acid } HA_2} = \sqrt{\frac{Ka_1}{Ka_2}} = \frac{\propto_1}{\propto_2}}$$

Thus the relative strength of two acids of equimolar concentration can be compared by taking square root of the ratio of their ionisation constants.

## Illustrations

**Illustration 25** Which of the following solvents will undergo self-ionization ?

(1) $H_2O$ (2) $NH_3$ (3) HF (4) All of these

**Solution** **Ans. (4)**

**Illustration 26** For two acids A and B, $pKa_1 = 1.2$, $pKa_2 = 2.8$ respectively in value, then which is true:-

(1) A & B both are equally acidic

(2) A is stronger than B

(3) B is stronger than A

(4) None of these

**Solution** **Ans. (2)**

**Illustration 27** pH values of two acids A and B are 4 and 5. The strengths of these two acids are related as :-

(1) The strengths of the two acids cannot be compared.

(2) Acid B is 10 times stronger than acid A.

(3) Strength of acid A : Strength of acid B = 4 : 5

(4) Acid A is 10 times stronger than acid B.

**Solution** **Ans. (4)**

## BEGINNER'S BOX-3

**1.** The units of solubility product of silver chromate ($Ag_2CrO_4$) will be–

(1) $mol^2L^{-2}$ (2) $mol^3L^{-3}$ (3) $mol\ L^{-1}$ (4) $mol^{-1}L$

**2.** At a certain temperature, the solubility of the salt $A_xB_y$ is S moles per liter. The general expression for its solubility product will be–

(1) $K_{sp} = x^y y^x S^{x+y}$ (2) $K_{sp} = (xy)^{x+y} S^{x+y}$ (3) $K_{sp} = (x^x y^y) S^{x+y}$ (4) $K_{sp} = x^x y^y S^{xy}$

**3.** The molar solubility of silver sulphate is $1.5 \times 10^{-2}$ mol $L^{-1}$. The solubility product of the salt will be–

(1) $2.25 \times 10^{-4}$ (2) $1.35 \times 10^{-5}$ (3) $1.7 \times 10^{-6}$ (4) $3.0 \times 10^{-3}$

**4.** The precipitate of $CaF_2$ ($K_{sp} = 1.7 \times 10^{-10}$) is obtained when equal volumes of the following are mixed:

(1) $10^{-3}$ M $Ca^{2+}$ + $10^{-5}$ M $F^-$ (2) $10^{-5}$ M $Ca^{2+}$ + $10^{-3}$ M $F^-$

(3) $10^{-2}$ M $Ca^{2+}$ + $10^{-3}$ M $F^-$ (4) $10^{-4}$ M $Ca^{2+}$ + $10^{-4}$ M $F^-$

**5.** If $S_0$, $S_1$, $S_2$ and $S_3$ are the solubilities of AgCl in water, 0.01 M $CaCl_2$, 0.01 M NaCl and 0.5 M $AgNO_3$ solutions, respectively, then which of the following is true?

(1) $S_0 > S_2 > S_1 > S_3$ (2) $S_0 = S_2 = S_1 > S_3$

(3) $S_3 > S_1 > S_2 > S_0$ (4) $S_0 > S_2 > S_3 > S_1$

**6.** Given $K_{sp}$ (AgI) = $8.5 \times 10^{-17}$. The solubility of AgI in 0.1 M KI solution is–

(1) 0.1 M (2) $8.5 \times 10^{-16}$ M

(3) $8.5 \times 10^{-17}$ M (4) $8.5 \times 10^{-18}$ M

## 4.9 pH

### 1. INTRODUCTION

**pH of some important substance :-**

(i) Blood ⟶ 7.4

(ii) Tear ⟶ 7.4

(iii) Gastric Juice ⟶ 1 to 3

(iv) Soft drink ⟶ 2 to 4

(v) Acidic rain ⟶ 6

(vi) Milk ⟶ 6.3 to 6.6

(vii) Water ⟶ 7

### 2. TYPES OF pH

(i) pH of SA and SB. $[H^+] = C$, $[OH^-] = C$

(ii) pH of WA and WB. $[H^+] = C\alpha$ , $[OH^-] = C\alpha$

(iii) pH of very dilute solution. (Consider $H^+$ & $OH^-$ of $H_2O$ also)

(iv) pH of salt solution.

- (a) SA SB Type Salt (always 7)
- (b) SA WB Type Salt (< 7)
- (c) WA SB Type Salt (> 7)
- (d) WA WB Type Salt (Almost 7)

### 3. pH of mixture of acid and base.

(a) pH of mixture of strong acids :

$N_1V_1 + N_2V_2 + N_3V_3 + \ldots\ldots\ldots\ldots\ldots\ldots = NV$

V = Volume of final solution $= V_1 + V_2 + V_3 + \ldots\ldots\ldots$

N = Normality of final solution = $[H^+]$ in final solution.

(b) pH of mixture of strong base :

$N_1V_1 + N_2V_2 + N_3V_3 + \ldots\ldots\ldots\ldots\ldots\ldots = NV$

V = Volume of final solution $= V_1 + V_2 + V_3 + \ldots\ldots\ldots$

N = Normality of final solution = $[OH^-]$ in final solution.

(c) pH of mixture of strong acids and strong bases :

For acid :

$$N_1V_1 + N_2V_2 + N_3V_3 + \ldots\ldots\ldots\ldots\ldots\ldots = (NV)_{Acid}$$

For base :

$$N_1V_1 + N_2V_2 + N_3V_3 + \ldots\ldots\ldots\ldots\ldots\ldots = (NV)_{Base}$$

(i) If $(NV)_{Acid} > (NV)_{Base}$ then solution is acidic.

$NV = (NV)_{Acid} - (NV)_{Base}$ and

$[H^+] = N$

(ii) If $(NV)_{Base} > (NV)_{Acid}$ then solution is basic.

$NV = (NV)_{Base} - (NV)_{Acid}$ and

$[OH^-] = N$

(iii) If, $(NV)_{acid} = (NV)_{base}$ then solution is neutral.

(d) $p^H$ of mixture of (WA + SB); when equivalent of WA > SB]

pH of mixture of (SA + WB); When equivalent of WB > SA]

⇓

gives rise to concept of buffer solution

## 4.10 BUFFER SOLUTION

**(A) Definition** : A solution which resist the change in pH and pH does not change significantly on addition of small amount of strong acid or strong base is called buffer solution.

**(B) Properties of buffer solution :**

**(i)** The pH of buffer solution does not change appreciably upon the addition of small amount of either strong acid or strong base.

**(ii)** The pH of buffer solution does not depend on the volume of solution. Hence, solution can be diluted without change in pH.

**(iii)** The pH of buffer solution remains constant even if it is kept for a long time.

**(C) Buffer solutions are used in**

(a) Qualitative analysis of mixture

(b) Quantitative analysis of mixture

(c) Digestion of food

(d) Preservation of food and fruits.

(e) Agriculture and dairy product preservation.

(f) Industrial process such as manufacture of paper, dyes inks, paints, drugs etc.

**TYPES OF BUFFER SOLUTION**

(A). Simple buffer solution

(B). Mixed buffer solution :- (i) Acidic buffer solution (ii) Basic buffer solution

**(A) Simple buffer solution :-** Aqueous solution of weak acid-weak base (WA – WB) types of salt.

**Ex.** $CH_3COONH_4$, $NH_4CN$, AgCN etc.

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

pH does not depend on concentration.

$$CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + NH_4OH.$$

**Buffer Action :**

**Case 1**

When mixing of acid [$H^+$]

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

$$NH_4OH + H^+ \rightleftharpoons NH_4^+ + H_2O$$

**Case 2.**

When mixing of base [$OH^-$]

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

$$CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$$

**(B) Mixed buffer solution :**

**(i) Acidic buffer solution :-** The solution in which weak acid and its conjugate base are present.

**OR**

Aqueous solution of mixture of weak acid and salt of same weak acid with any strong base is called acidic buffer solution.

**Ex.** $\underset{WA}{CH_3COOH} + \underset{WASB}{CH_3COONa}$

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

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+$$

$$\left[\underset{WA}{CH_3COOH} + \underset{C\,Base}{CH_3COO^-}\right] + Na^+ \text{ (Spectator ion)}$$

**BUFFER ACTION :-**

**Case 1.** If a small amount of acid is added to the buffer solution. The $H^+$ ions of the acid react with $CH_3COO^-$ ions and produce $CH_3COOH$ and most of the added $H^+$ ions are consumed so there is no appreciable change in pH.

$$CH_3COO^-_{(aq)} + H^+ \rightleftharpoons CH_3COOH_{(aq)}$$

**Case 2.** If small amount of base is added to the buffer solution, the $OH^-$ ions are consumed by $CH_3COOH$.

$$CH_3COOH_{(aq)} + OH^-_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H_2O_{(\ell)}$$

So, there is no appreciable change in pH.

**(a) pH of acidic buffer solution :**

$$\begin{array}{ccc} CH_3COOH & + & CH_3COONa \\ \text{Acid} & + & \text{Salt} \end{array}$$

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

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+$$

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

or $$[H^+] = \frac{Ka[CH_3COOH]}{[CH_3COO^-]} = \frac{Ka[\text{Acid}]}{[\text{Conjugate base}]}$$

taking –log on both sides

$$pH = pKa - \log\frac{[\text{Acid}]}{[\text{Conjugate base}]} \quad \text{or} \quad pH = pKa + \log\frac{[\text{Conjugate base}]}{[\text{Acid}]}$$

**Henderson's equation :**

$$\boxed{pH = pKa + \log\frac{[\text{Salt}]}{[\text{Acid}]} \text{ or } \frac{[\text{Conjugate base}]}{[\text{Acid}]}}$$

- $[\text{Conjugate base}]_{eq} \simeq [\text{Salt}]$ because $CH_3COO^-$ mainly comes from salt since dissociation of $CH_3COOH$ in presence of $CH_3COONa$ is appreciably decreased.
- $[\text{Acid}]_{aq} \simeq$ initial concentration of acid since it is almost unionised in presence of salt due to common ion effect.

**(b) pH range of acidic buffer solution :** It depends on $pK_a$ of acid and ratio of salt to acid concentrations.

$$[CH_3COOH] : [CH_3COONa] \Rightarrow pH = pKa + \log\frac{[CH_3COONa]}{[CH_3COOH]}$$

(i) If, 1 : 10 $\Rightarrow$ $pH = pKa + \log\frac{10}{1} = pKa + 1$

(ii) If, 10 : 1 $\Rightarrow$ $pH = pKa - 1$

So pH range

$$\boxed{pH = pKa \pm 1}$$

**(c) Maximum buffer action condition of acidic buffer solution :**

$[CH_3COOH]$ : $[CH_3COONa]$

1 : 1 $\Rightarrow$ $pH = pK_a + \log\frac{1}{1}$

$$\boxed{pH = pKa}$$

**(ii) Basic buffer solution :**

**Definition :** The solution in which weak base and its conjugate acid are present. OR

Aqueous solution of mixture of weak base and salt of same weak base with any strong acid is called basic buffer solution.

**Ex.** $NH_4OH + NH_4Cl$

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

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-$$

$$\left[\underset{\text{W.B.}}{NH_4OH} + \underset{\text{C.acid}}{NH_4^+}\right] + Cl^-[\text{Spectator ion}]$$

**BUFFER ACTION :-**

**Case 1.** If a small amount of the acid is added to the buffer solution, the $H^+$ ions are consumed by $NH_4OH$.

$$NH_4OH_{(aq)} + H^+_{(aq)} \rightleftharpoons NH^+_{4(aq)} + H_2O_{(\ell)}$$

So there is no appreciable change in pH.

**Case 2.** If small amount of base is added to the buffer solution, the $OH^-$ ions react with $NH_4^+$ ions to produce $NH_4OH$.

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightleftharpoons NH_4OH_{(aq)}$$

Hence most of the added $OH^-$ ions are consumed. so there is no appreciable change in pH.

**(a) pOH of basic buffer solution :**

$$[\underset{\text{Base}}{NH_4OH} + \underset{\text{Salt}}{NH_4Cl}]$$

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

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} \quad \text{or} \quad [OH^-] = \frac{K_b[NH_4OH]}{[NH_4^+]}$$

taking –log on both sides $\quad pOH = pK_b + \log\frac{[NH_4^+]}{[NH_4OH]}$

**Henderson's equation :**

$$\boxed{pOH = pK_b + \log\frac{[\text{Salt}]}{[\text{Base}]} \text{ or } \frac{[\text{Conjugate acid}]}{[\text{Base}]}}$$

**(b) pOH range of basic buffer solution :** It depends on $pK_b$ of base and ratio of salt to base concentrations.

$[NH_4OH] : [NH_4Cl] \Rightarrow pOH = pK_b + \log\frac{[NH_4Cl]}{[NH_4OH]}$

(i) If, 1 : 10 $\Rightarrow$ $pOH = pK_b + 1$

(ii) If, 10 : 1 $\Rightarrow$ $pOH = pK_b - 1$

So, pOH range : $\boxed{pOH = pK_b \pm 1}$

Z:\NODE02\B0AI-B0\TARGET\CHEM\ENG\MODULE-1\4.IONIC EQUILIBRIUM\01-THEORY.P65

**(c) Maximum buffer action condition of basic buffer solution :**

$$\begin{array}{ccc} NH_4OH & : & NH_4Cl \\ 1 & : & 1 \end{array}$$

$$\boxed{pOH = pK_b}$$

### 3. BUFFER CAPACITY

**Definition :-**

(i) It is defined as the number of moles of strong acid (or strong base) added to one litre of a buffer solution to change its pH by one unit.

(ii) It measures the effectiveness of a buffer.

(iii) Larger the value of buffer capacity more resistant is the solution to pH change.

$$\text{Buffer capacity} = \frac{\text{Number of moles of acid or base added per litre}}{\text{Change in pH of buffer solution}}$$

## Illustrations

**Illustration 28** When 2 moles of HCl is added to 1 L. of an acidic buffer solution, its pH changes from 3.9 to 3.4. Find its buffer capacity.

**Solution** $\text{B.C.} = \frac{2}{0.5} = 4$

## BEGINNER'S BOX-4

**1.** A buffer solution is one which has–

(1) reserved acid (2) reserved base (3) constant pH (4) pH equal to 7

**2.** Which of the following solutions cannot act as a buffer system ?

(1) $KH_2PO_4/H_3PO_4$ (2) $NaClO_4/HClO_4$ (3) $C_5H_5N/C_5H_5N^+HCl^-$ (4) $Na_2CO_3/NaHCO_3$

**3.** A buffer solution can not be prepared by mixing equimolar amounts of–

(1) $B(OH)_3$ and $Na_2B_4O_7.10H_2O$ (2) $NH_3$ and $NH_4Cl$

(3) HCl and NaCl (4) $CH_3COOH$ and $CH_3COONa$

**4.** Which of the following salt solution will act as a buffer ?

(1) $CH_3COONH_4$ (aq.) (2) $NH_4Cl$ (aq.) (3) $CH_3COONa$ (aq.) (4) NaCl (aq.)

**5.** Which of the following combinations will make a buffer solutions ?

(i) $CH_3COONa$ (2 mol) + HCl (1 mol) (ii) $CH_3COOH$ (2 mol) + NaOH (1 mol)

(iii) $CH_3COOH$ (1 mol) + $CH_3COONa$ (1 mol)

(1) (iii) (2) (i), (ii) (3) (ii), (iii) (4) (i), (ii), (iii)

**6.** The pH of blood circulating in a human body is maintained around 7.4 by the action of the buffer system–

(1) $CH_3COOH/CH_3COONa$ (2) $NH_4Cl/NH_3$

(3) $H_2PO_4^{2-}$ (4) $H_2CO_3/HCO_3^-$

## 4.11 INDICATOR

**(A) Definition :-** An indicator is a substance that undergoes change in its colour to show end point of an acid-base titration.

- Titration involves neutralization of an acid and base.

**(B) Types of indicators :**

**(a) Acidic indicator (HIn)**

**Ex.** Phenolphthalein (HPh)

$$\underset{\text{colourless}}{HPh} \rightleftharpoons H^+ + \underset{\text{pink}}{Ph^-}$$

**(b) Basic indicator (InOH)**

**Ex.** Methyl orange (MeOH)

$$\underset{\text{yellow}}{MeOH} \rightleftharpoons \underset{\text{orange red}}{Me^+} + OH^-$$

**OSTWALD THEORY OF INDICATORS**

(i) According to this theory, all the indicators are weak organic acid or base.

(ii) Every indicator has light colour in its unionised form and dark colour in ionised form.

(iii) Every indicator has one colour in acidic medium and an altogether different colour in basic medium due to common ion effect and odd ion effect.

**Ex.** Acidic indicator [HPh]

**Case 1.** $HPh \rightleftharpoons H^+ + Ph^-$

In acidic medium $[H^+]$ $[HPh] > [Ph^-]$

Colourless due to common ion effect.

**Case 2.** $HPh \rightleftharpoons H^+ + Ph^-$

In basic medium $[OH^-]$ $[Ph^-] > [HPh]$

Pink colour due to odd ion effect.

**(a) pH of acidic indicator -**

$$HIn \rightleftharpoons H^+ + In^-$$

$$K_a \text{ or } K_I = \frac{[H^+][In^-]}{[HIn]}$$

$$[H^+] = \frac{K_I[HIn]}{[In^-]}$$

Taking – log on both sides

$pH = pK_I - \log [HIn] + \log [In^-]$

$$pH = pK_I + \log\frac{[In^-]}{[HIn]} \qquad \boxed{pH = pK_I + \log\frac{[\text{Ionised form}]}{[\text{Unionised form}]}}$$

**(b) pOH of basic indicator -**

$$InOH \rightleftharpoons In^+ + OH^-$$

$$K_b \text{ or } K_I = \frac{[In^+][OH^-]}{[InOH]} \qquad \boxed{pOH = pK_I + \log\frac{[\text{Ionised form}]}{[\text{Unionised form}]}}$$

**(c) pH range of indicators :**

Indicators show colour change for a certain pH range not at a definite pH which is called pH range of indicator.

**(i) pH range of acidic indicator :**

[HIn] : $[In^-]$
1 : 10
10 : 1 $\boxed{pH = pK_I \pm 1}$

**(ii) pOH range of basic indicator :**

[InOH] : $[In^+]$
1 : 10
10 : 1 $\boxed{pOH = pK_I \pm 1}$

**(iii) Neutral condition for indicators (No working condition)**

For acidic indicator
[HIn] : $[In^-]$
1 : 1

$\boxed{pH = pK_I}$

For basic indicators
[InOH] : $[In^+]$
1 : 1

$\boxed{pOH = pK_I}$

| S.No. | Name of indicator | Colour in acidic medium | Colour in basic medium | Working pH range of indicators |
|---|---|---|---|---|
| **1.** | Methyl orange (MeOH) | Orange red | Yellow | 3.1 to 4.5 |
| **2.** | Methyl red | Red | Yellow | 4.2 to 6.2 |
| **3.** | Phenol red | Yellow | Red | 6.2 to 8.2 |
| **4.** | Phenolphthalein (HPh) | Colourless | Pink | 8.2 to 10.2 |

## 3. ACID-BASE TITRATION

| S.No. | Type of titration | pH range of titration | Suitable indicators |
|---|---|---|---|
| **1.** | SA/SB. | 3 – 11 | All indicators (MeOH, HPh etc. ) |
| **2.** | SA/WB | 3 – 7 | Methyl orange (MeOH)<br>and methyl red |
| **3.** | WA/SB | 7 – 11 | Phenolphthalein (HPh) |

**Important points of titration :**

(i). Titration is always possible in two opposite solution. (i.e. one is acidic and other is basic).

(ii). For any titration only that indicator is suitable if their working pH range is in pH range of titration.

(iii). At the end point of titration equivalents of acids and bases are always same.

(iv). At the euivalence point, nature of solution depends on the type of titration.

SA/SB - Neutral (pH = 7)
SA/WB - Acidic
WA/SB - Basic
WA/WB - Any one possible

## 4.12 ACID AND BASE

### (1) INTRODUCTION

The earliest classification of substances into acids and bases was bases upon the characteristic properties possessed by each one of them. For example,

Acid was defined as a substance whose aqueous solution possesses the following characteristic properties:

(i) conducts electricity,

(ii) reacts with active metals like zinc, magnesium etc. to give hydrogen,

(iii) turns blue litmus red.

(iv) has a sour taste.

(v) whose acidic properties disappear on reaction with a base.

Base was defined as a substance whose aqueous solution possessed the following characteristic properties:

(i) conducts electricity,

(ii) reacts with active metals like zinc, magnesium etc. to give hydrogen,

(iii) turns red litmus blue,

(iv) has a bitter taste

(v) whose basic properties are destroyed on reaction with an acid.

The above definitions of acids and bases are called operational definitions as they are based certain operations (i.e. tests) to be performed on the substances. However, these have been replaced by conceptual definitions (put forward by Arrhenius, Bronsted-Lowry and Lewis) which go into the causes of the observed behavior, based upon structure and composition of the substances.

### (2) Arrhenius Concept (1884)

**(a)** **Acid:-** According to this concept, those substances which produce free $H^+$ ions in aqueous solution are called acid.

**Example** $HCl$, $HNO_3$, $H_2SO_4$, $H_3PO_4$, $H_2CO_3$, $H_2S$, $CH_3COOH$ etc.

$$HCl + H_2O \longrightarrow H^+(aq.) + Cl^-(aq.)$$

$$H_2SO_4 + H_2O \longrightarrow 2H^+(aq.) + SO_4^{-2}(aq.)$$

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^-(aq.) + H^+(aq.)$$

**(b)** **Base :-** Those substances which produce free $OH^-$ ions in aqueous solution are called base.

**Example** $NaOH$, $KOH$, $Cs(OH)$, $Rb(OH)$, $NH_4OH$, $Ba(OH)_2$, $Ca(OH)_2$, $Al(OH)_3$ etc.

$$NaOH + H_2O \longrightarrow Na^+ (aq.) + OH^- (aq.)$$

$$Ba(OH)_2 + H_2O \longrightarrow Ba^{+2} (aq.) + 2OH^- (aq.)$$

**(c)** **Nature of water :-** According to this concept nature of water is neutral and act as a solvent.

**(d)** **Neutralisation Reaction :-** Those reactions in which acid and base react together to form water molecule are called neutralisation reactions.

i.e. $Na^+ + OH^- + H^+ + Cl^- \longrightarrow NaCl + H_2O$

$H^+ + OH^- \longrightarrow H_2O$

**(e)** **Strength of acids and bases :-** This concept explains the strength of acids and bases depending upon the basis of degree of ionisation.

**Example** For strong electrolytes $\alpha \simeq 100\%$

For weak electrolytes $\alpha < 100\%$

**(f)** **Advantage :-** This concept explains the acids and bases practically. i.e. To find out the pH, ionisation constant, hydrolysis constants, heat of neutralisation etc.

**(g) Disadvantage :-**

It explains the behaviour of acids and bases only in aqueous (water) solvents.

## Illustrations

**Illustration 29** Gaseous hydrogen chloride is a very poor conductor of electricity but a solution of hydrogen chloride in water is a good conductor. This is due to the fact that :-

(1) Water is a good conductor of electricity

(2) Hydrogen chloride ionises in water

(3) A gas cannot conduct electricity but a liquid can

(4) HCl does not obey Ohm's law where as the solution does

**Solution** (2)

**Illustration 30** Which is acid in the following pairs according to Arrhenius concept ?

(1) HCl(g) and HCl (aq) (2) $CH_3COOH(l)$ and $CH_3COOH(aq)$

**Solution** (1) HCl(aq.) (2) $CH_3COOH(aq.)$

---

### (3) Bronsted-Lowry Concept (1923)

It is based upon the exchange of proton.

**(a) Acid :-** According to this concept those substances which have tendency to donate the proton ($H^+$) by any method in any solvent are called acids.

**Example** (i) Neutral molecules - HCl, $HNO_3$, $H_2SO_4$, $H_2CO_3$, $H_2S$, $CH_3COOH$, $H_3PO_3$ etc.

(ii) Anions - $HS^-$, $HCO_3^-$, $HSO_4^-$, $H_2PO_4^-$, $HPO_4^{-2}$, $H_2O$ etc.

(iii) Cations - $NH_4^+$, $H_3O^+$, $PH_4^+$, $CH_3COOH_2^+$ etc.

$[Al(H_2O)_6]^{+3}$, $[Ag(H_2O)_2]^{+1}$, $[Fe(H_2O)_6]^{+3}$ etc.

Ex. 1 $HCl\ (Acid) + H_2O\ (Solvent) \longrightarrow H_3O^+ + Cl^-$

Ex.2 $HS^-\ (Acid) + H_2O\ (Solvent) \longrightarrow H_3O^+ + S^{-2}$

Ex.3 $NH_4^+\ (Acid) + H_2O\ (Solvent) \longrightarrow NH_3 + H_3O^+$

Ex.4 $[Al(H_2O)_6]^{+3}\ (Acid) + H_2O\ (Solvent) \longrightarrow [Al(H_2O)_5OH]^{+2} + H_3O^+$

**(b) Base :-** Those substances which have tendency to accept the proton by any method in any solvent are called the bases.

(1) $HS^-$, $HCO_3^-$, $HSO_4^-$, $H_2PO_4^-$, $HPO_4^{-2}$, $O^{-2}$, $SO_4^{-2}$, $CO_3^{-2}$, $Cl^-$, $Br^-$, $I^-$, $CN^-$ etc.

(2) $NH_3$, $RNH_2$, $R_2NH$, $R_3N$, $C_6H_5NH_2$, $C_5H_5N$, $H_2N\text{-}NH_2$ etc.

**Example** (i) $HS^-\ (Base) + H_2O\ (Solvent) \longrightarrow H_2S + OH^-$

(ii) $NH_3\ (Base) + H_2O\ (Solvent) \longrightarrow NH_4^+ + OH^-$

(iii) $CO_3^{-2}(Base) + H_2O\ (Solvent) \longrightarrow HCO_3^- + OH^-$

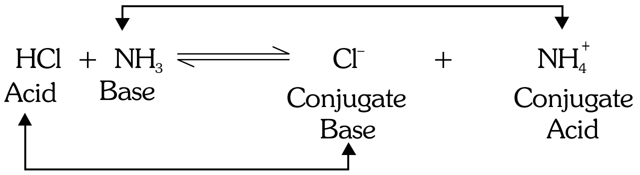
**(c) Nature of water :-** According to this concept nature of water is amphoteric or amphiprotic i.e. water can act as both acid and base.

(i) $HCl\ (Acid) + H_2O\ (Base) \longrightarrow Cl^- + H_3O^+$

(ii) $NH_3\ (Base) + H_2O\ (Acid) \longrightarrow NH_4^+ + OH^-$

**(d) Neutralisation Reaction :-** According to this concept those reaction in which acid and base react together and convert into their respective conjugate acid and base are called neutralisation reactions.

**Example**



**(e) Strength of acids and bases :-** This concept explain the strength of acid and base depending upon the basis of relative tendency to accept or donate the proton.

(i) $HClO_4$
(ii) HI
(iii) HBr
(iv) $H_2SO_4$
(v) HCl
(vi) $HNO_3$
(vii) $H_3O^+$
(viii) $HSO_4^-$
(ix) $H_3PO_4$
(x) HF
(xi) $CH_3COOH$
(xii) $H_2CO_3$
(xiii) $H_2S$
(xiv) $NH_4^+$
(xv) HCN
(xvi) $C_6H_5OH$
(xvii) H – OH
(xviii) $C_2H_5–OH$
(xix) $C_6H_5NH_2$
(xx) $NH_3$
(xxi) $R–NH_2$
(xxii) $CH_4$
(xxiii) $H_2$

**Example**

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| (i) | HCl | + | $H_2O$ | $\rightleftharpoons$ | $Cl^-$ | + | $H_3O^+$ |
| | Strong acid | | Strong base | | Weak base | | Weak acid |
| (ii) | $CH_3COOH$ | + | $H_2O$ | $\rightleftharpoons$ | $CH_3COO^-$ | + | $H_3O^+$ |
| | Weak acid | | Weak base | | Strong base | | Strong acid |
| (iii) | $NH_3$ | + | $H_2O$ | $\rightleftharpoons$ | $NH_4^+$ | + | $OH^-$ |
| | Weak base | | Weak acid | | Strong acid | | Strong base |
| (iv) | HCl | + | $CH_3COOH$ | $\rightleftharpoons$ | $Cl^-$ | + | $CH_3COOH_2^+$ |
| | Weak acid | | Weak base | | Strong base | | Strong acid |
| (v) | HCl | + | $NH_3$ | $\rightleftharpoons$ | $Cl^-$ | + | $NH_4^+$ |
| | Strong acid | | Strong base | | Weak base | | Weak acid |
| (vi) | $CH_3COOH$ | + | $NH_3$ | $\rightleftharpoons$ | $CH_3COO^-$ | + | $NH_4^+$ |
| | Strong acid | | Strong base | | Weak base | | Weak acid |

## GOLDEN KEY POINTS

**Types of Solvent**

**(a) Protogenic or acidic solvent :-** They have a tendency to generate or donate the protons.

**Example** $CH_3COOH$, HCl, $HNO_3$, $H_2SO_4$ etc.

**(b) Protophilic or basic solvent :-** They have a tendency to accept the protons.

**Example** $NH_3$, $CH_3OH$, $RNH_2$ etc.

**(c) Amphiprotic or amphoteric solvents :-** They have the tendency to accept or donate the protons or they are either protophilic or protogenic.

**Example** $H_2O$, $HS^-$, $HSO_4^-$, $HCO_3^-$ etc.

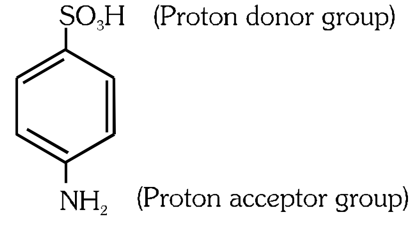
**(d) Aprotic solvents :-** They neither donate nor accept the protons.

**Example** $C_6H_6$, $CCl_4$, $CHCl_3$, $BrF_3$, $NO_2$, $COCl_2$ etc.

**(e) Ampholyte solvents-**Those solvents in which the same molecule acts as proton donor and proton acceptor.

**Example** (i) $\underset{\text{(Glycine)}}{H_2N - CH_2 - COOH} \longrightarrow \underset{\text{(Zwitter ion)}}{H_3^+N - CH_2 - COO^-}$

(ii)



(Sulphanilic acid)

**(f) Advantage :-**

(i) It explain the behaviour of acids and bases in any type of solvent.

(ii) It explain the stability of proton ($H^+$).

**(g) Disadvantage :-**

(i) It does not explain the acid and base practically.

(ii) It does not explain acidic behaviour of aprotic acid. ($SO_2$, $SO_3$, $CO_2$, $AlCl_3$, $SiCl_4$ etc.)

## Illustrations

**Illustration 31** In the process : $NH_3 + NH_3 \rightleftharpoons NH_2^- + NH_4^+$, The nature of ammonia is:-

(1) Acidic (2) Basic (3) Amphoteric (4) None

**Solution.** **Ans.(3)**

**Illustration 32** Which of the following behave both as Bronsted acid as well as Bronsted bases ?

$H_2O$, $HCO_3^-$, $H_2SO_4$, $H_3PO_4$, $HS^-$, $NH_3$

**Solution.** $H_2O$, $HCO_3^-$, $HS^-$, $NH_3$

### (4) Lewis Concept (1939)

**(a) Lewis Acid :-** According to this concept those species which have self tendency to accept the lone pair of electrons are called acids. i.e. Lewis acid is an electron pair acceptor (electrophilic).

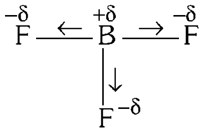
$H^+ + :\ddot{\underset{\cdot\cdot}{O}}{}_{\times}H^- \longrightarrow H:\ddot{\underset{\cdot\cdot}{O}}{}_{\times}H$

Acid Base

**Classification of Lewis Acids :-**

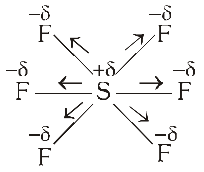
(i) Compounds whose central atom have an incomplete octet (electron deficient)

**Example** $BF_3$, $BBr_3$, $BCl_3$, $BI_3$, $B(CH_3)_3$, $B(OH)_3$, $AlCl_3$, $GeCl_3$ etc.

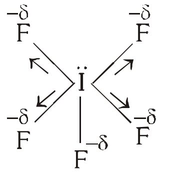


(ii) Compounds whose central atom have vacant d-orbitals and can accept one or more pair of lone electrons.

**Example** $SF_4$, $SF_6$, $SnCl_2$, $SnCl_4$, $PX_3$, $PX_5$, $GeX_4$, $TeX_4$, $IF_5$, $IF_7$, etc.



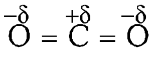
$SF_6$



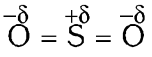
$IF_5$

(iii) Molecules with a multiple bond between atoms of different electronegativities.

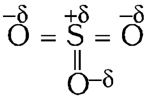
**Example** $CO_2$, $SO_2$, $SO_3$ etc.



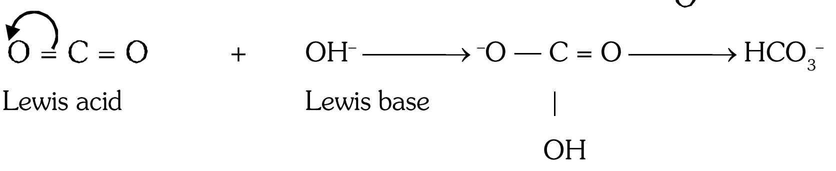
$\overset{-\delta}{O}=\overset{+\delta}{C}=\overset{-\delta}{O}$



$\overset{-\delta}{O}=\overset{+\delta}{S}=\overset{-\delta}{O}$



$\overset{-\delta}{O}=\underset{\overset{\|}{O^{-\delta}}}{\overset{+\delta}{S}}=\overset{-\delta}{O}$



$O=C=O$ + $OH^- \longrightarrow {}^-O-\underset{\underset{OH}{|}}{C}=O \longrightarrow HCO_3^-$

Lewis acid Lewis base

$SO_3$ + $H_2O \longrightarrow H_2SO_4$

Lewis acid Lewis base

(iv) **Cations :-**

**Example** $Ag^+$, $Na^+$, $Li^+$, $Al^{+3}$, $Be^{+2}$, $Mg^{+2}$, $I^+$, $Cl^+$, $H^+$ etc.

**False cations :- Example** $NH_4^+$, $H_3O^+$, $CH_3COOH_2^+$, $PH_4^+$ etc.

**Example** $Ag^+ + 2:NH_3 \rightleftharpoons [Ag(:NH_3)_2]^+$

$Cu^{+2} + 2(:H_2O)_2 \longrightarrow [Cu(:H_2O)_4]^{+2}$

(v) Elements which have six electrons in their outermost shell or valence shell.

O–Family (O, S, Se, Te) Po – Radioactive element

$SO_3^{-2} + \ddot{\underset{\cdot\cdot}{O}}: \longrightarrow [SO_3 \longrightarrow O]^{-2}$

Base Acid

**(b) Lewis Base :-** Those species which have self tendency to donate the lone pair of electrons are called bases. i.e. a base is an electron pair (lone pair) donor (nucleophile).

**Classification of Lewis Bases**

(i) Those species whose central atom have lone pair of electrons and have self tendency to donate them –

**Example** $\ddot{N}H_3$, $R-\ddot{N}H_2$, $R_2-\ddot{N}H$, $C_6H_5-\ddot{N}H_2$, $C_5H_5-\ddot{N}$, $H_2\ddot{N}-\ddot{N}H_2$,

$H-\underset{\cdot\cdot}{\ddot{O}}-H$, $R-\underset{\cdot\cdot}{\ddot{O}}-H$, $R-\underset{\cdot\cdot}{\ddot{O}}-R$, $R-\underset{\cdot\cdot}{\ddot{S}}-R$, $R-\underset{\cdot\cdot}{\ddot{S}}-H$, etc.

Explain :- $\underset{+I}{R}\xrightarrow{\quad}\overset{-\delta}{O}\xleftarrow{\quad}\underset{+I}{R}$ (Electron density on O, increases due to +I effect)

(ii) **Anions** –

**Example** $O^{-2}$, $SO_4^{-2}$, $CO_3^{-2}$, $Cl^-$, $Br^-$, $F^-$, $I^-$, $N^{-3}$, $P^{-3}$, $NO_3^-$, $S^{-2}$, $CN^-$, etc.

**(c) Nature of water :-** According to this concept nature of water is basic i.e. water is a base.

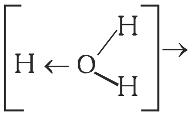
**(d) Neutralisation Reaction :-** According to this concept those reactions in which acid and base react together to form a co-ordinate bond between them are called neutralisation reactions.

**Example**

$$\underset{\text{Lewis acid}}{F_3B} + \underset{\text{Lewis base}}{\ddot{N}H_3} \rightarrow [F_3B\leftarrow NH_3]$$

$$H^+ + \ddot{N}H_3 \rightarrow [H\leftarrow NH_3]^+ \rightarrow NH_4^+$$

$$H^+ + H_2\ddot{O} \rightarrow$$



$$\rightarrow H_3O^+$$

$$Ag^+ + 2\ddot{N}H_3 \rightarrow [H_3\ddot{N}\rightarrow Ag\leftarrow \ddot{N}H_3]^+ \rightarrow [Ag(NH_3)_2]^+$$

**(e) Strength of acids and bases :-** This concept explain the strength of acids and bases depending upon the basis of self tendency. To accept or donate the lone pair of electrons ($e^-$).

**Example** $SO_2$ < $SO_3$ (Strong acid)

**(f) Advantage :-**

(a) This concept explains the acid and base without any solvent.

(b) This concept explains those acids and bases which are not explained by other concepts.

**(g) Disadvantage :-**

(a) This concept fails to explain the acid and base practically.

(b) According to this concept forming of co-ordinate bond in neutralisation reaction of acid and base is a slow process, whereas neutralisation reaction of acids and bases is actually fast process.

(c) This concept fails to explain the real acids and bases

**Example**HCl, $HNO_3$, $H_2SO_4$, NaOH, KOH, etc.

- All the lewis bases are bronsted bases but all the lewis acids are not bronsted acids.
- All Arrhenius acids are Bronsted acids but it is not so for bases.

### (6) POLYACIDIC BASE AND POLYBASIC ACIDS

**(a) Polyacidic Base**

$Al(OH)_3 \rightleftharpoons Al^{+3} + 3OH^- \longrightarrow K_b$

$Al(OH)_3 \rightleftharpoons [Al(OH)_2]^{+1} + OH^- \longrightarrow K_{b_1}$

$[Al(OH)_2]^{+1} \rightleftharpoons [Al(OH)]^{+2} + OH^- \longrightarrow K_{b_2}$

$[Al(OH)]^{+2} \rightleftharpoons Al^{+3} + OH^- \longrightarrow K_{b_3}$

$K_b = K_{b_1} \times K_{b_2} \times K_{b_3}$

taking –log on both sides

$pK_b = pK_{b_1} + pK_{b_2} + pK_{b_3}$

i.e. $\boxed{K_{b_1} > K_{b_2} > K_{b_3}} \Rightarrow \boxed{pK_{b_1} < pK_{b_2} < pK_{b_3}}$

**(b) Polybasic Acid**

$H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{-3} \longrightarrow K_a$

$H_3PO_4 \rightleftharpoons H_2PO_4^{-1} + H^+ \longrightarrow K_{a_1}$

$H_2PO_4^{-1} \rightleftharpoons HPO_4^{-2} + H^+ \longrightarrow K_{a_2}$

$HPO_4^{-2} \rightleftharpoons PO_4^{-3} + H^+ \longrightarrow K_{a_3}$

$K_a = K_{a_1} \times K_{a_2} \times K_{a_3}$

taking –log on both sides

$pK_a = pK_{a_1} + pK_{a_2} + pK_{a_3}$

$\boxed{K_{a_1} > K_{a_2} > K_{a_3}} \Rightarrow \boxed{pK_{a_1} < pK_{a_2} < pK_{a_3}}$

## Illustrations

**Illustration 33** In the dissociation, $H_2A \overset{K_1}{\rightleftharpoons} H^+ + HA^-$; $HA^- \overset{K_2}{\rightleftharpoons} H^+ + A^{-2}$

(1) $K_1$ is equal to $K_2$
(2) $K_1$ is smaller than $K_2$
(3) $K_1$ is greater than $K_2$
(4) $K_1$ is negligible

**Solution.** **Ans. (3)**

**Illustration 34** Calculate the pH of $2.0 \times 10^{-4}$ M $H_3X$ Solution assuming first dissociation to be 100%, second to be 50% and third to be negligible :

**Solution.** First dissociation : $H_3X \rightleftharpoons H_2X^- + H^+$

$\therefore$ $H^+$ ion concentration $= 2 \times 10^{-4}$ M

Second dissociation : $H_2X^- \rightleftharpoons HX^- + H^+$

$H^+$ ion concentration $= C\alpha = 2 \times 10^{-4} \times 0.5$

$\because$ Third dissociation is negligible

$\therefore$ $[H^+] = 2 \times 10^{-4} + 10^{-4} = 3 \times 10^{-4}$

pH = 4 – log3 = 3.52

## BEGINNER'S BOX–5

**1.** Which of the following is a Bronsted acid ?

(i) HCN (ii) $H_2PO_4^-$ (iii) $NH_4^+$ (iv) HCl

(1) (i), (iii) (2) (i), (ii), (iii), (iv) (3) (ii), (iii) (4) (i), (iii),(iv)

**2.** Which of the following is a Bronsted base ?

(i) $NH_3$ (ii) $CH_3NH_2$ (iii) $HCO_3^-$ (iv) $SO_4^{2-}$

(1) (i), (ii), (iii), (iv) (2) (i), (ii) (3) (i), (ii), (iii) (4) (i), (iii),(iv)

**3.** The conjugate base of hydroxide ion is–

(1) $H_2O$ (2) $H_3O^+$ (3) $O^{2-}$ (4) $O_2$

**4.** The conugate acid of amide ion ($NH_2^-$) is–

(1) $N_2H_4$ (2) $NH_2OH$ (3) $NH_4^+$ (4) $NH_3$

**5.** Which of the following can act both as a Bronsted acid as well as a Bronsted base ?

(1) $H_2SO_4$ (2) $HCO_3^-$ (3) $O^{2-}$ (4) $NH_4^+$

**6.** Which of the following acid-base reactions cannot be explained by the Bronsted theory ?

(1) $CO_2 + CaO \longrightarrow CaCO_3$ (2) $BF_3 + NH_3 \longrightarrow \overset{-}{B}F_3\overset{+}{N}H_3$

(3) $Ni + 4CO \longrightarrow Ni(CO)_4$ (4) All of these

**7.** Which of the following Bronsted acid has the weakest conjugate base ?

(1) $H_2O$ (2) HCN (3) HCOOH (4) HF

# ANSWER KEY

| | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|
| **BEGINNER'S BOX-1** | **Que.** | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 |
| | **Ans.** | **3** | **1** | **3** | **4** | **1** | **1** | **4** | **4** | **2** | **1** |
| | **Que.** | 11 | | | | | | | | | |
| | **Ans.** | **2** | | | | | | | | | |
| **BEGINNER'S BOX-2** | **Que.** | 1 | 2 | 3 | 4 | 5 | 6 | 7 | | | |
| | **Ans.** | **4** | **2** | **1** | **2** | **3** | **2** | **1** | | | |
| **BEGINNER'S BOX-3** | **Que.** | 1 | 2 | 3 | 4 | 5 | 6 | | | | |
| | **Ans.** | **2** | **3** | **2** | **3** | **1** | **2** | | | | |
| **BEGINNER'S BOX-4** | **Que.** | 1 | 2 | 3 | 4 | 5 | 6 | | | | |
| | **Ans.** | **3** | **2** | **3** | **1** | **4** | **4** | | | | |
| **BEGINNER'S BOX-5** | **Que.** | 1 | 2 | 3 | 4 | 5 | 6 | 7 | | | |
| | **Ans.** | **2** | **1** | **3** | **4** | **2** | **4** | **4** | | | |

## EXERCISE-I (Conceptual Questions) Build Up Your Understanding

### INTRODUCTION

**1.** The formula weight of $H_2SO_4$ is 98. The weight of the acid in 400mL of 0.1 M solution is:-
(1) 2.45 g (2) 3.92 g (3) 4.90 g (4) 9.8 g

**2.** Normality of 2M sulphuric acid is:-
(1) 2 N (2) 4N (3) N/2 (4) N/4

**3.** If pH = 3.31, then find out $[H^+]$ (Approxy)
(1) $3.39 \times 10^{-4}$ (2) $5 \times 10^{-4}$
(3) $3.0 \times 10^{-3}$ (4) None

**4.** If $[OH^-] = 5.0 \times 10^{-5}$ M then pH will be :-
(1) 5 - log 5 (2) 9 + log 5
(3) log 5 - 5 (4) log 5 - 9

**5.** Basicity of $H_3PO_3$ and $H_3PO_2$ are respectivily :-
(1) 1 and 2 (2) 2 and 3
(3) 3 and 2 (4) 2 and 1

**6.** Find out pH of solution having $2 \times 10^{-3}$ moles of $OH^-$ ion's in 2 litre solution :-
(1) pH = 3 (2) pH = 3 + log2
(3) pH = 3 – log2 (4) pH = 11

**7.** pH of tomato juice is 4.4. Then concentration of $H_3O^+$ will be :-
(1) $39 \times 10^{-4}$ (2) $3.9 \times 10^{-5}$
(3) $3.9 \times 10^{-4}$ (4) $3.9 \times 10^{5}$

**8.** 8 g NaOH is dissolved in one litre of solution, the molarity of the solution is:-
(1) 0.2 M (2) 0.4 M (3) 0.02 M (4) 0.8 M

**9.** The amount of acetic acid present in 100 mL of 0.1M solution is :-
(1) 0.30 g (2) 3.0 g (3) 0.60 g (4) None

**10.** The number of milli equivalents of acid in 100 mL of 0.5N HCl solution is:-
(1) 50 (2) 100 (3) 25 (4) 200

**11.** If the molar concentration of $PbI_2$ is $1.5 \times 10^{-3}$ mol $L^{-1}$, the concentration of iodide ions in g ion $L^{-1}$ is:-
(1) $3.0 \times 10^{-3}$ (2) $6.0 \times 10^{-3}$
(3) $0.3 \times 10^{-3}$ (4) $0.6 \times 10^{-6}$

### OSTWALD'S DILUTION LAW

**12.** Order of dissociation of 0.1 N $CH_3COOH$ is :- (Dissociation constant = $1 \times 10^{-5}$)
(1) $10^{-5}$ (2) $10^{-4}$ (3) $10^{-3}$ (4) $10^{-2}$

**13.** If $\alpha$ is the degree of dissociation of weak dibasic organic acid and y is the hydrogen ion concentration, what is the initial concentration of acid :-
(1) $\frac{\alpha(y)^{-1}}{2}$ M (2) $y(\alpha)^{-1}$M
(3) $\frac{y(\alpha)^{-1}}{2}$ M (4) None of them

**14.** The degree of dissociation of acetic acid is given by the expression $\alpha = 0.1 \times C^{-1}$ (where C = concentration of the acid) What is the pH of the solution :-
(1) 1 (2) 2 (3) 3 (4) 4

**15.** Ostwald's dilution law is not applicable for strong electrolytes because:-
(1) Strong electrolytes are completely ionised
(2) Strong electrolytes are volatile
(3) Strong electrolytes are unstable
(4) Strong electrolytes often contain metal ions

**16.** The degree of ionisation of a compound depends upon :
(1) Size of the solute molecules
(2) Nature of the solute molecules
(3) Nature of the container taken
(4) The amount of current passed

**17.** Find out $K_a$ for $10^{-2}$ M HCN acid, having pOH is 10 :-
(1) $K_a = 10^{-4}$ (2) $K_a = 10^{-2}$
(3) $K_a = 10^{-5}$ (4) None of them

**18.** Which of the following will occur if a 1.0 M solution of a weak acid is diluted to 0.01 M at constant temperature:-
(1) Percentage ionisation will increase
(2) $[H^+]$ will decrease to 0.01M
(3) $K_a$ will increase
(4) pH will decrease by 2 units

**19.** The pH of 0.15 M solution of HOCl ($K_a = 9.6 \times 10^{-6}$) is:-
(1) 4.42 (2) 2.92 (3) 3.42 (4) None

**20.** The extent of ionisation increases (weak electrolytes)
(1) With the increase in concentration of solute
(2) On decreasing the temp. of solution
(3) On addition of excess of water to the solution
(4) On stirring the solution vigorously

**21.** If Ka of HCN = $4 \times 10^{-10}$, then the pH of $2.5 \times 10^{-1}$ molar HCN (aq) is:-
(1) 4.2 (2) 4.7 (3) 0.47 (4) 5.0

**22.** The molarity of nitrous acid at which its pH becomes 2.($K_a = 4.5 \times 10^{-4}$) :-
(1) 0.3333 (2) 0.4444
(3) 0.6666 (4) 0.2222

**23.** Correct statement for HCN weak acid at 25°C temperature.:-
(1) $\alpha = \frac{K_a}{[H^+]}$ (2) $\alpha = \frac{K_a \times [OH^-]}{K_w}$
(3) (1) & (2) both (4) $K_b = C\alpha^2$

## EXPLANATION OF WATER

**24.** Ionic product of water will increase, if :-
(1) Dissociation the pressure
(2) Add $H^+$
(3) Add $OH^-$
(4) Increase the temperature

**25.** For water at 25° C, $2 \times 10^{-7}$ moles per litre is the correct answer for which one of the followoing
(1) $[H^+] + [OH^-]$ (2) $[H^+]^2$
(3) $[OH^-]^2$ (4) $[H^+] - [OH^-]$

**26.** At 25°C, the dissociation constant for pure water is given by :-
(1) $(55.4 \times 10^{14})^{-1}$ (2) $1 \times 10^{-14}$
(3) $\frac{1 \times 10^{-14}}{18}$ (4) None of these

**27.** Ionic product of water is equal to :-
(1) Dissociation constant of water $\times [H_2O]$
(2) Dissociation constant of water $\times [H^+]$
(3) Product of $[H_2O]$ and $[H^+]$
(4) Product of $[OH^-]^2$ and $[H^+]$

**28.** Addition of $H^+$ and $OH^-$ ion's concentration at 90°C
(1) $10^{-14}$ (2) $10^{-12}$
(3) $2 \times 10^{-6}$ (4) $2 \times 10^{-7}$

**29.** At 90°C, pure water has $[H_3O^+] = 10^{-6.7}$ mol $L^{-1}$ what is the value of $K_W$ at 90°C:-
(1) $10^{-6}$ (2) $10^{-12}$
(3) $10^{-67}$ (4) $10^{-13.4}$

**30.** At 373 K, temp. the pH of pure $H_2O$ can be:-
(1) < 7 (2) > 7 (3) = 7 (4) = 0

**31.** The common ion effect is shown by which of the following sets of solutions :-
(1) $BaCl_2 + Ba(NO_3)_2$ (2) NaCl + HCl
(3) $NH_4OH + NH_4Cl$ (4) None

**32.** Basic strength of $NH_4OH$ in presence of $NH_4Cl$
(1) Increases
(2) Remains unchanged
(3) Decreases
(4) Some times increases or sometimes decreases

**33.** Which of the following is a true statement :
(1) The ionisation constant and ionic product of water are same.
(2) Water is a strong electrolyte.
(3) The value of ionic product of water is less than that of its ionisation constant.
(4) At 298K, the number of $H^+$ ions in a litre of water is $6.023 \times 10^{16}$.

**34.** If it is known that $H_2S$ is a weak acid and it is ionised into $2H^+$ and $S^{-2}$. Then in this solution HCl is added so, pH becomes less, then what will happen :-
(1) Decrease in $S^{-2}$ ion concentration
(2) Concentration of $S^{-2}$ is not affected
(3) Increase in $S^{-2}$ ion concentration
(4) It is not possible, to add HCl in solution

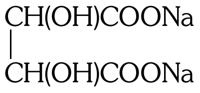
## SALTS, TYPES OF SALT & CONJUGATE THEORY

**35.** Which of the following is not an acidic salt :-
(1) $NaHSO_4$ (2) HCOONa
(3) $NaH_2PO_3$ (4) None of them

**36.** Which is a basic salt :–
(1) PbS (2) $PbCO_3$
(3) $PbSO_4$ (4) $2PbCO_3\ Pb(OH)_2$

**37.** The process of neutralisation invariably results in the production of :-
(1) $H^+$ ions
(2) $OH^-$ ions
(3) Both $H^+$ and $OH^-$ ions
(4) Molecules of water

**38.** Which of the following is an acid salt :-
(1) $Na_2S$ (2) $Na_2SO_3$
(3) $NaHSO_3$ (4) $Na_2SO_4$

**39.** The mixed salt among the following is :-
(1) 
$\begin{array}{l} CH(OH)COONa \\ | \\ CH(OH)COONa \end{array}$
(2) $NaKSO_4$
(3) $CaCl_2$ (4) All

## HYDROLYSIS OF SALTS

**40.** At $90^0$ C, the pH of 0.1M NaCl aqueous solution is :-
(1) < 7 (2) > 7 (3) 7 (4) 0.1

**41.** What will be the pH of 1.0 M ammonium formate solution, If $K_a = 1 \times 10^{-4}$ acid $K_b = 1 \times 10^{-5}$:-
(1) 6.5 (2) 7.5 (3) 8.0 (4) 9.0

**42.** Which salt will not undergo hydrolysis :-
(1) KCl (2) $Na_2SO_4$ (3) NaCl (4) All

**43.** Maximum efficiency of cationic hydrolysis will be shown by :-
(1) $A\ell^{+3}$ (2) $Ga^{+3}$ (3) $T\ell^{+1}$ (4) $T\ell^{+3}$

**44.** $HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$ is related:-
(1) $h = \sqrt{K_h}$ (2) $h = \sqrt{\frac{K_h}{C}}$
(3) $h = \sqrt{\frac{K_h}{V}}$ (4) $K_h = \sqrt{hc}$

**45.** The pH of aqueous solution of sodium acetate is
(1) 7 (2) Very low
(3) > 7 (4) < 7

**46.** If $pK_b$ for $CN^-$ at 25°C is 4.7. The pH of 0.5M aqueous NaCN solution is :-
(1) 12 (2) 10 (3) 11.5 (4) 11

**47.** The highest pH value is of :-
(1) 0.1 M NaCl
(2) 0.1 M $NH_4Cl$
(3) 0.1 M $CH_3COONa$
(4) 0.1 M $CH_3COONH_4$

**48.** pH of $K_2S$ solution is:-
(1) 7 (2) Less than 7
(3) More than 7 (4) 0

**49.** For anionic hydrolysis, pH is given by:-

(1) $pH = \frac{1}{2} pK_W - \frac{1}{2} pK_b - \frac{1}{2} \log c$

(2) $pH = \frac{1}{2} pK_W + \frac{1}{2} pK_a - \frac{1}{2} pK_b$

(3) $pH = \frac{1}{2} pK_W + \frac{1}{2} pK_a + \frac{1}{2} \log c$

(4) None of above

**50.** A weak acid react with strong base, ionisation constant of weak acid is $10^{-4}$. Find out equilibrium constant for this reaction :-
(1) $10^{-10}$ (2) $10^{10}$ (3) $10^{-9}$ (4) $10^{9}$

**51.** Hydroxyl ion concentration $[OH^-]$ in the case of sodium acetate can be expressed as (where $K_a$ is dissociation constant of $CH_3COOH$ and C is the concentration of sodium acetate):-
(1) $[OH^-] = (CK_w . K_a)^{1/2}$ (2) $[OH^-] = C.K_w\sqrt{K_a}$
(3) $[OH^-] = \left(\frac{C.K_w}{K_a}\right)^{1/2}$ (4) $[OH^-] = C. K_a. K_w.$

**52.** Consider :-
(a) $FeCl_3$ in water - Basic
(b) $NH_4Cl$ in water - Acidic
(c) Ammonium acetate in water - Acidic
(d) $Na_2CO_3$ in water - Basic
Which is/are not correctly matched:-
(1) b and d (2) b only
(3) a and c (4) d only

**53.** Which of the following salts undergoes hydrolysis in water:-
(1) $Na_3PO_4$ (2) $CH_3COONa$
(3) $NaNO_3$ (4) Both of (1) and (2)

**54.** A salt 'X' is dissolved in water of pH = 7. The resulting solution becomes alkaline in nature. The salt is made up of:-
(1) A strong acid and strong base
(2) A strong acid and weak base
(3) A weak acid and weak base
(4) A weak acid and strong base

**55.** $K_a$ for cyano acetic acid is $3.5 \times 10^{-3}$. Then the degree of hydrolysis of 0.05 M. sodium cyano acetate solution will have the following value :-
(1) $4.559 \times 10^{-6}$ (2) $5.559 \times 10^{-6}$
(3) $6.559 \times 10^{-6}$ (4) $7.559 \times 10^{-6}$

**56.** Degree of Hydrolysis of $\frac{N}{100}$ solution of KCN is
(Given Ka = $1.4 \times 10^{-9}$)
(1) $2.7 \times 10^{-3}$ (2) $2.7 \times 10^{-2}$
(3) $2.7 \times 10^{-4}$ (4) $2.7 \times 10^{-5}$

## SOLUBILITY & SOLUBILITY PRODUCT($K_{sp}$)

**57.** The solubility product of sparingly soluble univalent salt is defined as the product of ionic concentration in a:-
(1) 1 M solution
(2) Concentration solution
(3) Very dilute solution
(4) Saturated solution

**58.** In solubility of salts $M_2X$, $QY_2$ and $PZ_2$ equal, then the relation between their $K_{sp}$ will be :-
(1) $K_{sp}(M_2X) > K_{sp}(QY_2) > K_{sp}(PZ_2)$
(2) $K_{sp}(M_2X) = K_{sp}(QY_2) < K_{sp}(PZ_2)$
(3) $K_{sp}(M_2X) > K_{sp}(QY_2) = K_{sp}(PZ_2)$
(4) $K_{sp}(M_2X) = K_{sp}(QY_2) = K_{sp}(PZ_2)$

**59.** The expression of solubility product of mercurous iodide is :-
(1) $[2\ Hg^+]^2 \times 2\ [I^-]^2$
(2) $[Hg^{++}]^2 \times [\ 2I^-]^2$
(3) $[Hg_2^{2+}] \times [I^-]^2$
(4) $[Hg^{2+}]^2 \times [I^-]^2$

**60.** At 25°C, the $K_{sp}$ value of AgCl is $1.8 \times 10^{-10}$. If $10^{-5}$ moles of $Ag^+$ are added to solution then $K_{sp}$ will be :-
(1) $1.8 \times 10^{-15}$
(2) $1.8 \times 10^{-10}$
(3) $1.8 \times 10^{-5}$
(4) $18 \times 10^{+10}$

**61.** At 25°C, required volume of water, to dissolve 1g $BaSO_4$ ($K_{sp} = 1.1 \times 10^{-10}$) will be (Molecular weight of $BaSO_4$ = 233) :-
(1) 820 L.
(2) 1 L.
(3) 205 L.
(4) 430 L.

**62.** Concentration of $Ag^+$ ions in saturated solution of $Ag_2CrO_4$ at 20°C is $1.5 \times 10^{-4}$ mol $L^{-1}$. At 20°C, the solubility product of $Ag_2CrO_4$ is :-
(1) $3.3750 \times 10^{-12}$
(2) $1.6875 \times 10^{-10}$
(3) $1.68 \times 10^{-12}$
(4) $1.6875 \times 10^{-11}$

**63.** How many grams of $CaC_2O_4$ will dissolve in distilled water to make one litre saturated solution? solubility product of $CaC_2O_4$ is $2.5 \times 10^{-9}$ mol$^2$ L$^{-2}$ and its molecular weight is 128.
(1) 0.0064 g
(2) 0.0128 g
(3) 0.0032 g
(4) 0.0640 g

**64.** If the concentration of $CrO_4^{2-}$ ion in a saturated solution of silver chromate will be $2 \times 10^{-4}$ M, solubility product of silver chromate will be -
(1) $4 \times 10^{-8}$
(2) $8 \times 10^{-12}$
(3) $32 \times 10^{-12}$
(4) $6 \times 10^{-12}$

**65.** If the solubility of AgCl (formula mass=143) in water at 25°C is $1.43 \times 10^{-4}$ g/100 mL of solution then the value of $K_{sp}$ will be :–
(1) $1 \times 10^{-5}$
(2) $2 \times 10^{-5}$
(3) $1 \times 10^{-10}$
(4) $2 \times 10^{-10}$

**66.** If the salts $M_2X$, $QY_2$ and $PZ_3$ have the same solubilities, their $K_{sp}$ values are related as - (S < 1)
(1) $K_{sp}(M_2X) = K_{sp}(QY_2) < K_{sp}(PZ_3)$
(2) $K_{sp}(M_2X) > K_{sp}(QY_2) = K_{sp}(PZ_3)$
(3) $K_{sp}(M_2X) = K_{sp}(QY_2) > K_{sp}(PZ_3)$
(4) $K_{sp}(M_2X) > K_{sp}(QY_2) > K_{sp}(PZ_3)$

**67.** The solubility product of $As_2S_3$ is given by the expression :-
(1) $K_{sp} = [As^{3+}] \times [S^{-2}]$
(2) $K_{sp} = [As^{3+}]^1\ [S^{-2}]^1$
(3) $K_{sp} = [As^{3+}]^3\ [S^{-2}]^2$
(4) $K_{sp} = [As^{3+}]^2\ [S^{-2}]^3$

**68.** If the solubility of $PbBr_2$ is 'S' g molecules per litre, considering 100% ionisation its solubility product is :-
(1) $2S^3$
(2) $4S^2$
(3) $4S^3$
(4) $2S^4$

**69.** If the solubility of lithium sodium hexeafluoro aluminate $Li_3Na_3(AlF_6)_2$ is 'S' mol $L^{-1}$. Its solubility product is equal to :-
(1) $S^8$
(2) $12\ S^3$
(3) $18S^3$
(4) $2916\ S^8$

**70.** One litre of saturated solution of $CaCO_3$ is evaporated to dryness, when 7.0 g of residue is left. The solubility product for $CaCO_3$ is:-
(1) $4.9 \times 10^{-3}$
(2) $4.9 \times 10^{-5}$
(3) $4.9 \times 10^{-9}$
(4) $4.9 \times 10^{-7}$

## APPLICATION OF SOLUBILITY PRODUCT($K_{sp}$)

**71.** At 30°C, In which of the one litre solution, the solubility of $Ag_2CO_3$ (solubility product = $8 \times 10^{-12}$) will be maximum :-
(1) 0.05 M $Na_2CO_3$
(2) Pure water
(3) 0.05 M $AgNO_3$
(4) 0.05 M $NH_3$

**72.** Solubility of AgBr will be minimum in :-
(1) Pure water
(2) 0.1 M $CaBr_2$
(3) 0.1 M NaBr
(4) 0.1 M $AgNO_3$

**73.** In which of the following, the solution of AgSCN will be unsaturated :-
(1) $[Ag^+]\ [SCN^-] = K_{sp}$
(2) $[Ag^+] \times [SCN^-] < K_{sp}$
(3) $[Ag^+] \times [SCN^-] > K_{sp}$
(4) $[Ag^+]\ [SCN^-]^2 < K_{sp}$

**74.** If 's' and 'S' are respectively solubility and solubility product of a sparingly soluble binary electrolyte then :–
(1) s = S
(2) $s = S^2$
(3) $s = S^{1/2}$
(4) $s = \frac{1}{2}S$

**75.** The solubility product of CuS , $Ag_2S$ and HgS are $10^{-37}$, $10^{-44}$ and $10^{-54}$ respectively. The solubility of these sulphides will be in the order
(1) $HgS > Ag_2S > CuS$
(2) $Ag_2S > HgS > CuS$
(3) $CuS > Ag_2S > HgS$
(4) $Ag_2S > CuS > HgS$

**76.** If the maximum concentration of $PbCl_2$ in water is 0.01 M at 298 K, Its maximum concentration in 0.1 M NaCl will be:-
(1) $4 \times 10^{-3}$ M
(2) $0.4 \times 10^{-4}$ M
(3) $4 \times 10^{-2}$ M
(4) $4 \times 10^{-4}$ M

**77**. $M_2SO_4$ ($M^+$ is a monovalent metal ion) has a $K_{sp}$ of $1.2 \times 10^{-5}$ at 298 K. The maximum concentration of $M^+$ ion that could be attained in a saturated solution of this solid at 298 K is :-

(1) $3.46 \times 10^{-3}$ M (2) $2.89 \times 10^{-2}$ M
(3) $2.8 \times 10^{-3}$ M (4) $7.0 \times 10^{-3}$ M

**78**. Which of the following has maximum solubility ($K_{sp}$ value is given in brackets) :-

(1) HgS ($1.6 \times 10^{-54}$) (2) $PbSO_4$ ($1.3 \times 10^{-8}$)
(3) ZnS ($7.0 \times 10^{-26}$) (4) AgCl ($1.7 \times 10^{-10}$)

**79**. Maximum soluble is :- ($K_{sp}$ is given)

(1) CuS ($8.5 \times 10^{-36}$) (2) CdS ($3.6 \times 10^{-28}$)
(3) ZnS ($1.2 \times 10^{-28}$) (4) MnS ($1.4 \times 10^{-10}$)

**80**. In which of the following, the solubility of AgCl will be maximum :-

(1) 0.1 M $AgNO_3$ (2) Water
(3) 0.1 M NaCl (4) 0.1 M KCl

**81**. The solubility product of three sparingly soluble salts are given below :

| No. | Formula | Solubility product |
|---|---|---|
| 1 | PQ | $4.0 \times 10^{-20}$ |
| 2 | $PQ_2$ | $3.2 \times 10^{-14}$ |
| 3 | $PQ_3$ | $2.7 \times 10^{-35}$ |

The correct order of decreasing molar solublity is:-

(1) 1, 2, 3 (2) 2, 1, 3 (3) 3, 2, 1 (4) 2, 3, 1

**82.** The $K_{sp}$ value for $Gd(OH)_3$ is $2.8 \times 10^{-23}$, the pH of $Gd(OH)_3$ in saturated solution :-

(1) 6.08 (2) 5.08 (3) 8.47 (4) 4.08

**83.** If the solubility product of $AgBrO_3$ and $Ag_2SO_4$ are $5.5 \times 10^{-5}$ and $2 \times 10^{-5}$ respectively, the relationship between the solubilities of these can be correctly represented as:-

(1) $sAgBrO_3 > sAg_2SO_4$
(2) $sAgBrO_3 = sAg_2SO_4$
(3) $sAgBrO_3 < sAg_2SO_4$
(4) Can't predict

**84.** 0.5 M HCl solution has ions- $Hg^{++}$, $Cd^{++}$, $Sr^{++}$, $Fe^{++}$, $Cu^{++}$. To pass the $H_2S$ gas in this solution, which are precipitated out :-

(1) $Cd^{+2}$, $Fe^{+2}$, $Sr^{+2}$ (2) $Cd^{+2}$, $Hg^{+2}$, $Cu^{+2}$
(3) $Hg^{+2}$, $Cu^{+2}$, $Fe^{+2}$ (4) $Cu^{+2}$, $Sr^{+2}$, $Fe^{+2}$

**85.** Solubility product of $Mg(OH)_2$ is $1 \times 10^{-11}$. At what pH, precipitation of $Mg(OH)_2$ will begin from 0.1 M $Mg^{2+}$ solution :-

(1) 9 (2) 5 (3) 3 (4) 7

**86**. A solution, containing 0.01 M $Zn^{+2}$ and 0.01 M $Cu^{2+}$ is saturated by passing $H_2S$ gas. The $S^{-2}$ concentration is $8.1 \times 10^{-21}$ M, Ksp for ZnS and CuS are $3.0 \times 10^{-22}$ and $8.0 \times 10^{-36}$ respectively. Which of the following will occur in the solution:-

(1) ZnS will precipitate
(2) CuS will precipitate
(3) Both ZnS and CuS will precipitate
(4) Both $Zn^{2+}$ and $Cu^{2+}$ will remain in the solution

**87**. Consider (i) $Zn(OH)_2$ (ii) $Cr(OH)_3$ (iii) $Mg(OH)_2$ (iv) $Al(OH)_3$ which hydroxide is precipitated by $NH_4OH$ containing $NH_4Cl$ :-

(1) i, ii (2) ii, iv
(3) Only iv (4) i, ii, iii and iv

**88.** What will happen if the pH of the solution of 0.001 M $Mg(NO_3)_2$ solution is adjusted to pH = 9 ($K_{sp}Mg(OH)_2$ = $8.9 \times 10^{-12}$)

(1) ppt will take place
(2) ppt will not take place
(3) Solution will be saturated
(4) None of these

**89**. The solubility product constant Ksp of $Mg(OH)_2$ is $9.0 \times 10^{-12}$. If a solution is 0.010 M with respect to $Mg^{2+}$ ion, what is the maximum hydroxide ion concentration which could be present without causing the precipitation of $Mg(OH)_2$ :-

(1) $1.5 \times 10^{-7}$ M (2) $3.0 \times 10^{-7}$ M
(3) $1.5 \times 10^{-5}$ M (4) $3.0 \times 10^{-5}$ M

**90**. When HCl gas is passed through a saturated solution of common salt, pure NaCl is precipitated because:-

(1) The impurities dissolve in HCl
(2) HCl is slightly soluble in water
(3) The ionic product $[Na^+] \times [Cl^-]$ exceeds the solubility product of NaCl
(4) The solubility product of NaCl is lowered by $Cl^-$ from aq. HCl

**91**. A solution is a mixture of 0.06 M KCl and 0.06 M KI. $AgNO_3$ solution is being added drop by drop till AgCl starts precipitating ($K_{sp}$ AgCl = $1 \times 10^{-10}$ and $K_{sp}$ AgI = $4 \times 10^{-16}$). The concentration of Iodide ion at this stage will be nearly equal to :-

(1) $4.0 \times 10^{-5}$ M (2) $2.4 \times 10^{-7}$ M
(3) $2.0 \times 10^{-8}$ M (4) $4 \times 10^{-8}$ M

**92**. Why only $As^{3+}$ gets precipitated as $As_2S_3$ and not $Zn^{2+}$ as ZnS when $H_2S$ is passed through an acidic solution containing $As^{3+}$ and $Zn^{2+}$ :-

(1) Enough $As^{3+}$ are present in acidic medium
(2) Zinc salt does not ionise in acidic medium
(3) Solubility product of $As_2S_3$ is less than that of ZnS
(4) Solubility product changes in presence of an acid

**93**. $H_2S$ is passed through a solution of cations in HCl medium to precipitate cation of :-
(1) II-A group of cation analysis
(2) II - B group of cation analysis
(3) IV group of cation analysis
(4) Both II - A and II-B gps.

**94**. To have more sulphide ion concentration, $H_2S$ should be passed through :-
(1) 1 N HCl solution
(2) 0.1 M HCl solution
(3) A neutral solution such as water
(4) An ammonical solution

**95**. The solubility product of hydroxides of $Mg^{+2}$, $Zn^{+2}$, and $Fe^{+3}$ decreases as
$K_{sp}$ $Mg(OH)_2 > K_{sp}$ $Zn(OH)_2 > K_{sp}$ $Fe(OH)_3$ The order of precipitation of hydroxides is:-
(1) $Fe(OH)_3$, $Zn(OH)_2$, $Mg(OH)_2$
(2) $Mg(OH)_2$, $Zn(OH)_2$, $Fe(OH)_3$
(3) $Zn(OH)_2$, $Fe(OH)_3$, $Mg(OH)_2$
(4) $Zn(OH)_2$, $Mg(OH)_2$, $Fe(OH)_3$

## FEW IMPORTANT POINTS

**96.** Two monobasic weak acids have the same concentration of $H^+$ ions. What is the relationship between dissociation constant and dilution:-
(1) $Ka_1V_1 = Ka_2V_2$ (2) $Ka_1V_2 = Ka_2V_1$
(3) $[Ka_1V_1]^{\frac{1}{2}} = Ka_2V_2$ (4) $Ka_1V_1 = [Ka_2V_2]^{\frac{1}{2}}$

**97.** What is the molar concentration of chloride ion in the solution obtained by mixing 300 mL of 3.0M NaCl and 200 mL of 4.0 M solution of $BaCl_2$ :-
(1) 5.0 M (2) 1.8 M
(3) 1.6 M (4) None of these

**98**. The pH of a 0.1 M formic acid 0.1% dissociated is equal to 4. What will be the pH of another weak monobasic acid (same concentration) which is 1% dissociated
(1) 2 (2) 3 (3) 1 (4) 4

## pH

**99.** pH of water is 7. When any substance Y is dissolved in water then pH becomes 13. Substance Y is a salt of :-
(1) Strong acid and strong base
(2) Weak acid and weak base
(3) Strong acid and weak base
(4) Weak acid and strong base

**100.** Minimum pH is shown by aqueous solution of :-
(1) 0.1 M $BaCl_2$ (2) 0.1 M $Ba(NO_3)_2$
(3) 0.1 M $BeCl_2$ (4) 0.1 M $Ba(OH)_2$

**101.** Given :-
(a) 0.005 M $H_2SO_4$ (b) 0.1 M $Na_2SO_4$
(c) $10^{-2}$ M NaOH (d) 0.01 M HCl
Choose the correct code having same pH :-
(1) a, c, d (2) b, d (3) a, d (4) a, c

**102.** What is $H^+$ ion concentration of $5 \times 10^{-3}$ M $H_2CO_3$ solution having a 10% dissociation:-
(1) $10^{-3}$ (2) $10^{-2}$
(3) $10^{-1}$ (4) $5 \times 10^{-2}$

**103.** A metal hydroxide of molecular formula $M(OH)_4$ is 50% ionised. Its 0.0025M solution will have the pH :-
(1) 12 (2) 2 (3) 4 (4) 11.7

**104.** In the following solutions, the conc. of different acids are given, which mixture of the acid has highest pH :-
(1) $\frac{M}{10}H_2SO_4$, $\frac{M}{20}HNO_3$, $\frac{M}{10}HClO_4$
(2) $\frac{M}{20}H_2SO_4$, $\frac{M}{10}HNO_3$, $\frac{M}{20}HClO_4$
(3) $\frac{M}{20}H_2SO_4$, $\frac{M}{10}HNO_3$, $\frac{M}{40}HClO_4$
(4) $\frac{M}{20}H_2SO_4$, $\frac{M}{5}HNO_3$, $\frac{M}{5}HClO_4$

**105.** If 100 mL of pH = 3 and 400 mL of pH = 3 is mixed, what will be the pH of the mixture
(1) 3.2 (2) 3.0 (3) 3.5 (4) 2.8

**106.** $10^{-6}$ M HCl is diluted to 100 times. Its pH is :-
(1) 6.0 (2) 8.0 (3) 6.95 (4) 9.5

**107.** pH of 0.001M acetic acid would be :-
(1) 2 (2) > 3 (3) 7 (4) 14

**108.** At 90°C, the pH of 0.001M KOH solution will be
(1) 3 (2) 11 (3) 5 (4) 9

**109.** The pH of solution is increased from 3 to 6. Its $H^+$ ion concentration will be :-
(1) Reduced to half
(2) Doubled
(3) Reduced by 1000 times
(4) Increased by 1000 times

**110**. A solution has pOH equal to 13 at 298 K. The solution will be:-
(1) Highly acidic (2) Highly basic
(3) Moderately basic (4) Unpredictable

**111.** The pH of the solution containing 10 mL of a 0.1M NaOH and 10 mL of 0.05M $H_2SO_4$ would be
(1) Zero (2) 1 (3) >7 (4) 7

**112.** Which of the following statements are (is) correct
(a) The pH of $1.0 \times 10^{-8}$ M solution of HCl is 8.
(b) The conjugate base of $H_2PO^-_4$ is $HPO_4^{2-}$.
(c) Autoprotolysis constant of water increases with temperature.
(d) When a solution of a weak monoprotic acid is titrated against a strong base, at half neutralization point pH = 1/2 pKa
(1) a (2) a, b (3) a, b, d (4) b, c

**113.** In a solution of pH = 5, more acid is added in order to reduce the pH = 2. The increase in hydrogen ion concentration is:-
(1) 100 times (2) 1000 times
(3) 3 times (4) 5 times

**114.** The hydrogen ion concentration in a given solution is $6 \times 10^{-4}$ M. Its pH will be :-
(1) 6 (2) 3.22 (3) 4 (4) 2.

**115.** The pOH of beer is 10.0. The hydrogen ion concentration will be :–
(a) $10^{-10}$ (b) $\frac{Kw}{10^{-10}}$ (c) $\frac{Kw}{10^{-8}}$ (d) $10^{-4}$
(1) a, d (2) b, c (3) a, b, c (4) None

**116.** An aqueous solution whose pH = 0 is :-
(1) Basic (2) Acidic
(3) Neutral (4) Amphoteric

**117.** The pH of the solution produced when an aqueous solution of strong acid pH 5 is mixed with equal volume of an aqueous solution of strong acid of pH 3 is :-
(1) 3.3 (2) 3.5 (3) 4.5 (4) 4.0

**118.** Following five solution of KOH were prepare as–
First → 0.1 moles in 1 L
Second → 0.2 moles in 2 L
Third → 0.3 moles in 3 L
Fourth → 0.4 moles in 4 L
Fifth → 0.5 moles in 5 L
The pH of resultant solution is :–
(1) 2 (2) 1 (3) 13 (4) 7

**119.** The pH of a 0.02 M ammonia solution which is 5% ionised will be :-
(1) 2 (2) 11 (3) 5 (4) 7

**120.** For $\frac{N}{10}$ $H_2SO_4$, pH value is :-
(1) 1 (2) 0.586 (3) 0.856 (4) None

**121.** An aqueous solution of HCl is $10^{-9}$ M HCl. The pH of the solution should be:-
(1) 9 (2) Between 6 and 7
(3) 7 (4) Unpredictable

**122.** $H_2X$ is a dibasic acid which dissociates completely in water. Which one of the following is the molarity of an aqueous solution of this acid which has a pH of 1 :-
(1) 0.1 (2) 0.05 (3) 0.2 (4) 0.5

**123.** How many moles of HCl must be removed from 1 litre of aqueous HCl solution to change its pH from 2 to 3 :-
(1) 1 (2) 0.02 (3) 0.009 (4) 0.01

**124.** 8 g NaOH and 4.9 g $H_2SO_4$ are present in one litre of the solution. What is its pH
(1) 1 (2) 13 (3) 12 (4) 2

**125.** Calculate pH of a solution whose 100 mL contains 0.2 g NaOH dissolved in it :-
(1) 10.699 (2) 11.699
(3) 12.699 (4) 13.699

**126.** One litre solution contains 1M HOCl [$K_a = 10^{-8}$] and 1 M NaOH. What is the pH of the solution : -
(1) 8 (2) 11 (3) 5 (4) 2

**127.** What is the quantity of NaOH present in 250 cc of the solution, so that it gives a pH = 13 :-
(1) $10^{-13}$ g (2) $10^{-1}$ g (3) 1.0 g (4) 4.0 g

**128.** 0.001 mol of the strong electrolyte $M(OH)_2$ has been dissolved to make a 20 mL of its saturated solution. Its pH will be : - [$K_w = 1 \times 10^{-14}$]
(1) 13 (2) 3.3 (3) 11 (4) 9.8

**129.** Choose the wrong statement :-
(1) For a neutral solution : $[H^+] = [OH^-] = \sqrt{K_w}$
(2) For an acidic solution : $[H^+] > \sqrt{K_w}$ and $[OH^-] < \sqrt{K_w}$
(3) For a basic solution : $[H^+] < \sqrt{K_w}$ and $[OH^-] > \sqrt{K_w}$
(4) For a neutral solution at all temperatures : $[H^+] = [OH^-] = 10^{-7}$ M

**130.** The pH of 0.1 M solution of the following salts increases in order :-
(1) NaCl < $NH_4Cl$ < NaCN < HCl
(2) NaCN < $NH_4Cl$ < NaCl < HCl
(3) HCl < NaCl < NaCN < $NH_4Cl$
(4) HCl < $NH_4Cl$ < NaCl < NaCN

## BUFFER SOLUTIONS and INDICATOR

**131.** In a buffer solution the ratio of concentration of $NH_4Cl$ and $NH_4OH$ is 1 : 1 when it changes in 2 : 1 what will be the value of pH of buffer :-
(1) Increase (2) Decrease
(3) No effect (4) N.O.T.

**132.** To a 50 mL of 0.05M formic acid how much volume of 0.10M sodium formate must be added to get a buffer solution of pH = 4.0 ? ($pK_a$ of the acid is 3.8)
(1) 50 mL (2) 4 mL (3) 39.6 mL (4) 100 mL

**133.** In the volumetric estimation of HCl, if we make use of phenolphthalein as an indicator, which base is unsuitable for the titration :-
(1) NaOH (2) RbOH
(3) KOH (4) $NH_4OH$

**134.** In a mixture of weak acid and its salt, the ratio of concentration of acid to salt is increased ten-fold. The pH of the solution :-
(1) Decreases by one (2) Increases by one-tenth
(3) Increases by one (4) Increases ten-fold

**135**. $pK_b$ for $NH_4OH$ at certain temperature is 4.74. The pH of basic buffer containing equimolar concentration of $NH_4OH$ and $NH_4Cl$ will be:-
(1) 7.74 (2) 4.74 (3) 2.37 (4) 9.26

**136**. What is the suitable indicator for titration of NaOH and oxalic acid:-
(1) Methyl orange (2) Methyl red
(3) Phenolphthalein (4) Starch solution

**137**. Phenolphthalein does not act as an indicator for the titration between :-
(1) KOH and $H_2SO_4$
(2) NaOH and $CH_3COOH$
(3) Oxalic acid and $KMnO_4$
(4) $Ba(OH)_2$ and HCl

**138**. Which can act as buffer :-
(1) $NH_4OH$ + NaOH
(2) HCOOH + $CH_3COONa$
(3) 40 mL 0.1 M NaCN + 20 mL of 0.1 M HCl
(4) None of them

**139**. The buffer solution play an important role in :-
(1) Increasing the pH value
(2) Decreasing the pH value
(3) Keeping the pH constant
(4) Solution will be neutral

**140**. Ka for HCN is 5 x $10^{-10}$ at $25^0C$. For maintaining a constant pH of 9, the volume of 5M KCN solution required to be added to 10mL of 2M HCN solution is-
(1) 4 mL (2) 7.95 mL
(3) 2 mL (4) 9.3 mL

**141.** Buffering action of a mixture of $CH_3COOH$ and $CH_3COONa$ is maximum when the ratio of salt to acid is equal to -
(1) 1.0 (2) 100.0 (3) 10.0 (4) 0.1

**142** The pink colour of phenolphthalein in alkaline medium is due to -
(1) Negative ion (2) Positive ion
(3) $OH^-$ ions (4) Neutral form

**143**. Which indicator works in the pH range 8 – 9.8
(1) Phenolphthalein (2) Methyl orange
(3) Methyl red (4) Litmus

**144.** A basic - buffer will obey the equation pOH - $pK_b$ = 1 only under condition:-
(1) [Conjugate acid] : [base] = 1 : 10
(2) [Conjugate acid] = [base]
(3) [Conjugate acid] : [base] = 10 : 1
(4) N.O.T

**145.** For weak acid strong base titration, the indicator used is :-
(1) Potassium di-chromate
(2) Methyl orange
(3) Litmus
(4) Phenolphthalein

**146.** From the following in which titration methyl orange is a best indicator :-
(1) $CH_3COOH$ + NaOH (2) $H_2C_2O_4$ + NaOH
(3) HCl + NaOH (4) $CH_3COOH + NH_4OH$

**147**. The total number of different kind of buffers obtained during the titration of $H_3PO_4$ with NaOH are :-
(1) 3 (2) 1 (3) 2 (4) 0

**148**. The $H^+$ ion concentration in 0.001 M acetic acid is $1.34 \times 10^{-4}$ g ion $L^{-1}$. The $H^+$ ion concentration of 0.164 g of $CH_3COONa$ is added to a litre of 0.001 M $CH_3COOH$ will be :-
(1) $9 \times 10^{-6}$ (2) $18 \times 10^{-6}$
(3) $4.5 \times 10^{-6}$ (4) $5 \times 10^{-6}$

**149**. A certain acidic buffer solution contains equal concentration of $X^-$ and HX. The $K_b$ for $X^-$ is $10^{-10}$. The pH of the buffer is :-
(1) 4 (2) 7 (3) 10 (4) 14

**150**. When 1.0 mL of dil. HCl acid is added to 100 mL of a buffer solution of pH 4.0. The pH of the solution
(1) Becomes 7 (2) Does not change
(3) Becomes 2 (4) Becomes 10

**151**. The pH of blood is maintained by $CO_2$ and $H_2CO_3$ in the body and chemical constituents of blood. This phenomenon is called :-
(1) Colloidal (2) Buffer action
(3) Acidity (4) Salt balance

**152**. Phenolphthalein is not a good indicator for titrating
(1) NaOH against oxalic acid
(2) NaOH against HCl
(3) NaOH against $H_2SO_4$
(4) Ferrous sulphate against $KMnO_4$

**153**. Which of the following solutions does not act as buffer :–
(1) $H_3PO_4 + NaH_2PO_4$
(2) $NaHCO_3 + H_2CO_3$
(3) $NH_4Cl + HCl$
(4) $CH_3COOH + CH_3COONa$

**154**. 50 mL of 2N acetic acid mixed with 10 mL of 1N sodium acetate solution will have an approximate pH of ($K_a = 10^{-5}$) :–
(1) 4 (2) 5 (3) 6 (4) 7

**155**. On addition of NaOH to $CH_3COOH$ solution, 60% of the acid is neutralised. If $pK_a$ of $CH_3COOH$ is 4.7 then the pH of the resulting solution is :-
(1) More than 4.7 but less than 5.0
(2) Less than 4.7 but more than 4.0
(3) More than 5.0
(4) Remains unchanged

**156**. 500 mL of 0.2 M acetic acid are added to 500 mL of 0.30 M sodium acetate solution. If the dissociation constant of acetic acid is $1.5 \times 10^{-5}$ then pH of the resulting solution is:-
(1) 5.0 (2) 9.0 (3) 3.0 (4) 4.0

**157**. Half of the formic acid solution is neutralised on addition of a KOH solution to it. If $K_a$ (HCOOH) = $2 \times 10^{-4}$ then pH of the solution is : (log 2 = 0.3010)
(1) 3.6990 (2) 10.3010 (3) 3.85 (4) 4.3010

**158**. A solution contains 0.2M $NH_4OH$ and 0.2M $NH_4Cl$. If 1.0 mL of 0.001 M HCl is added to it. What will be the $[OH^-]$ of the resulting solution $[K_b = 2 \times 10^{-5}]$ :-
(1) $2 \times 10^{-5}$ (2) $5 \times 10^{-10}$
(3) $2 \times 10^{-3}$ (4) None of these

**159**. Hunderson equation $pH - pK_a = 1$ will be applicable to an acidic buffer when :-
(1) [Acid] = [Conjugate base]
(2) [Acid] × 10 = [Conjugate base]
(3) [Acid] = [Conjugate base] × 10
(4) None of these

**160**. 0.05 M ammonium hydroxide solution is dissolved in 0.001 M ammonium chloride solution. What will be the $OH^-$ ion concentration of this solution: $K_b(NH_4OH) = 1.8 \times 10^{-5}$
(1) $3.0 \times 10^{-3}$ (2) $9.0 \times 10^{-4}$
(3) $9.0 \times 10^{-3}$ (4) $3.0 \times 10^{-4}$

**161**. When 0.02 moles of NaOH are added to a litre of buffer solution, its pH changes from 5.75 to 5.80. What is its buffer capacity :-
(1) 0.4 (2) 0.05 (3) – 0.05 (4) 2.5

**162**. Calculate the pH of a buffer prepared by mixing 300 cc of 0.3 M $NH_3$ and 500 cc of 0.5 M $NH_4Cl$. $K_b$ for $NH_3 = 1.8 \times 10^{-5}$ :-
(1) 8.1187 (2) 9.8117
(3) 8.8117 (4) None of these

**163**. What amount of sodium propanoate should be added to one litre of an aqueous solution containing 0.02 mole of propanoic acid (Ka = $1.34 \times 10^{-5}$ at 25°C) to obtain a buffer solution of pH 4.75 :-
(1) $4.52 \times 10^{-2}$ M (2) $3.52 \times 10^{-2}$ M
(3) $2.52 \times 10^{-2}$ M (4) $1.52 \times 10^{-2}$ M

**164**. What will be the pH of the solution, if 0.01 moles of HCl is dissolved in a buffer solution containing 0.02 moles of propanoic acid ($K_a = 1.34 \times 10^{-5}$) and 0.0152 moles of salt, at 25°C : [log(0.173) = –0.76]
(1) 3.11 (2) 4.11 (3) 5.11 (4) 6.11

**165**. Calculate the pH of a buffer solution prepared by dissolving 30g of $Na_2CO_3$ in 500 mL of an aqueous solution containing 150 mL of 1M HCl. Ka for $HCO_3^- = 5.63 \times 10^{-11}$ $\left[\log\left(\frac{133}{150}\right) = -0.05\right]$
(1) 8.197 (2) 9.197
(3) 10.197 (4) 11.197

**166**. Calculate the ratio of pH of a solution containing 1 mole of $CH_3COONa$ + 1 mole of HCl per litre and of other solution containing 1 mole $CH_3COONa$ + 1mole of acetic acid per litre :-
(1) 1 : 1 (2) 2 : 1 (3) 1 : 2 (4) 2 : 3

**167.** When 20 mL of $\frac{M}{20}$ NaOH are added to 10 mL of $\frac{M}{10}$ HCl, the resulting solution will:-
(1) Turn blue litmus red
(2) Turn phenolphthalein solution
(3) Turn methyl orange red
(4) Will have no effect on either red or blue litmus

**168**. 10 mL of a solution contains 0.1 M $NH_4Cl$+ 0.01 M $NH_4OH$. Which addition would not change the pH of solution :–
(1) Adding 1 mL water
(2) Adding 5 mL of 0.1 M $NH_4Cl$
(3) Adding 5 mL of 0.1 M $NH_4OH$
(4) Adding 10 mL of 0.1 M $NH_4Cl$

**169.** $\frac{N}{10}$ acetic acid was titrated with $\frac{N}{10}$ NaOH. When 25%, 50% and 75% of titration is over then the pH of the solution will be :- $[K_a = 10^{-5}]$
(1) 5 + log 1/3, 5, 5 + log 3
(2) 5 + log 3, 4, 5 + log 1/3
(3) 5 – log 1/3, 5, 5 – log 3
(4) 5 – log 1/3, 4, 5 + log 1/3

## ACID AND BASE

**170.** The conjugated acid of $O^{-2}$ ion's is :-
(1) $O_2^+$ (2) $H^+$ (3) $H_3O^+$ (4) $OH^-$

**171.** Ionization constant of AOH and BOH base are $K_{b_1}$ and $K_{b_2}$. Their relation is $pK_{b_1} < pK_{b_2}$. Conjugate of following base, does not show maximum pH :
(1) AOH (2) BOH
(3) Both of them (4) NOT

**172.** Select the species which can function as - Lewis base, bronsted acid and bronsted base:-
(a) $H_2O$ (b) $NH_4^+$ (c) $N^{-3}$
Correct code is :-
(1) Only a (2) a, b (3) a, c (4) b, c

**173.** Which ion does not show acid behaviour :-
(1) $\left[Al(H_2O)_6\right]^{+3}$ (2) $\left[Fe(H_2O)_6\right]^{+3}$
(3) $HPO_4^{-2}$ (4) $ClO_3^-$

**174.** An example of Lewis acid is:-
(1) CaO (2) $CH_3NH_2$
(3) $SO_3$ (4) None of these

**175.** In the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ water behaves as :-
(1) Acid (2) Base
(3) Neutral (4) Both acid & Base

**176.** Which acts as Lewis base in the reaction $BCl_3 + :PH_3 \rightarrow Cl_3B \leftarrow PH_3$
(1) $PH_3$ (2) $BCl_3$
(3) Both 1 & 2 (4) None

**177.** Which acts as Lewis acid in the reaction
$SnCl_2 + 2Cl^- \longrightarrow [SnCl_4]^{-2}$
(1) $Cl^-$ (2) $SnCl_2$ (3) $SnCl_4$ (4) None

**178.** The conjugated base of $(CH_3)_2\overset{+}{N}H_2$ is :-
(1) $CH_3NH_2$ (2) $(CH_3)_2N^+$
(3) $(CH_3)_2N$ (4) $(CH_3)_2NH$

**179.** Which equilibrium can be described as Lewis acid base reaction but not Bronsted acid base reaction:-
(1) $H_2O + CH_3COOH \rightleftharpoons H_3O^+ + CH_3COO^-$
(2) $2NH_3 + H_2SO_4 \rightleftharpoons 2NH_4^+ + SO_4^{2-}$
(3) $NH_3 + CH_3COOH \rightleftharpoons NH_4^+ + CH_3COO^-$
(4) $Cu^{+2} + 4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{2+}$

**180.** Conjugate base of hydrazoic acid is :-
(1) $HN_3^-$ (2) $N_3^-$ (3) $N^{3-}$ (4) $N_2^-$

**181.** $NH_3$ gas dissolves in water to give $NH_4OH$, in this reaction, water acts as :-
(1) An acid (2) A base
(3) A salt (4) A conjugate base

**182.** Conjugate acid of $Zn(OH)_2$ is :-
(1) $Zn(OH)^+$ (2) $Zn(OH_3)^-$ (3) $Zn^{2+}$ (4) None

**183.** When ammonia is added to water it decreases the concentration of which of the following ion
(1) $OH^-$ (2) $H_3O^+$ (3) $NH_4^+$ (4) None

**184.** The strongest acid among the following is -
(1) $ClO_3(OH)$ (2) $ClO_2(OH)$
(3) $SO(OH)_2$ (4) $SO_2(OH)_2$

**185.** Which of the following is not a Bronsted acid :-
(1) $CH_3NH_4^+$ (2) $CH_3COO^-$
(3) $H_2O$ (4) $HSO_4^-$

**186.** Which of the following example behave as a Lewis acid $BF_3$ , $SnCl_2$ , $SnCl_4$ :-
(1) Stanus chloride , stanic chloride
(2) $BF_3$ , Stanus chloride
(3) Only $BF_3$
(4) $BF_3$ , stanus chloride, stanic chloride

**187.** In the reaction
$HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$, the conjugate base of $HNO_3$ is :-
(1) $H_2O$ (2) $H_3O^+$
(3) $NO_3^-$ (4) $H_3O^+$ and $NO_3^-$

**188.** The conjugate base of the weak acid in the reaction
$HBr + H_2O \rightleftharpoons H_3O^+ + Br^-$ is
(1) HBr (2) $H_2O$ (3) $Br^-$ (4) $H_3O^+$

**189.** In the reaction, $AlCl_3 + Cl^- \rightarrow [AlCl_4]^-$, $AlCl_3$ acts as :-
(1) Salt (2) Lewis base
(3) Bronsted acid (4) Lewis acid

**190.** $Mg^{2+}$ is ------- than $Al^{3+}$ :-
(1) Strong Lewis acid (2) Strong Lewis base
(3) Weak Lewis acid (4) Weak Lewis base

**191**. The two Bronsted bases in the reaction

$HC_2O_4^- + PO_4^{3-} \rightleftharpoons HPO_4^{2-} + C_2O_4^{2-}$ are
(1) $HC_2O_4^-$ and $PO_4^{2-}$
(2) $HPO_4^{2-}$ and $C_2O_4^{2-}$
(3) $PO_4^{3-}$ and $C_2O_4^{-2}$
(4) $HC_2O_4^-$ and $HPO_4^{2-}$

**192**. The compound HCl behaves as ---- in the reaction,

$HCl + HF \rightleftharpoons H_2^+Cl + F^-$
(1) Strong acid (2) Strong base
(3) Weak acid (4) Weak base

**193.** Which of the following is not a lewis base:-
(1) $NH_3$ (2) $O^{2-}$ (3) $H_2O$ (4) $I^+$

**194.** Which of the following is bronsted Lowry acid:-
(1) $SO_4^{-2}$ (2) $H_3O^+$ (3) $OH^-$ (4) $Cl^-$

**195.** The conjugated base for bicarbonate ion is:-
(1) $CO_3^{2-}$ (2) $HCO_3^-$
(3) $CO_2$ (4) $H_2CO_3$

**196.** Conjugated base of $OH^-$ is :-
(1) $H_2O$ (2) $H_3O^+$ (3) $H^+$ (4) $O^{2-}$

**197.** HCl does not behave as acid in :-
(1) $NH_3$ (2) $C_2H_5OH$ (3) $H_2O$ (4) $C_6H_6$

**198**. Which of the following is a base according to Lowry-Bronsted concept :-
(1) $I^-$ (2) $H_3O^+$ (3) HCl (4) $NH_4^+$

**199**. In which of the following reactions $NH_3$ acts as acid
(1) $NH_3 + HCl \rightarrow NH_4Cl$
(2) $NH_3 + H^+ \rightarrow NH_4^+$
(3) $NH_3 + Na \rightarrow NaNH_2 + \frac{1}{2}H_2$
(4) $NH_3$ cannot act as acid

**200**. According to Bronsted concept, the acids in the following reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ are :-
(1) $NH_3$ and $NH_4^+$ (2) $H_2O$ and $OH^-$
(3) $H_2O$ and $NH_4^+$ (4) $NH_3$ and $OH^-$

**201**. Consider the following reactions :-

(i) $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$

(ii) $CO_2 + H_2O \rightleftharpoons H_2CO_3$

(iii) $NH_3 + H_2O \rightleftharpoons NH_4OH$

(iv) $HCl + H_2O \rightleftharpoons Cl^- + H_3O^+$
Which of the pairs of reactions proves that water is amphoteric in character :-
(1) (i) and (ii) (2) (ii) and (iii)
(3) (iii) and (iv) (4) (i) and (iii)

**202**. $CH_3COO^-$ ion is a :-
(1) Weak conjugate base
(2) Strong conjugate base
(3) Weak conjugate acid
(4) Strong conjugate acid

**203**. Which of the following is strongest conjugate base
(1) $ClO_4^-$ (2) $HCO_3^-$ (3) $F^-$ (4) $HSO_4^-$

**204**. Which of the following species can act as Lewis base :-
(1) $Cu^{2+}$ (2) $AlCl_3$ (3) $NH_3$ (4) $BF_3$

**205**. A compound having the formula $NH_2CH_2COOH$ may behave :-
(1) Only as an acid
(2) Only as a base
(3) Both as an acid and base
(4) Neither acid nor base

**206**. $BF_3$ is acid according to :-
(1) Lewis
(2) Arrhenius
(3) Bronsted and Lowery
(4) Madam Curie

**207**. Which of the following can act both as Bronsted acid and Bronsted base :-
(1) $Na_2CO_3$ (2) $O^{2-}$
(3) $CO_3^{-2}$ (4) $NH_3$

**208**. The strongest conjugate base is :-
(1) $NO_3^-$ (2) $Cl^-$
(3) $SO_4^{2-}$ (4) $CH_3COO^-$

**209**. Aluminium chloride is :-
(1) Bronsted Lowry acid (2) Arrhenius acid
(3) Lewis acid (4) Lewis base

**210**. Water is a :-
(1) Protogenic solvent (2) Protophilic solvent
(3) Amphiprotic solvent (4) Aprotic solvent

**211**. Ammonium ion is :-
(1) A conjugate acid
(2) A conjugate base
(3) Neither an acid nor a basic
(4) Both an acid and a base

**212**. Species which do not act both as Bronsted acid and base is :-
(1) $(HSO_4)^{-1}$ (2) $Na_2CO_3$
(3) $NH_3$ (4) $OH^{-1}$

**213.** Which one of the following is strong Lewis base & Bronsted acid & bronsted base:-
(1) $NH_3$ (2) $PH_3$ (3) $CH_4$ (4) $BH_3$

**214.** Which of the following pair is Lewis acid & Lewis base & product of these is also Lewis base

(1) $BF_3$ , $NH_3$ (2) $SiCl_4$ , $2Cl^-$

(3) $CH_3^{\oplus}$ , $^{\ominus}OC_2H_5$ (4) None of these

**215.** Which of the following is not a correct statement

(1) Arrhenius theory of acids-bases is capable of explaining the acidic or basic nature of the substances in the solvents other than water

(2) Arrhenius theory does not explain acidic nature of $AlCl_3$

(3) The aqueous solution of $Na_2CO_3$ is alkaline although it does not contain $OH^-$ ions

(4) Aqueous solution of $CO_2$ is acidic although it does not contain $H^+$ ions

**216.** For the reaction $NH_4^+ + S^{-2} \rightleftharpoons NH_3 + HS^-$, $NH_3$ and $S^{-2}$ are a group of :-

(1) Acids (2) Bases

(3) Acid-base pair (4) None of these

**217.** According to Lewis concept acid & base pair is–

(1) $HO^-$, $H^+$ (2) $Ag^+$,$Cl^-$

(3) $BF_3$,$NH_3$ (4) None of these

## EXERCISE-I (Conceptual Questions) ANSWER KEY

| Que. | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| **Ans.** | **2** | **2** | **2** | **2** | **4** | **4** | **2** | **1** | **3** | **1** | **1** | **4** | **3** | **1** | **1** |
| Que. | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | 29 | 30 |
| **Ans.** | **2** | **4** | **1** | **2** | **3** | **4** | **4** | **3** | **4** | **1** | **1** | **1** | **3** | **4** | **1** |
| Que. | 31 | 32 | 33 | 34 | 35 | 36 | 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 |
| **Ans.** | **3** | **3** | **4** | **1** | **2** | **4** | **4** | **3** | **2** | **1** | **1** | **4** | **1** | **2** | **3** |
| Que. | 46 | 47 | 48 | 49 | 50 | 51 | 52 | 53 | 54 | 55 | 56 | 57 | 58 | 59 | 60 |
| **Ans.** | **3** | **3** | **3** | **3** | **2** | **3** | **3** | **4** | **4** | **4** | **2** | **4** | **4** | **3** | **2** |
| Que. | 61 | 62 | 63 | 64 | 65 | 66 | 67 | 68 | 69 | 70 | 71 | 72 | 73 | 74 | 75 |
| **Ans.** | **4** | **3** | **1** | **3** | **3** | **3** | **4** | **3** | **4** | **1** | **4** | **2** | **2** | **3** | **4** |
| Que. | 76 | 77 | 78 | 79 | 80 | 81 | 82 | 83 | 84 | 85 | 86 | 87 | 88 | 89 | 90 |
| **Ans.** | **4** | **2** | **2** | **4** | **2** | **4** | **3** | **3** | **2** | **1** | **2** | **2** | **2** | **4** | **3** |
| Que. | 91 | 92 | 93 | 94 | 95 | 96 | 97 | 98 | 99 | 100 | 101 | 102 | 103 | 104 | 105 |
| **Ans.** | **2** | **3** | **4** | **4** | **1** | **2** | **1** | **2** | **4** | **3** | **3** | **1** | **4** | **3** | **2** |
| Que. | 106 | 107 | 108 | 109 | 110 | 111 | 112 | 113 | 114 | 115 | 116 | 117 | 118 | 119 | 120 |
| **Ans.** | **3** | **2** | **4** | **3** | **1** | **4** | **4** | **2** | **2** | **4** | **2** | **1** | **3** | **2** | **1** |
| Que. | 121 | 122 | 123 | 124 | 125 | 126 | 127 | 128 | 129 | 130 | 131 | 132 | 133 | 134 | 135 |
| **Ans.** | **2** | **2** | **3** | **2** | **3** | **2** | **3** | **1** | **4** | **4** | **2** | **3** | **4** | **1** | **4** |
| Que. | 136 | 137 | 138 | 139 | 140 | 141 | 142 | 143 | 144 | 145 | 146 | 147 | 148 | 149 | 150 |
| **Ans.** | **3** | **3** | **3** | **3** | **3** | **1** | **1** | **1** | **3** | **4** | **3** | **1** | **1** | **1** | **2** |
| Que. | 151 | 152 | 153 | 154 | 155 | 156 | 157 | 158 | 159 | 160 | 161 | 162 | 163 | 164 | 165 |
| **Ans.** | **2** | **4** | **3** | **1** | **1** | **1** | **1** | **1** | **2** | **2** | **1** | **3** | **4** | **2** | **3** |
| Que. | 166 | 167 | 168 | 169 | 170 | 171 | 172 | 173 | 174 | 175 | 176 | 177 | 178 | 179 | 180 |
| **Ans.** | **3** | **4** | **1** | **1** | **4** | **2** | **1** | **4** | **3** | **1** | **1** | **2** | **4** | **4** | **2** |
| Que. | 181 | 182 | 183 | 184 | 185 | 186 | 187 | 188 | 189 | 190 | 191 | 192 | 193 | 194 | 195 |
| **Ans.** | **1** | **1** | **2** | **1** | **2** | **4** | **3** | **2** | **4** | **3** | **3** | **4** | **4** | **2** | **1** |
| Que. | 196 | 197 | 198 | 199 | 200 | 201 | 202 | 203 | 204 | 205 | 206 | 207 | 208 | 209 | 210 |
| **Ans.** | **4** | **4** | **1** | **3** | **3** | **3** | **2** | **2** | **3** | **3** | **1** | **4** | **4** | **3** | **3** |
| Que. | 211 | 212 | 213 | 214 | 215 | 216 | 217 | | | | | | | | |
| **Ans.** | **1** | **2** | **1** | **3** | **1** | **2** | **3** | | | | | | | | |

## EXERCISE-II (Previous Year Questions) AIPMT/NEET & AIIMS (2006-2018)

### AIPMT 2006

**1.** Which of the following pairs constitutes a buffer :-

(1) $HNO_2$ & $NaNO_2$ (2) NaOH & NaCl
(3) $HNO_3$ & $NH_4NO_3$ (4) HCl & KCl

**2.** The hydrogen ion concentration of a $10^{-8}$ M HCl aqueous solution at 298 K ($K_W = 10^{-14}$) is :-

(1) $1.0 \times 10^{-6}$ M (2) $1.0525 \times 10^{-7}$ M
(3) $9.525 \times 10^{-8}$ M (4) $1.0 \times 10^{-8}$ M

### AIIMS 2006

**3.** 40 mL of 0·1 M ammonia solution is mixed with 20 mL of 0·1 M HCl. What is the pH of the mixture ($pK_b$ of ammonia solution is 4·74) :–

(1) 4·74 (2) 2·26 (3) 9·26 (4) 5·00

### AIPMT 2007

**4.** Calculate the pOH of a solution at $25^0$ C that contains $1 \times 10^{-10}$M of hydronium ions, i.e., $H_3O^+$

(1) 7.000 (2) 4.000 (3) 9.000 (4) 1.000

**5.** A weak acid *HA* has a $K_a$ of $1.00 \times 10^{-5}$. If 0.100 moles of this acid is dissolved in one litre of water the percentage of acid dissociated at equilibrium is closed to :-

(1) 99.0% (2) 1.00%
(3) 99.9% (4) 0.100%

### AIPMT 2008

**6.** Equimolar solutions of the following were prepared in water separately. Which one of the solutions will record the highest pH?

(1) $BaCl_2$ (2) $MgCl_2$ (3) $CaCl_2$ (4) $SrCl_2$

**7.** Equal volumes of three acid solutions of pH 3, 4 and 5 are mixed in a vessel. What will be the $H^+$ ion concentration in the mixture ?

(1) $3.7 \times 10^{-4}$ M (2) $3.7 \times 10^{-3}$ M
(3) $1.11 \times 10^{-3}$ M (4) $1.11 \times 10^{-4}$ M

### AIPMT 2009

**8.** The ionization constant of ammonium hydroxide is $1.77 \times 10^{-5}$ at 298 K. Hydrolysis constant of ammonium chloride is :-

(1) $5.65 \times 10^{-12}$ (2) $5.65 \times 10^{-10}$
(3) $6.50 \times 10^{-12}$ (4) $5.65 \times 10^{-13}$

**9.** What is the $[OH^-]$ in the final solution prepared by mixing 20.0 mL of 0.050 M HCl with 30.0 mL of 0.10M $Ba(OH)_2$ ?

(1) 0.12 M (2) 0.10 M
(3) 0.40 M (4) 0.0050M

**10.** The dissociation constants for acetic acid and HCN at 25°C are $1.5 \times 10^{-5}$ and $4.5 \times 10^{-10}$, respectively. The equilibrium constant for the equilibrium

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

would be :-

(1) $3.0 \times 10^{4}$ (2) $3.0 \times 10^{5}$
(3) $3.0 \times 10^{-5}$ (4) $3.0 \times 10^{-4}$

### AIPMT 2010

**11.** If pH of a saturated solution of $Ba(OH)_2$ is 12, the value of its $K_{(sp)}$ is :-

(1) $5.00 \times 10^{-7}$ $M^3$ (2) $4.00 \times 10^{-6}$ $M^3$
(3) $4.00 \times 10^{-7}$ $M^3$ (4) $5.00 \times 10^{-6}$ $M^3$

**12.** In a buffer solution containing equal concentration of $B^-$ and HB the $K_b$ for $B^-$ is $10^{-10}$. The pH of buffer solution is :-

(1) 4 (2) 10 (3) 7 (4) 6

### AIIMS 2010

**13.** The pH of blood from 7.26 to 7.42 is maintained by buffer :-

(1) $H_2CO_3/HCO_3^-$
(2) $H_3PO_4/H_2PO_4^-$
(3) $CH_3COOH/CH_3COO^-$
(4) $NH_2–CH_2COOH/NH_2CH_2COO^-$

**14.** $K_{sp}$ of a salt $Ni(OH)_2$ is $2 \times 10^{-15}$ then molar solubility of $Ni(OH)_2$ in 0.01M NaOH is :-

(1) $2 \times 10^{-15}$ M (2) $2^{1/3} \times 10^{-5}$ M
(3) $2 \times 10^{-11}$ M (4) $10^{-7}$ M

### AIPMT Mains 2011

**15.** In qualitative analysis, the metals of Group I can be separated from other ions by precipitating them as chloride salts. A solution initially contains $Ag^+$ and $Pb^{2+}$ at a concentration of 0.10 M. Aqueous HCl is added to this solution until the $Cl^-$ concentration is 0.10 M. What will the concentrations of $Ag^+$ and $Pb^{2+}$ be at equilibrium ?

($K_{sp}$ for AgCl=$1.8 \times 10^{-10}$, $K_{sp}$ for $PbCl_2$=$1.7 \times 10^{-5}$)

(1) $[Ag^+] = 1.8 \times 10^{-11}$ M ; $[Pb^{2+}] = 1.7 \times 10^{-4}$ M;
(2) $[Ag^+] = 1.8 \times 10^{-7}$ M ; $[Pb^{2+}] = 1.7 \times 10^{-6}$ M;
(3) $[Ag^+] = 1.8 \times 10^{-11}$ M ; $[Pb^{2+}] = 8.5 \times 10^{-5}$ M;
(4) $[Ag^+] = 1.8 \times 10^{-9}$ M ; $[Pb^{2+}] = 1.7 \times 10^{-3}$ M;

**16.** A buffer solution is prepared in which the concentration of $NH_3$ is 0.30 M and the concentration of $NH_4^+$ is 0.20 M. If the equilibrium constant, $K_b$ for $NH_3$ equals $1.8 \times 10^{-5}$, what is the pH of this solution ? (log 2.7 = 0.43)

(1) 9.08 (2) 9.43
(3) 11.72 (4) 8.73

### AIIMS 2011

**17.** $K_{sp}$ of a $CaSO_4.5H_2O$ is $9 \times 10^{-6}$, Find the volume of $CaSO_4$ for 1gm (Mw = 136)

(1) 2.45 litre (2) 5.1 litre
(3) 4.52 litre (4) 3.2 litre

**18.** 25ml, 0.2M $Ca(OH)_2$ is neutralised by 10ml of 1M HCl. Then pH of resulting solution is–

(1) 1.37 (2) 9
(3) 12 (4) 7

### AIPMT Mains 2012

**19.** Buffer solutions have constant acidity and alkalinity because:

(1) they have large excess of $H^+$ or $OH^-$ ions
(2) they have fixed value of pH
(3) these give unionised acid or base on reaction with added acid or alkali
(4) acids and alkalies in these solutions are shielded from attack by other ions

**20.** Equimolar solutions of the following substances were prepared separately. Which one of the these will record the highest pH value?

(1) LiCl (2) $BeCl_2$
(3) $BaCl_2$ (4) $AlCl_3$

### AIIMS 2012

**21.** Which will not be precipitated after addition of $(NH_4)_2CO_3$ in presence of $NH_4Cl$ :-

(1) Mg (2) Ba
(3) Ca (4) Sr

### NEET UG 2013

**22.** Which is the strongest acid in the following ?

(1) $H_2SO_3$ (2) $H_2SO_4$
(3) $HClO_3$ (4) $HClO_4$

### AIIMS 2013

**23.** What will be the solubility product of $AX_3$ :-

(1) $27\ s^4$ (2) $4s^3$
(3) $36\ s^4$ (4) $9s^3$

### AIPMT 2014

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

(1) KCl (2) NaCl
(3) $Na_2CO_3$ (4) $CuSO_4$

### AIPMT 2015

**25.** The $K_{sp}$ of $Ag_2CrO_4$, AgCl, AgBr and AgI are respectively, $1.1 \times 10^{-12}$, $1.8 \times 10^{-10}$, $5.0 \times 10^{-13}$, $8.3 \times 10^{-17}$. Which one of the following salts will precipitate last if $AgNO_3$ solution is added to the solution containing equal moles of NaCl, NaBr, NaI and $Na_2CrO_4$ ?

(1) AgCl (2) AgBr
(3) $Ag_2CrO_4$ (4) AgI

### Re-AIPMT 2015

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

(1) $H_2CO_3$ and $Na_2CO_3$
(2) $H_3PO_4$ and $Na_3PO_4$
(3) $HClO_4$ and $NaClO_4$
(4) $CH_3COOH$ and $CH_3COONa$

**27.** What is the pH of the resulting solution when equal volumes of 0.1 M NaOH and 0.01 M HCl are mixed?

(1) 7.0 (2) 1.04
(3) 12.65 (4) 2.0

### AIIMS 2015

**28.** What will be solubility of $A_2X_3$ if its solubility product ($K_{sp}$) is equal to $1.08 \times 10^{-23}$

(1) $10^{-5}$ mol $L^{-1}$
(2) $3.7 \times 10^{-4}$ mol $L^{-1}$
(3) $1.2 \times 10^{-3}$ mol $L^{-1}$
(4) $7.5 \times 10^{-4}$ mol $L^{-1}$

**29.** pH of salt of weak acid with strong base at 25° C.

(1) = 7
(2) > 7
(3) < 7
(4) None

## NEET-I 2016

**30.** MY and $NY_3$, two nearly insoluble salts, have the same $K_{sp}$ values of $6.2 \times 10^{-13}$ at room temperature. Which statement would be **true** in regard to MY and $NY_3$ ?

(1) The molar solubilities of MY and $NY_3$ in water are identical.

(2) The molar solubility of MY in water is less than that of $NY_3$

(3) The salts MY and $NY_3$ are more soluble in 0.5 M KY than in pure water.

(4) The addition of the salt of KY to solution of MY and $NY_3$ will have no effect on their solubilities.

## NEET-II 2016

**31.** 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

(1) 0.77% (2) 1.6%

(3) 0.0060% (4) 0.013%

**32.** The solubility of AgCl(s) with solubility product $1.6 \times 10^{-10}$ in 0.1 M NaCl solution would be

(1) $1.6 \times 10^{-11}$ M (2) zero

(3) $1.26 \times 10^{-5}$ M (4) $1.6 \times 10^{-9}$ M

## AIIMS 2016

**33.** Calculate the degree of ionization of 0.04 M HOCl solution having ionization constant $1.25 \times 10^{-4}$?

(1) 0.025 (2) 0.25 (3) 0.5 (4) 0.055

**34.** Calcualte pH of a salt of weak monobasic acid and weak monoacidic base having concentration 0.1M at 25ºC (Given :- $pk_a = 4.8$, $pk_b = 5.2$)

(1) 7.2 (2) 6.8 (3) 9.6 (4) 7

**35.** The $K_{sp}$ of salt AgCl at 25°C is $2.56 \times 10^{-10}$. Then how much volume of $H_2O$ is required to dissolve 0.01 mole of salt.

(1) 800 L (2) 400 L

(3) 625 L (4) 50 L

**36.** Calculate pOH of $10^{-2}$M monobasic acid if value of $K_w$ at 90°C is $10^{-12}$?

(1) 2 (2) 10

(3) 12 (4) 7

**37.** $K_{sp}$ of $Sr(OH)_2$ is x, then calculate value of $[OH^-]$ in saturated solution?

(1) $\left(\frac{3x}{4}\right)^{1/3}$ (2) $2\left(\frac{x}{4}\right)^{1/3}$

(3) $\left(\frac{x}{4}\right)^{1/3}$ (4) $\left(\frac{x}{3}\right)^{1/4}$

## NEET(UG) 2017

**38.** 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 :-

(1) $2.66 \times 10^{-12}$ (2) $4.5 \times 10^{-11}$

(3) $5.3 \times 10^{-12}$ (4) $2.42 \times 10^{-8}$

## AIIMS 2017

**39.** Calculate the ratio of solubility of AgCl in 0.1M $AgNO_3$ and in pure water.

Given : $K_{sp}$ of AgCl = $1.8 \times 10^{-10}$ :-

(1) $1.34 \times 10^{-5}$ M (2) $1.34 \times 10^{-4}$ M

(3) $2 \times 10^{-8}$ M (4) $3 \times 10^{-6}$ M

**40.** 200 ml of 0.005 M $AgNO_3$ reacts with 300 mL of 0.01 M KCl. If $K_{sp}$ of AgCl is $1.8 \times 10^{-10}$. Then maximum conc. of $Ag^+$ in mixture is :-

(1) $2 \times 10^{-8}$ (2) $4.5 \times 10^{-8}$

(3) $4.8 \times 10^{-5}$ (4) $1.34 \times 10^{-5}$

## NEET(UG) 2018

**41.** Following solutions were prepared by mixing different volumes of NaOH and HCl of different concentrations :

a. $60\text{mL}\frac{M}{10}\text{HCl} + 40\text{mL}\frac{M}{10}\text{NaOH}$

b. $55\text{mL}\frac{M}{10}\text{HCl} + 45\text{mL}\frac{M}{10}\text{NaOH}$

c. $75\text{mL}\frac{M}{5}\text{HCl} + 25\text{mL}\frac{M}{5}\text{NaOH}$

d. $100\text{mL}\frac{M}{10}\text{HCl} + 100\text{mL}\frac{M}{10}\text{NaOH}$

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

(1) b (2) a

(3) d (4) c

**42.** The solubility of $BaSO_4$ in water $2.42 \times 10^3$ $gL^{-1}$ at 298 K. The value of solubility product ($K_{sp}$) will be
(Given molar mass of $BaSO_4$ = 233 g $mol^{-1}$)
(1) $1.08 \times 10^{-10}$ $mol^2$ $L^{-2}$
(2) $1.08 \times 10^{-12}$ $mol^2$ $L^{-2}$
(3) $1.08 \times 10^{-14}$ $mol^2$ $L^{-2}$
(4) $1.08 \times 10^{-8}$ $mol^2$ $L^{-2}$

## AIIMS 2018

**43.** Mixture of ester and HCl is titrated with NaOH using phenolphthalein as an acidic indicator at end point pink color dissappear after some time due to :-
(1) $CH_3COOH$ is formed
(2) Due to weak acedic nature of $CH_3OH$
(3) Regeneration of HCl
(4) Due to ionization of phenolphthalein

**44.** 50 ml of 0.1 M $NH_3$ solution and 10 mL of 0.1 M HCl mixed together then calculate pH of mixure. If $pk_b$ of $NH_3$ is 4.75 :-
(1) 8.86 (2) 4.74 (3) 9.86 (4) 5.32

**45.** Which of following will have maximum pH :-
(1) Human saliva (2) Black coffee
(3) Human blood (4) Gastric juice

**46.** Find concentration of $[H^+]$ after mixing of 15 mL, 0.1M $H_2SO_4$ and 15ml 0.1M NaOH :-
(1) $5 \times 10^{-2}$ (2) $10^{-7}$
(3) $1 \times 10^{-2}$ (4) $2 \times 10^{-2}$

**47.** Find the concentration of $[H^+]$ after mixing 20 ml, 0.1M $CH_3COOH$ and 10 ml, 0.1 M NaOH. Given $K_a$ for $CH_3COOH$ $1.7 \times 10^{-5}$.
(1) $1.7 \times 10^{-5}$ (2) $3.4 \times 10^{-5}$
(3) $3.4 \times 10^{-6}$ (4) $1.7 \times 10^{-4}$

**48.** Solubility of $XB_2$ is X mol/lit in water find out solubility of $XB_2$ in $10^{-3}$ M YB. Consider YB is completely soluble salt :-
(1) $4X^3 \times 10^{+6}$ (2) $4X^3 \times 10^{-6}$
(3) $10^{-3}$ (4) $X^3$

**49.** 0.02M of pyridium hydrochloride having $[H^+]$ is $5.55 \times 10^{-6}$ what is ionisation constant of pyridine.
(1) $1.54 \times 10^{-9}$ (2) $1.54 \times 10^{-7}$
(3) $1.54 \times 10^{-6}$ (4) $1.54 \times 10^{-5}$

## EXERCISE-II (Previous Year Questions) ANSWER KEY

| Que. | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| **Ans.** | 1 | 2 | 3 | 2 | 2 | 1 | 1 | 2 | 2 | 1 | 1 | 1 | 1 | 3 | 4 |
| **Que.** | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | 29 | 30 |
| **Ans.** | 2 | 1 | 4 | 3 | 3 | 1 | 4 | 1 | 3 | 3 | 3 | 3 | 1 | 2 | 2 |
| **Que.** | 31 | 32 | 33 | 34 | 35 | 36 | 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 |
| **Ans.** | 4 | 4 | 4 | 2 | 3 | 2 | 2 | 3 | 2 | 2 | 4 | 1 | 4 | 3 | 3 |
| **Que.** | 46 | 47 | 48 | 49 | | | | | | | | | | | |
| **Ans.** | 1 | 1 | 1 | 1 | | | | | | | | | | | |

## EXERCISE-III (Analytical Questions)

## Check Your Understanding

**1.** The concentration of $[H^+]$ and concentration of $[OH^-]$ of a 0.1 M aqueous solution of 2% ionised weak acid is [ionic product of water = $1 \times 10^{-14}$] :-
(1) $0.02 \times 10^{-3}$ M and $5 \times 10^{-11}$ M
(2) $1 \times 10^{-3}$ M and $3 \times 10^{-11}$ M
(3) $2 \times 10^{-3}$ M and $5 \times 10^{-12}$ M
(4) $3 \times 10^{-2}$ M and $4 \times 10^{-13}$ M

**2.** The solubility of $BaSO_4$ in water, is $2.33 \times 10^{-3}$ $gL^{-1}$. Its solubility product will be (molecular weight of $BaSO_4$ = 233) :-
(1) $1 \times 10^{-5}$ (2) $1 \times 10^{-10}$
(3) $1 \times 10^{-15}$ (4) $1 \times 10^{-20}$

**3.** What will be the $H^+$ ion concentration, when 4 g NaOH dissolved in 1000 mL of water:
(1) $10^{-1}$ (2) $10^{-13}$ (3) $10^{-4}$ (4) $10^{-10}$

**4.** When 10 mL of 0.1 M acetic acid (pKa=5.0) is titrated against 10 mL of 0.1 M ammonia solution ($pK_b$ = 5.0), the equivalence point occurs at pH:
(1) 5.0 (2) 6.0 (3) 7.0 (4) 9.0

**5.** At 25°C, the dissociation constant of a base, BOH, is $1.0 \times 10^{-12}$. The concentration of hydroxyl ions in 0.01M aqueous solution of the base would be:
(1) $1.0 \times 10^{-6}$ mol $L^{-1}$ (2) $1.0 \times 10^{-7}$ mol $L^{-1}$
(3) $2.0 \times 10^{-6}$ mol $L^{-1}$ (4) $1.0 \times 10^{-5}$ mol $L^{-1}$

**6.** The solubility product of $BaSO_4$ at 25°C is $1.0 \times 10^{-9}$. What would be the concentration of $H_2SO_4$ necessary to precipitate $BaSO_4$ from a solution of 0.01 M $Ba^{+2}$ ions :-
(1) $10^{-9}$ (2) $10^{-8}$ (3) $10^{-7}$ (4) $10^{-6}$

**7.** pH of the solution of $HCOONH_4$ is 6.48 this can be explained by :-
(1) Hydrolysis of both cation and anion
(2) Hydrolysis of cation
(3) Hydrolysis of anion
(4) Hydrolysis of water

**8.** The correct representation of solubility product of $SnS_2$ is :-
(1) $[Sn^{4+}][S^{2-}]^2$ (2) $[Sn^{4+}][S^{2-}]$
(3) $[Sn^{4+}][2S^{2-}]$ (4) $[Sn^{4+}][2S^{2-}]^2$

**9.** A solution of $FeCl_3$ in water acts as acidic due to:-
(1) Acidic impurities
(2) Ionisation
(3) Hydrolysis of $Fe^{3+}$
(4) Dissociation

**10.** The pKa of $HNO_2$ is 3.37. The pH of $HNO_2$ in its 0.01 mol $L^{-1}$ aqueous solution will be :-
(1) 5.37 (2) 2.69 (3) 1.69 (4) 0.69

**11.** When 0.01 M HCl is added in aqueous solution of acetic acid then :-
(1) $[CH_3COO^-]$ decreases
(2) $[CH_3COOH]$ decreases
(3) $[CH_3COO^-]$ increases
(4) NOT

**12.** On passing $H_2S$ gas through a solution of $Cu^{+2}$ and $Zn^{+2}$ ions, CuS is precipitated first because:-
(1) Solubility product of CuS is equal to the ionic product of ZnS ;
(2) Solubility product of CuS is equal to the solubility product of ZnS.
(3) Solubility product of CuS is lower than the solubility product of ZnS.
(4) Solubility product of CuS is greater than the solubility product of ZnS.

**13.** Solubility of $MX_2$ – type electrolytes is $0.5 \times 10^{-4}$ mol $L^{-1}$ then find out $K_{sp}$ of electrolytes:-
(1) $5 \times 10^{-12}$ (2) $25 \times 10^{-10}$
(3) $1 \times 10^{-13}$ (4) $5 \times 10^{-13}$

**14.** A solution of $MgCl_2$ in water has pH :
(1) < 7 (2) > 7 (3) 7 (4) 14.2

**15.** When $H_2S$ gas is passed through the HCl containing aqueous solutions of $CuCl_2$, $HgCl_2$, $BiCl_3$ and $CoCl_2$, it does not precipitate out :
(1) CuS (2) HgS (3) $Bi_2S_3$ (4) CoS

**16.** Mark the correct statement :
(1) I group basic radicals precipitate as chloride.
(2) IV group basic radicals precipitate as sulphides
(3) V group basic radicals precipitate as carbonates
(4) All the above statements are correct

**17.** The role of $NH_4Cl$ in group III for analysis of cations is :–
(1) that it acts as a catalyst.
(2) that it increases the solubility of hydroxides.
(3) that it lowers the $OH^-$ concentration.
(4) that it causes the precipitation of cations as chlorides.

**18.** $K_{sp}$ of $Ca_3(PO_4)_2$ is :-
(1) $[Ca^{+2}][PO_4^{-3}]^2$ (2) $[Ca^{+2}]^3[PO_4^{-3}]^2$
(3) $[Ca^{+2}]^2[PO_4^{-3}]^3$ (4) $[Ca^{+2}][PO_4^{-3}]$

**19.** An acid HA has dissociated as following manner

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

It has concentration 1 M and pH = 5 then find out dissociation constant :-

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

(3) $5 \times 10^{-5}$ (4) 5

**20.** Which of the following group of cations will be precipitated when passing of $H_2S$ gas in the presence of acidic medium:-

(1) $Cu^{+2}$, $Cr^{+3}$ (2) $Zn^{+2}$, $Co^{+2}$

(3) $Cu^{+2}$, $Cd^{+2}$ (4) $Al^{+3}$, $Cd^{+2}$

**21.** Solution of sodium carbonate is :-

(1) Strongly acidic (2) Weakly basic

(3) Strongly basic (4) Weakly acidic

**22.** Aqueous solution of $Al_2(SO_4)_3$ is :–

(1) Basic & acidic (2) Neutral

(3) Basic (4) Acidic

**23.** In a saturated solution of the sparingly soluble strong electrolyte $AgIO_3$ (Moleculatr mass =283). The equilibrium which sets in is $AgIO_{3(S)} \rightleftharpoons Ag^+_{(aq)} + IO^-_{3(aq)}$ If the solubility product constant $K_{sp}$ of $AgIO_3$ at a given temperature is $1.0 \times 10^{-8}$, what is the mass of $AgIO_3$ contained in 100 mL of its saturated solution :

(1) $2.83 \times 10^{-3}$ g (2) $1.0 \times 10^{-7}$ g

(3) $1.0 \times 10^{-4}$ g (4) $28.3 \times 10^{-2}$ g

**24.** If $k_{sp}$ of $CaF_2$ in pure water is $1.70 \times 10^{-10}$, then find the solubility of $CaF_2$ in 0.10M NaF solution :-

(1) $1.70 \times 10^{-10}$ (2) $1.70 \times 10^{-9}$

(3) $1.70 \times 10^{-8}$ (4) 0.10 M

**25.** The solubility product of AgCl is $1 \times 10^{-10}$, then solubility of AgCl is :-

(1) $1 \times 10^{-10}$ (2) $1 \times 10^{-7}$

(3) $1 \times 10^{-5}$ (4) $1 \times 10^{-8}$

**26.** To precipitate Zn in form of ZnS, Why $NH_4OH$ is first added in the solution before $H_2S$ gas is passed through it :-

(1) To convert Zn into $Zn^{+2}$

(2) To reduce Zinc

(3) To decrease the dissociation of $H_2S$

(4) To increase the dissociation of $H_2S$

**27.** Ka for $CH_3COOH$ is $1.8 \times 10^{-5}$. Find out the percentage dissociation of 0.2M $CH_3COOH$ in 0.1M HCl solution

(1) 0.018 (2) 0.36 (3) 18 (4) 36

**28.** The pKa of a weak acid HA is 4.80. The $pK_b$ of weak base BOH is 4.78. The pH of an aqueous solution of the corresponding salt BA will be :

(1) 9.58 (2) 4.79 (3) 7.01 (4) 9.22

**29.** 0.2M solution of HCOOH is 3.2% ionised then find ionisation constant of acid :-

(1) $4.2 \times 10^{-4}$ (2) $4.2 \times 10^{-5}$

(3) $2.1 \times 10^{-4}$ (4) $2.1 \times 10^{-5}$

**30.** Three reactions involving $H_2PO_4^-$ are given below :-

(i) $H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$

(ii) $H_2PO_4^- + H_2O \rightarrow HPO_4^{2-} + H_3O^+$

(iii) $H_2PO_4^- + OH^- \rightarrow H_3PO_4 + O^{2-}$

In which of the above does $H_2PO_4^-$ act as an acid ?

(1) (i) only (2) (ii) only

(3) (i) and (ii) (4) (iii) only

**31.** Given that for HA acid, $K_a = 10^{-6}$ and for MOH base $K_b = 10^{-6}$. The pH of 0.1 M MA salt solution will be :-

(1) 5 (2) 7 (3) 9 (4) 2

**32.** The $K_{sp}$ for $Cr(OH)_3$ is $1.6 \times 10^{-30}$. The molar solubility of this compound in water is :-

(1) $\sqrt[2]{1.6 \times 10^{-30}}$ (2) $\sqrt[4]{1.6 \times 10^{-30}}$

(3) $\sqrt[4]{1.6 \times 10^{-30}/27}$ (4) $1.6 \times 10^{-30}/27$

**33.** An acid HA ionises as

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

The pH of 1.0 M solution is 5. Its dissociation constant would be :-

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

(3) $5 \times 10^{-8}$ (4) $1 \times 10^{-5}$

**34.** The pH of a 0.1 molar solution of the acid HQ is 3. The value of the ionization constant, Ka of this acid is :-

(1) $1 \times 10^{-7}$ (2) $3 \times 10^{-7}$

(3) $1 \times 10^{-3}$ (4) $1 \times 10^{-5}$

**35.** How many litres of water must be added to 1 litre of an aqueous solution of HCl with a pH of 1 to create an aqueous solution with pH of 2 ?

(1) 0.1 L (2) 0.9 L (3) 2.0 L (4) 9.0 L

**36.** What volume of 0.1M $H_2SO_4$ is needed to completely neutralize 40 mL of 0.2M NaOH solution–

(1) 10 mL (2) 40 mL (3) 20 mL (4) 80 mL

**37.** If pH value of a solution is 3 and by adding water, it becomes 6, then the dilution is increased by :
(1) 10 times
(2) 100 times
(3) 500 times
(4) 1000 times

**38.** The aqueous solution of which of the following salt will have the lowest pH
(1) NaClO
(2) $NaClO_2$
(3) $NaClO_3$
(4) $NaClO_4$

**39.** In the reaction $I_2 + I^- \rightarrow I_3^-$, the Lewis base is :
(1) $I^-$
(2) $I_2$
(3) $I_3^-$
(4) None of these

**40.** In which of the following acid–base titration, pH is greater than 8 at equivalence point:-
(1) Acetic acid versus ammonia
(2) Acetic acid versus sodium hydroxide
(3) Hydrochloric acid versus ammonia
(4) Hydrochloric acid versus sodium hydroxide

**41.** Which one of the following compounds is not a protonic acid :-
(1) $B(OH)_3$
(2) $PO(OH)_3$
(3) $SO(OH)_2$
(4) $SO_2(OH)_2$

**42.** Which one of the following is NOT a buffer solution:-
(1) 0.8M $H_2S$ + 0.8M KHS
(2) 2M $C_6H_5NH_2$ + 2M $C_6H_5\overset{+}{N}H_3Br$
(3) 3M $H_2CO_3$ + 3M $KHCO_3$
(4) 0.05M $KClO_4$ + 0.05M $HClO_4$

**43.** The rapid change of pH near the stoichiometric point of an acid–base titration is the basis of indicator detection. pH of the solution is related to ratio of the concentrations of the conjugate acid (HIn) and base ($In^-$) forms of the indicator by the expression:-

(1) $\log \dfrac{[HIn]}{[In^-]} = pK_{In} - pH$

(2) $\log \dfrac{[HIn]}{[In^-]} = pH - pK_{In}$

(3) $\log \dfrac{[In^-]}{[HIn]} = pH + pK_{In}$

(4) $\log \dfrac{[In^-]}{[HIn]} = pK_{In} - pH$

**44.** What is the pH of 0.01M glycine solution? For glycine, $Ka_1 = 4.5 \times 10^{-3}$ and $Ka_2 = 1.7 \times 10^{-10}$ at 298 K.
(1) 3·0
(2) 10·0
(3) 6.1
(4) 7·2

**45.** The correct order of acid strength is –
(1) $HClO_4 < HClO_3 < HClO_2 < HClO$
(2) $HClO_2 < HClO_3 < HClO_4 < HClO$
(3) $HClO_4 < HClO < HClO_2 < HClO_3$
(4) $HClO < HClO_2 < HClO_3 < HClO_4$

**46.** Concentrations of $NH_4Cl$ and $NH_4OH$ in a buffer solution are in the ratio 1 : 10. If $K_b$ for $NH_4OH$ is $10^{-10}$, then pH of the buffer is :–
(1) 4
(2) 5
(3) 9
(4) 11

**47.** 10 mL concentrated $H_2SO_4$ (18 molar) is diluted to 1 litre.Concentration of diluted acid is:-
(1) 0.18N
(2) 0.09N
(3) 0.36N
(4) 18N

**48.** When HF is dissolved in formic acid, the equilibrium established is :–

$$HF + HCOOH \rightleftharpoons F^- + HCOOH_2^+$$

the true pair of conjugate acid base is as a [acid, conjugate acid] [Base, conjugate base]:–
(1) (HF, HCOOH) and ($HCOOH_2^+$, $F^-$)
(2) (HF, $HCOOH_2^+$) and (HCOOH, $F^-$)
(3) ($HCOOH_2^+$, HF) and ($F^-$, HCOOH)
(4) (HF, $F^-$) and ($HCOOH_2^+$, HCOOH)

**49.** How many grams of dibasic acid (mol. wt. = 200) should be present in 100 mL the aqueous solution to give strength of (N/10) :–
(1) 1g.
(2) 2g.
(3) 5g.
(4) 10g.

**50.** In a mixture of equimolar solutions of $NaHCO_3$ and NaOH the species present in solution shall be:–
(1) $Na_2CO_3$
(2) $NaHCO_3$ + NaOH
(3) NaOH
(4) $NaHCO_3 + Na_2CO_3$

**51.** An aqueous solution contains $10^{-4}$ [$H^+$]. If it is diluted by mixing equal volume of water then the concentration of $OH^-$ in mol $dm^{-3}$ will be :–
(1) $0·5 \times 10^{-10}$
(2) $2 \times 10^{-10}$
(3) $10^{-6}$
(4) $10^{-8}$

**52.** Which of the following is right for diprotic acid :
(1) $Ka_2 > Ka_1$
(2) $Ka_1 > Ka_2$
(3) $Ka_2 > \dfrac{1}{Ka_1}$
(4) $Ka_2 = Ka_1$

**53.** The first and second dissociation constants of an acid $H_2A$ are $1.0 \times 10^{-5}$ and $5.0 \times 10^{-10}$ respectively. The overall dissociation constant of the acid will be:-
(1) $5.0 \times 10^{15}$ (2) $5.0 \times 10^{-15}$
(3) $0.2 \times 10^{5}$ (4) $5.0 \times 10^{-5}$

**54.** The pKa of a weak acid (HA) is 4.5. The pOH of an aqueous beffered solution of HA in which 50% of the acid is ionized is :-
(1) 2.5 (2) 9.5 (3) 7.0 (4) 4.5

**55.** 50 mL solution of 0.1M $CH_3COOH$ ($pK_a = 4.73$) is titrated with 0.1M NaOH solution, pH of solution when half of $CH_3COOH$ is neutralized
(1) 4.53 (2) 4.63 (3) 4.73 (4) 4.83

**56.** $pK_a$ of Quinoline base is 4.88. What will be the $pK_a$ of 0.01M solution of it
(1) 4.88 (2) 0.01 (3) 9.12 (4) 14

**57.** The pH of an aqueous solution of a $1 \times 10^{-7}$ M solution of HCl will be :-
(1) 7 (2) slightly less than 7
(3) slightly greater than 7 (4) 1

**58.** What will be the concentration of $H^{\oplus}$ ions in 0.1M acetic acid and 0.1M sodium acetate solution, if the dissociation constant of acetic acid is $1.8 \times 10^{-5}$ ?
(1) $1.8 \times 10^{-7}$ (2) $1.8 \times 10^{-5}$
(3) $1.8 \times 10^{-2}$ (4) $1.8 \times 10^{-3}$

**59.** Calculate the pH of $4 \times 10^{-3}$ M, $Y(OH)_2$ solution assuming the first dissociation to be 100% second dissociation to be 50%:-
(1) 11.78 (2) 9.9 (3) 2.5 (4) 2.22

## EXERCISE-III (Analytical Questions) ANSWER KEY

| Que. | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| **Ans.** | **3** | **2** | **2** | **3** | **2** | **4** | **1** | **1** | **3** | **2** | **1** | **3** | **4** | **1** | **4** |
| Que. | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | 29 | 30 |
| **Ans.** | **4** | **3** | **2** | **1** | **3** | **2** | **4** | **1** | **3** | **3** | **4** | **1** | **3** | **3** | **2** |
| Que. | 31 | 32 | 33 | 34 | 35 | 36 | 37 | 38 | 39 | 40 | 41 | 42 | 43 | 44 | 45 |
| **Ans.** | **2** | **3** | **1** | **4** | **4** | **2** | **4** | **4** | **1** | **2** | **1** | **4** | **1** | **3** | **4** |
| Que. | 46 | 47 | 48 | 49 | 50 | 51 | 52 | 53 | 54 | 55 | 56 | 57 | 58 | 59 | |
| **Ans.** | **2** | **3** | **2** | **1** | **1** | **2** | **2** | **2** | **2** | **3** | **1** | **2** | **2** | **1** | |

## EXERCISE-IV (Assertion & Reason) Target AIIMS

### Directions for Assertion & Reason questions

These questions consist of two statements each, printed as Assertion and Reason. While answering these Questions you are required to choose any one of the following four responses.

**(A)** If both Assertion & Reason are True & the Reason is a correct explanation of the Assertion.

**(B)** If both Assertion & Reason are True but Reason is not a correct explanation of the Assertion.

**(C)** If Assertion is True but the Reason is False.

**(D)** If both Assertion & Reason are false.

**1.** ***Assertion :–*** When small amount of acid or base is added to pure water, its pH undergoes a change.
***Reason :–*** Addition of an acid or base increases the degree of ionisation of water.
(1) A (2) B (3) C (4) D

**2.** ***Assertion :–*** Solubility of AgBr decreases in the presence of sodium bromide.
***Reason :–*** Sodium bromide undergoes hydrolysis in water.
(1) A (2) B (3) C (4) D

**3.** ***Assertion :–*** The pH of an aqueous solution of acetic acid remains unchanged on the addition of sodium acetate.
***Reason :–*** The ionisation of acetic acid is suppressed by the addition of sodium acetate.
(1) A (2) B (3) C (4) D

**4.** ***Assertion :–*** If HCl gas is passed through saturated NaCl solution, solid NaCl starts to precipitate.
***Reason :–*** HCl decreases the solubility product of NaCl.
(1) A (2) B (3) C (4) D

**5.** ***Assertion :–*** The aqueous solution of $Na_3PO_4$ is alkaline in nature.
***Reason :–*** $Na_3PO_4$ in its aqueous solution undergoes anionic hydrolysis.
(1) A (2) B (3) C (4) D

**6.** ***Assertion :–*** If $K_{sp}$ < ionic product, precipitate is formed.
***Reason :–*** Solubility product ($K_{sp}$) is the highest limit of ionic product of the electrolyte in saturated solutions.
(1) A (2) B (3) C (4) D

**7.** ***Assertion :–*** To precipitate the cations of fourth group in qualitative analysis, medium is made alkaline before passing $H_2S$ gas.
***Reason :–*** This is done to suppress the ionisation of $H_2S$.
(1) A (2) B (3) C (4) D

**8.** ***Assertion :–*** When an acid or a base is added to water at constant temperature the pH changes.
***Reason :–*** This is due to change in ionic product of water.
(1) A (2) B (3) C (4) D

**9.** ***Assertion*** : Addition of $NH_4OH$ to an aqueous solution of $BaCl_2$ in the presence of $NH_4Cl$(excess) precipitates $Ba(OH)_2$.
***Reason*** : $Ba(OH)_2$ is insoluble in water.
(1) A (2) B (3) C (4) D

**10.** ***Assertion :–*** Sb(III) is not precipitated as sulphide when in its alkaline solution $H_2S$ is passed.
***Reason :–*** The concentration of $S^{2-}$ ion in alkaline medium is inadequate for precipitation.
(1) A (2) B (3) C (4) D

**11.** ***Assertion :-*** For $H_2SO_4$, $K_{a_1}$ value is much greater than $Ka_2$.
***Reason :-*** $H_2SO_4$ is a strong acid.
(1) A (2) B (3) C (4) D

**12.** ***Assertion :-*** Solubility of salt of SB and WA changes when pH of the solution is changed.
***Reason :-*** Solubility product ($K_{sp}$) depends on pH.
(1) A (2) B (3) C (4) D

**13.** ***Assertion :–*** A mixture of a weak acid $CH_3COOH$ and sodium acetate forms a buffer solution.
***Reason :–*** A buffer solution reacts with small quantities of hydrogen or hydroxyl ions and keeps the pH almost same.
(1) A (2) B (3) C (4) D

**14.** ***Assertion :–*** At 25°C the pH of $10^{-8}$ M HCl is 8.
***Reason :–*** pH of acidic solution is always below 7 at 25°C.
(1) A (2) B (3) C (4) D

**15.** ***Assertion :–*** In the acid base titration involving a strong base and a weak acid methyl orange can be used as an indicator.

***Reason :–*** Methyl orange changes its colour in pH range 3 to 5.

(1) A (2) B (3) C (4) D

**16.** ***Assertion :–*** pH of a buffer changes with temperature.

***Reason :–*** Ionic product of water ($K_w$) changes with temperature.

(1) A (2) B (3) C (4) D

**17.** ***Assertion :–*** $H_3PO_3$ is a dibasic compound.

***Reason :–*** The two H-atom are directly attached to P.

(1) A (2) B (3) C (4) D

**18.** ***Assertion :*** Boric acid behaves as a weak monobasic acid.

***Reason :*** Boric acid contains hydrogen bonds in its structure.

(1) A (2) B (3) C (4) D

**19.** ***Assertion :*** $H_2O$ is amphoteric in nature.

***Reason :*** $H_2O$ can accept a proton to form $H_3O^+$ ion and can donate a proton to form $OH^-$ ion.

(1) A (2) B (3) C (4) D

**20.** ***Assertion :*** All Arrhenius acids are also Bronsted acids.

***Reason :*** All Bronsted bases are also Lewis bases.

(1) A (2) B (3) C (4) D

**21.** ***Assertion :–*** The buffer solution has a capacity to resist the change in pH value on addition of small amount of acid or base to it.

***Reason :–*** pH value of buffer solution does not change on dilution or on keeping for long.

(1) A (2) B (3) C (4) D

**22.** ***Assertion :–*** The species in the buffer must not react with each other.

***Reason :–*** The pH of a buffer depends on the value of $K_a$ of the weak acid and the relative concentration of that acid and its conjugate base.

(1) A (2) B (3) C (4) D

**23.** ***Assertion :–*** Only a very small amount of indicator should be used.

***Reason :–*** So that addition of the indicator does not effect the pH of the solution.

(1) A (2) B (3) C (4) D

**24.** ***Assertion :-*** Solubility of sparingly soluble salt increases with increase in temperature.

***Reason :-*** Enthalpy of solution is negative for all salts.

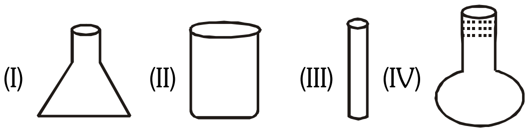
(1) A (2) B (3) C (4) D

**25.** ***Assertion :-*** HCl is titrated with NaOH using phenolphthalein as indicator, pink colour appears in solution when $OH^-$ ion is present in excess in solution.

***Reason :-*** At end point, both reacting species, neutralises each other.

(1) A (2) B (3) C (4) D

**26.** ***Assertion :-***



IInd is more suitable for titration.

***Reason :-*** It is difficult to add titrant in first.

(1) A (2) B (3) C (4) D

**27.** ***Assertion :*** In $NaHCO_3$ solution phenolpthalein is colorless.

***Reason :*** Phenolpthalein is colorless in basic medium and pH of $NaHCO_3$ is less than 10.

(1) A (2) B (3) C (4) D

**28.** ***Assertion :-*** $Na_2SO_3$ is alkaline towards litmus.

***Reason :-*** $H_2SO_3$ is formed due to hydrolysis.

(1) A (2) B (3) C (4) D

## EXERCISE-IV (Assertion & Reason) ANSWER KEY

| Que. | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| **Ans.** | **3** | **3** | **4** | **3** | **1** | **1** | **3** | **3** | **4** | **4** | **2** | **3** | **1** | **4** | **4** |
| Que. | 16 | 17 | 18 | 19 | 20 | 21 | 22 | 23 | 24 | 25 | 26 | 27 | 28 | | |
| **Ans.** | **2** | **3** | **2** | **1** | **3** | **2** | **2** | **1** | **3** | **2** | **4** | **4** | **2** | | |