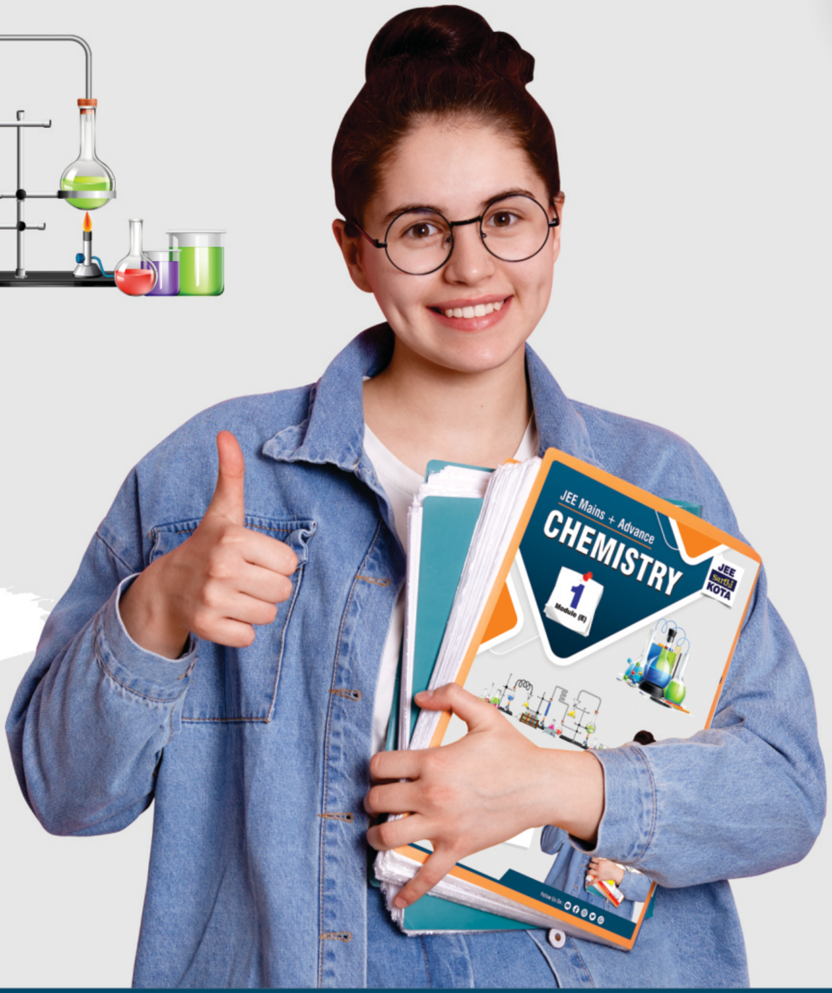
JEE Mains + Advance

CHEMISTRY

JEE Sarthi KOTA

Sample Module



Chapter

IONIC EQUILIBRIUM

JEE Module Details

(Total = 24)

CLASS - XI : 12 MODULES

PHYSICS

Module - 1

Ch. No.	Chapter Name
1.	Mathematical Tools
2.	Vector
3.	Unit, Dimension and Measurement
4.	Kinematics
5.	Newton's Laws of Motion

Module - 2

Ch. No.	Chapter Name
1.	Work Power and Energy
2.	Center of Mass & Collision
3.	Rotational Motion
4.	Gravitation

Module - 3

Ch. No.	Chapter Name
1.	Fluid Mechanics
2.	Surface Tension
3.	Elasticity & Viscosity
4.	Simple Harmonic Motion

Module - 4

Ch. No.	Chapter Name
1.	Thermometry & Calorimetry
2.	Thermal Expansion
3.	Kinetic Theory of Gases
4.	Thermodynamics
5.	Heat Transfer

CHEMISTRY

Module - 1

Ch. No.	Chapter Name
1.	Some Basic Concept of Chemistry
2.	Atomic Structure
3.	Redox Reactions
4.	States of Matter

Module - 2

Ch. No.	Chapter Name
1.	Chemical Equilibrium
2.	Ionic Equilibrium
3.	Chemical Thermodynamics & Energetics

Module - 3

Ch. No.	Chapter Name
1.	Periodic Table and Periodic Properties
2.	Chemical Bonding
3.	Hydrogen and its compounds
4.	s-Block elements
5.	p-Block (13 to 14 groups)

Module - 4

Ch. No.	Chapter Name
1.	IUPAC
2.	Isomerism
3.	GOC-I
4.	Hydrocarbons
5.	Environmental Chemistry

MATHEMATICS

Module - 1

Ch. No.	Chapter Name
1.	Set & Relations
2.	Trigonometric Ratios
3.	Trigonometric Equation
4.	Solution of a Triangle

Module - 2

Ch. No.	Chapter Name
1.	Sequence and Series
2.	Quadratic Equations and Inequalities
3.	Complex Numbers
4.	Limits & Derivative

Module - 3

Ch. No.	Chapter Name
1.	Binomial Theorem
2.	Permutations and Combinations
3.	Straight Lines
4.	Circle

Module - 4

Ch. No.	Chapter Name
1.	Parabola
2.	Hyperbola
3.	Ellipse

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JEE : Chemistry

Sample Module

STUDENT NAME:____

SECTION: ___________________________ROLL NO: ____________

CONTENTS

❖ PREFACE ❖

This module covers the theoretical concepts associated with NEET syllabus and contain sufficient multiple choice and previous year questions. We are confident that students would find this module helpful for their preparations.

Research & Development team of NEET Sarthi keeps working to improve the study material. Suggestions and inputs from students and readers are always welcome.

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"If you can dream it, you can do it" **-Dr. A.P.J. Abdul kalam**

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Chapter-01

Ionic Equilibrium

- **OSTWALD'S DILUTION LAW**
- **BIODATA OF WATER**
- **SALTS, TYPES OF SALT AND CONJUGATE THEORY**
- **HYDROLYSIS OF SALTS**
- **SOLUBILITY AND SOLUBILITY PRODUCT (Ksp)**
- **Buffer solution**
- **Acid and Base concept**

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

$= \log 3 + \log 10$

$= \log 0.4471 + 1 = 1.4471$

(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) = $10^{0.3010} = 2$

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. $\quad pH = -\log[H^+] = \log \frac{1}{[H^+]}$

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

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

Conclusion:

If $\quad pH = x \quad$ then $\quad [H^+] = 10^{-x}$

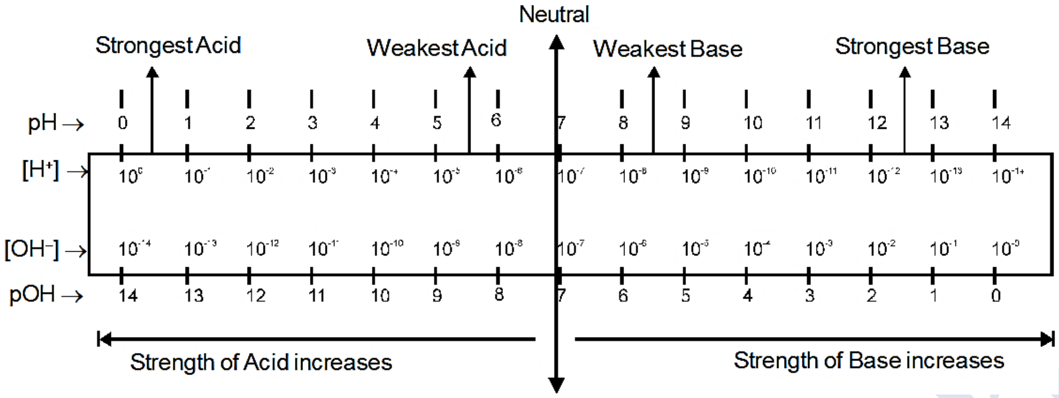
or $\quad$ Vice versa

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

pOH ⟶ It is equal to $-\log [OH^-]$

i.e. $pOH = -\log [OH^-] = \log \frac{1}{[OH^-]}$

pH scale at 25°C :



Ex.1 If $X = \frac{a}{b} \times 10^{-c}$, then find pX?

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

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

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

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.

Ex. Na^+Cl^-, K^+Cl^-, etc.

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 α 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.

↓

H_3BO_3

(boric acid)

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

(A) ARRHENIUS CONCEPT:

(1) According to Arrhenius when an electrolyte dissolves in aqueous solution it spilts up into two oppositely charged particles i.e. cation and anion.

(2) In an electrolytic solution (Aqueous solution of electrolyte), total +ve charge = 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.

(3) Properties of an electrolytic solution are defind by its ions.

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

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

(5) Ionization is the reversible reaction. In it 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 Law of Mass Action (L.O.M.A.)

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

(B) OSTWALD'S DILUTION LAW

⇒ Ostwald was the first to apply law of mass action to ionic equilibrium.

⇒ 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.O.M.A.

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

$$K = \frac{C\alpha \times C\alpha}{C(1-\alpha)} \qquad = \frac{C\alpha^2}{(1-\alpha)}$$

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

$K = C\alpha_2$ or $\alpha = \sqrt{\dfrac{K}{C}}$ (K = constant, At constant temperature)

$\alpha \propto \dfrac{1}{\sqrt{C}}$ $(C \propto \dfrac{1}{V})$

$\alpha \propto$

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

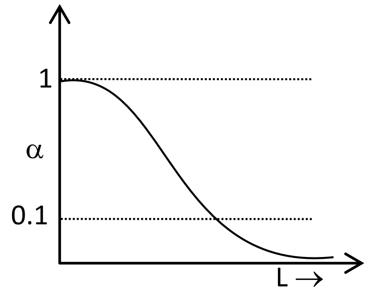
dilution ↑ α ↑

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.

Note: (1) At infinite dilution the value of α becomes equal to one.

(2) The value of concentration should be in terms of normality.



Application of Ostwald's Dilution Law :

$$\boxed{K = C\alpha^2}$$

(a) For mono basic weak acid (HA)	**(b)** For mono acidic weak base (BOH)
(i). Ionisation constant (K_a)	(i). Ionisation constant (K_b)
$HA \rightleftharpoons H^+ + A^-$	$BOH \rightleftharpoons B^+ + OH^-$
Initial concentration C 0 0	Initial concentration C 0 0
At equilibrium $C - C\alpha$ $C\alpha$ $C\alpha$	At equilibrium $C - C\alpha$ $C\alpha$ $C\alpha$
If degree of ionisation is α	If degree of ionisation is α

$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha \times C\alpha}{C - C\alpha}$	$K_b = = \frac{C\alpha \times C\alpha}{C - C\alpha}$
$K_a = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$	$K_b = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$
$\because \alpha <<<< 1 \quad \therefore (1 - \alpha) \approx 1$	$\because \alpha <<<< 1 \quad \therefore (1 - \alpha) \approx 1$
$\therefore K_a = C\alpha^2$	$\therefore K_b = C\alpha^2$

Where $K_a \rightarrow$ Ionisation constant of weak acid

$K_b \rightarrow$ Ionisation constant of weak base

(ii). $[H^+]$ (Concentration of H^+)	(ii). $[OH^-]$ (Concentration of OH^-)
$[H^+] = C\alpha$ ....(1)	$[OH^-] = C\alpha$ ....(1)
$K_a = C\alpha^2$ or $\alpha = \sqrt{\frac{K_a}{C}}$ ....(2)	$K_b = C\alpha^2$ or $\alpha = \sqrt{\frac{K_b}{C}}$ ....(2)
from Eq. (1) and (2)	from Eq. (1) and (2)
$[H^+] = C \times \frac{\sqrt{K_a}}{\sqrt{C}}$	$[OH^-] = C \times \frac{\sqrt{K_b}}{\sqrt{C}}$
$[H^+] = \sqrt{K_a \times C}$	$[OH^-] = \sqrt{K_b \times C}$
(iii). $pH = -\log [H^+]$	(iii). $pOH = -\log [OH^-]$
put the value of $[H^+]$	put the value of $[OH^-]$
$pH = -\log(\sqrt{K_a \times C}) = -\log(K_a \times C)^{1/2}$	$pOH = -\log(\sqrt{K_b \times C}) = -\log (K_b \times C)^{1/2}$
$pH = -\frac{1}{2}[\log K_a + \log C]$	$pOH = -\frac{1}{2}[\log K_b + \log C]$
$pH = -\frac{1}{2}\log K_a - \frac{1}{2}\log C$	$pOH = -\frac{1}{2}\log K_b - \frac{1}{2}\log C$
$pH = \frac{1}{2}pK_a - \frac{1}{2}\log C$	$pOH = \frac{1}{2}pK_b - \frac{1}{2}\log C$
In summary :	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 :

(1) It is not applicable for strong electrolytes.

(2) It is not applicable for saturated solutions.

Factors affecting the Value of Degree of ionisation :

(1) Temperature → On increasing temperature, ionization increases so α increases.

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

(3) Nature of electrolytes

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

(4) Nature of solvent

If Dielectric constant μ of solvent increases then the value of α increases.

$H_2O \rightarrow \mu = 81$

$D_2O \rightarrow \mu = 79$

$C_6H_6 \rightarrow \mu = 2.5$

$CCl_4 \rightarrow \mu = 0$

BIODATA OF WATER

(a) Nature of water is neutral.

i.e. at 25°C pH = 7 and pOH = 7

$\therefore$ pH = pOH

(b) Concentration of H^+ and OH^- ions in 1 litre water

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

(c) Number of H^+ and OH^- ions in 1 litre water

Number of H^+ ions $= 10^{-7}N_A$ and number of OH^- ions $= 10^{-7}\ N_A$

(d) Number of H_2O moles in 1 litre water 1000/18 = 55.5 moles

(e) Number of H_2O molecules in 1 litre water $= 55.5\ N_A$ (N_A= avogadro's number)

(f) Molar concentration of H_2O molecules in water = 55.5 mol L^{-1}

(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 55.5×10^7 H_2O molecules

So, degree of ionisation of water

$$\alpha = \frac{x}{a} = \frac{1}{55.5\times10^{+7}} = 18 \times 10^{-10} = 1.8 \times 10^{-9}$$

$$\boxed{\alpha\% = 1.8\times10^{-7}\%}$$

Hence, water is a very weak electrolyte.

K (Ionisation constant of water)

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

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

$$K = \frac{10^{-7}\times10^{-7}}{55.5} \quad \text{or} \quad \boxed{K = 1.8 \times 10^{-16}}$$

(h) Ionic product of water

Represented by [K_w]

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

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

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

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

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

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]

$pK_W = pH + pH$ $\quad$ $pK_W = pOH + pOH$

$2pH = pK_W$ $\quad$ $2\,pOH = pK_W$

$pH = \dfrac{pK_W}{2}$ $\quad$ $pOH = \dfrac{pK_W}{2}$

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

At 25°C, $K_W = 10^{-14}$ or $pK_W = 14$

pH + pOH = 14 or pH = pOH = 7

(i) Main Points :-

$pH = pOH = \dfrac{pK_W}{2}$

At 25°C, pK_W = pH + pOH =14

(j) Effect of temperature: -

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

On increasing temperature rate of the forward reaction increases i.e.α 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, $K_W = (100 \times 10^{-14}) = 10^{-12}$

Parameter	At 25°C	At 90°C
K_W	10^{-14}	10^{-12}
pK_w	14	12
$pH = \dfrac{pK_w}{2} = pOH$	7	6
$[H^+] = [OH^-]$	10^{-7}	10^{-6}
pH + pOH	14	12

Important point:

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

Mixing of ions:

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

Common ion:

On mixing CH_3COONa with CH_3COOH solution

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

$$\begin{array}{lll} \text{On mixing} & CH_3COONa \rightleftharpoons \boxed{CH_3COO^-} + Na^+ & CH_3COO^- \text{ is common ion} \\ & CH_3COOH \rightleftharpoons \boxed{CH_3COO^-} + H^+ & \end{array}$$

Odd ion:

On mixing, NaOH with CH_3COOH solution

$$\begin{array}{lllll} \text{Initially} & CH_3COOH & \rightleftharpoons & CH_3COO^- + H^+ & \\ \text{On mixing NaOH} & NaOH & \rightleftharpoons & Na^+ \quad + & \boxed{\begin{array}{l} OH^- \\ \\ H^+ \end{array}} \rightarrow H_2O \\ & CH_3COOH & \rightleftharpoons & CH_3COO^- + & \end{array}$$

- **pH calculation of different type of solution**

(a) Strong acid Solution :

(i) If concentration is greater than 10^{-6} M, in this case H^+ ions coming from water can be neglected.
So $[H^+]$ = normality of strong acid Solution

(ii) If concentration is less than 10^{-6} M, in this case H^+ ions coming from water cannot be neglected.
So $[H^+]$ = normality of strong acid + H^+ ions coming from water in presence of this strong acid

Ex-2 Calculate pH of 10^{-8}M HCl Solution.

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

$\qquad\qquad 10^{-8} + x$

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

$10^{-14} = x(x \times 10^{-8})$

$\Rightarrow x^2 + x \times 10^{-8} - 10^{-14} = 0$

$$x = \frac{-10^{-8} \pm \sqrt{10^{-16} + 4 \quad x \quad 10^{-14}}}{2} = \frac{-10^{-8} + 10^{-7}\sqrt{4 + \frac{1}{100}}}{2} = \frac{(\sqrt{401} - 1)10^{-8}}{2} = 0.95 \text{ x } 10^{-7}$$

$[H^+] = 10.5 \text{ x } 10^{-8} = 1.05 \text{ x } 10^{-7}$

$[pH] = 7 - \log 1.05 \approx 6.98$

10^{-9} M HCl pH ≈ 7

10^{-16} M HCl pH ≈ 7

(b) Strong base Solution :

Calculate the $[OH^-]$ which will be equal to normality of the strong base solution and then use $K_W = [H^+] \times [OH^-] = 10^{-14}$, to calculate $[H^+]$.

Ex-3 Calculate pH of 10^{-7} M of NaOH Solution

Sol. $[OH^-]$ from NaOH = 10^{-7}

$[OH^-]$ from water = $x < 10^{-7}$ M (due to common ion effect)

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

$$- \qquad (x + 10^{-7}) \qquad x$$

$$K_w = [H^+][OH^-] = 10^{-14} = x(x + 10^{-7})$$

$$x^2 + 10^{-7}x - 10^{-14} = 0$$

$$\Rightarrow \quad x = \frac{\sqrt{5}-1}{2} \times 10^{-7} = 0.618 \times 10^{-7} \qquad (\sqrt{5} = 2.236)$$

$$[OH^-] = 10^{-7} + 0.618 \times 10^{-7} = 1.618 \times 10^{-7}$$

pOH = 7 – log (1.618) = 6.79

pH = 14 – 6.79 = 7.21

(c) pH of mixture of two strong acids :

If V_1 volume of a strong acid solution of normality N_1 is mixed with V_2 volume of another strong acid solution of normality N_2, then

Number of H^+ ions from I-solution = N_1V_1

Number of H^+ ions from II-solution = N_2V_2

If final normality is N and final volume is V, then

$$NV = N_1V_1 + N_2V_2$$

[dissociation equilibrium of none of these acids will be disturbed as both are strong acid]

$$[H^+] = N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$$

(d) pH of mixture of two strong bases :

Similar to above calculation

$$[OH^-] = N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2} \quad [H^+] = \frac{10^{-14}}{[OH^-]}$$

Ex-4 Calculate pH of mixture of (400ml , $\frac{1}{200}$ M H_2SO_4) + (400ml, $\frac{1}{100}$ M HCl) +(200 ml of water)

Sol. $N_1V_1 = \frac{1}{100} \times \frac{400}{1000} = \frac{4}{1000}$, $N_2V_2 = \frac{4}{1000}$, H^+ ions from water will be neglected

$N_1V_1 + N_2V_2 = 8 \times 10^{-3}$ $\qquad [H^+] = \frac{8 \times 10^{-3}}{1} = 8 \times 10^{-3}$

pH = 3 – log 8 =2.1

(e) pH of mixture of a strong acid and a strong base :

- Acid Base neutralisation reaction will take place.
- The solution will be acidic or basic depending on which component has been taken in excess.
- If V_1 volume of a strong acid solution of normality N_1 is mixed with V_2 volume of a strong base solution of normality N_2 , then

Number of H^+ ions from I-solution = N_1V_1

Number of OH^- ions from II-solution = N_2V_2

If $N_1V_1 > N_2V_2$

$$[H^{\cdot}] = N = \frac{N_1V_1 - N_2V_2}{V_1 + V_2}$$

solution will be acidic in nature

If $N_2V_2 > N_1V_1$

$$[OH^{\cdot}] = N = \frac{N_2V_2 - N_1V_1}{V_1 + V_2}$$

solution will be basic in nature

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

Ex-5 Calculate pH of mixture of (400ml, $\frac{1}{200}$ M $Ba(OH)_2$) + (400ml, $\frac{1}{50}$ M HCl) + (200ml of Water)

Sol. $[H^+] = \dfrac{400\times\frac{1}{50} - 400\times\frac{1}{200}\times 2}{1000} = 4\times10^{-3}$, so pH = 3 – 2 log 2 = 2.4

Weak Acid, Weak Base and their pH, Polyprotic acid :

(f) pH of a weak acid (monoprotic) Solution :

- Weak acid does not dissociate 100% therefore we have to calculate the percentage dissociation using K_a, dissociation constant of the acid.
- We have to use Ostwald's Dilution law (as have been derived earlier)

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

t = 0 C 0 0

$t = t_{eq}$ $C(1-\alpha)$ $C\alpha$ $C\alpha$ $K_a = \dfrac{[H^+]\ [OH^-]}{[HA]} = \dfrac{C\ \alpha^2}{1-\alpha}$

If $\alpha << 1 \Rightarrow (1-\alpha) \approx 1$ $\Rightarrow$ $K_a \approx C\alpha^2$ $\Rightarrow$ $\alpha = \sqrt{\dfrac{K_a}{C}}$ (is valid if α < 0.1 or 10%)

$[H^+] = C\alpha = C\sqrt{\dfrac{K_a}{C}} = \sqrt{K_a \times C}$ So **pH =** $\frac{1}{2}(pK_a - \log C)$

on increasing the dilution $\Rightarrow C\downarrow \Rightarrow \alpha\uparrow$ and $[H^+]\downarrow \Rightarrow pH\uparrow$

Ex-6 Calculate pH of 10^{-1} M CH_3COOH

Take $K_a = 2\times10^{-5}$.

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

C 0 0

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

$$K_a = \frac{C\alpha^2}{1-\alpha} \Rightarrow \alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2\times10^{-5}}{10^{-1}}} = \sqrt{2\times10^{-4}} \quad (\alpha << 0.1)$$

So, $[H^+] = 10^{-1}\times\sqrt{2}\times10^{-2} \Rightarrow pH = 3 - \frac{1}{2}\log 2 = 2.85$ **Ans.**

(g) pH of a mixture of weak acid(monoprotic) and a strong acid Solution :

- Weak acid and Strong acid both will contribute H^+ ion.
- For the first approximation we can neglect the H^+ ions coming from the weak acid Solution and calculate the pH of the Solution from the concentration of the strong acid only.

- To calculate exact pH, we have to take the effect of presence of strong acid on the dissociation equilibrium of the weak acid.

  If [SA] = C_1 and [WA] = C_2, then $[H^+]$ from SA = C_1

  the weak acid will dissociate as follows.

$$\begin{array}{lll} HA & \rightleftharpoons H^+ & + \quad A^- \\ C_2 & 0 & 0 \\ C_2(1-\alpha) & C_2\alpha + C_1 & C_2\alpha \end{array} \qquad K_a = \frac{(C_2\alpha + C_1)\,C_2\alpha}{C_2(1-\alpha)} \quad (\alpha <<< 1)$$

  (The weak acids dissociation will be further suppressed because of presence of strong acid, common ion effect)

$$K_a = (C_2\alpha + C_1)\alpha$$

  Total H^+ ion concentration = $C_1 + C_2\alpha$

- If the total $[H^+]$ from the acid is less than 10^{-6} M, then contribution from the water can be neglected at 25ºC temp., if not then we have to take $[H^+]$ from the water also.

(h) pH of a Solution **of a polyprotic weak acid :**

- Diprotic acid is the one, which is capable of giving 2 protons per molecule in water.

  Let us take a weak diprotic acid (H_2A) in water whose concentration is cM.

  In an aqueous Solution , following equilbria exist.

If

α_1 = degree of ionization of H_2A in presence of HA^- $\qquad K_{a_1}$ = first ionisation constant of H_2A.

α_2 = degree of ionisation of HA^- in presence of H_2A. $\qquad K_{a_2}$ = second ionisation constant of H_2A.

I step

$$\begin{array}{llll} & H_2A + H_2O & \rightleftharpoons HA^- & + \quad H_3O^+ \\ \text{at eq.} & c(1-\alpha_1) & c\alpha_1(1-\alpha_2) & (c\alpha_1 + c\alpha_1\alpha_2) \end{array}$$

$$(K_{eq})_1\,(H_2O) = \frac{[H_3O^+][HA^-]}{[H_2A]} = K_{a_1}$$

$$\therefore\ K_{a_1} = \frac{(c\alpha_1 + c\alpha_1\alpha_2)[c\alpha_1(1-\alpha_2)]}{c(1-\alpha_1)}$$

$$= \frac{[c\alpha_1(1+\alpha_2)][\alpha_1(1-\alpha_2)]}{1-\alpha_1} \qquad \text{........(i)}$$

II step

$$\begin{array}{llll} & HA^- + H_2O & \rightleftharpoons A^{2-} & + \quad H_3O^+ \\ \text{at eq.} & c\alpha_1(1-\alpha_2) & c\alpha_1\alpha_2 & (c\alpha_1 + c\alpha_1\alpha_2) \end{array}$$

$$[K_{eq}]_2\,[H_2O] = \frac{[H_3O^+][A^{2-}]}{[HA^-]} = K_{a_2}$$

$$K_{a_2} = \frac{(c\alpha_1 + c\alpha_1\alpha_2)(c\alpha_1 c\alpha_2)}{c\alpha_1\ (1-\alpha_2)}$$

$$= \frac{[c\alpha_1(1+\alpha_2)]\ \alpha_2}{1-\alpha_2} \qquad \text{..........(ii)}$$

Knowing the values of K_{a_1}, K_{a_2} and c, the values of α_1 and α_2 can be calculated using equations (i) and (ii).

After getting the values of α_1 and α_2 $[H_3O^+]$ can be calculated as.

$[H_3O^+]_T = c\alpha_1 + c\alpha_1\alpha_2$

Finally, for calculating pH

- If the total $[H_3O^+] < 10^{-6}$ M, the contribution of H_3O^+ from water should be added
- If the total $[H_3O^+] > 10^{-6}$ M, then $[H_3O^+]$ contribution from water can be ignored.

Using this $[H_3O^+]$, pH of the Solution can be calculated.

Approximation

For diprotic acids, $K_{a_2} << K_{a_1}$ and α_2 would be even smaller than α_1 .

$\therefore \quad 1 - \alpha_2 \approx 1$ and $1 + \alpha_2 \approx 1$

Thus, equation (i) can be reduced to $\quad K_{a_1} = \dfrac{C\alpha_1 \times \alpha_1}{1-\alpha_1}$

This is an expression similar to the expression for a weak monoprotic acid.

- Hence, for a diprotic acid (or a polyprotic acid) the $[H_3O^+]$ can be calculated from its first equilibrium constant expression alone provided . $K_{a_2} << K_{a_1}$

Ex:7 Calculate pH, $[HS^-]$, S^{2-} , $[Cl^-]$ in a Solution which is 0.1 M HCl & 0.1 M H_2S given that $Ka_1 (H_2S) = 10^{-7}$, $Ka_2 (H_2S) = 10^{-14}$ also calculate α_1 & α_2

Sol. HCl + H_2S

0.1 0.1 C = 0.1

$\therefore$ pH = 1 (most of $[H^+]$ comes from HCl)

$$H_2S \rightleftharpoons H^+ + SH^-$$

C_1 +0.1

$C_1 - C_1\alpha_1$ $C_1\alpha_1$, $C_1\alpha_1$

$$Ka_1 = \frac{C_1\alpha_1 \times 10^{-1}}{C_1(1-\alpha_1)}$$

$$\alpha_1 = \frac{10^{-7}}{10^{-1}} = 10^{-6}$$

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

$C_1\alpha_1(1 - \alpha_2)$ $C_1\alpha_1\alpha_2$ 0.1

$10^{-14} = 0.1 \times \alpha_2$

$\Rightarrow \quad \alpha_2 = 10^{-13}$

$[S_2^{2-}] = C\alpha_1\ \alpha_2 = 10^{-1} \times 10^{-6} \times 10^{-13} = 10^{-20}$ M

Ex.8 Calculate pH 10^{-1} M HCl in 10^{-3} M CH_3COOH $[K_a = 2 \times 10^{-5}]$

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

$C(1 - \alpha)C\alpha$ $10^{-1} + C\alpha$

; C ; 10^{-1}

H^+ ion can be treated completely from HCl due to less dissociation of CH_3COOH and its low conc.

$$2 \times 10^{-5} = \frac{C\alpha \times 10^{-1}}{C}$$

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

$[H^+]_{CH_3COOH}\ C\alpha = 2 \times 10^{-7}$

(J) Mixture of a polyprotic weak Acid & Strong Acid

(i) pH can be calculated by taking the concentration of strong acid only

(ii) For other calculations we should consider the dissociation equilibria of weak polyprotic acid as done in equation.

Ex.9 At dissociation constant of heavy water is 4 × 10^{-15} at 35°c Its density is 1.04 g/ml calculate its ionic product & degree of dissociation.

Sol. kw = kd[$D_2$0] = $\left(4\times10^{-15}\times\frac{1040}{20}\right)$ $\quad d=\sqrt{\frac{kw}{c}}=\sqrt{\frac{2.08\times10^{-13}}{5^2}}=12.64\times10^{-8}$

Ex.10 Calculate ionic product of H2O at 50°c

Sol. ΔH = 13.7 × 10^3 cal

$$\log=\frac{k_2}{10^{-14}}=\frac{13.7\times10^3}{2}\left(\frac{1}{290}-\frac{1}{323}\right)$$

Ex.11 The hydronium ion conc. in an aq. H_2Co_3 solution is 4 × 10^{-4}m at 25°c OH^- ion the sol is.

(1) 0 (2) 2.5 × 10^{-10} (3) 2.5 × 10^{-3} (4) 2.5 × 10^{-11} m

Sol. 4

Ex.12 Select correct option from the following ?

(1) pk_w increase with increase of temperature (2) pk_w decrcase with increase of temperature

(3) pk_w = 14 at all temperature (4) pk_w = pH at all temperature

Sol 2

PRACTICE SECTION-01

Q.1 Ionisation constant of CH_3COOH is 1.7 × 10^{-5} and concentration of H^+ ions is 3.4 × 10^{-4}. Then initial concentration of CH_3COOH molecules is

(1) 3.4 × 10^{-4} (2) 3.4 × 10^{-3} (3) 6.8 × 10^{-3} (4) 1.7 × 10^{-3}

Q.2 At 25°C, the dissociation constant of a base, BOH is 1.0 × 10^{-12}. The concentration of hydroxyl ions in 0.01 M aqueous solution of the base would be :

(1) 1.0 × 10^{-6} mol L^{-1} (2) 1.0 × 10^{-7} mol L^{-1} (3) 2.0 × 10^{-6} mol L^{-1} (4) 1.0 × 10^{-5} mol L^{-1}

Q.3 Which of the following concentrations has the largest degree of dissociation for a weak acid?

(1) 1.0 M (2) 0.5 M (3) 0.10 M (4) 0.01 M

Q.4 Calculate the amount of acetic acid in one litre of its solution α = 1% and K_1 = 1.8 × 10^{-5}.

(1) 0.18 gm (2) 10.8 gm (3) 108 gm (4) 18 gm

Q.5 The K_w of $2H_2O \rightleftharpoons H_3O^+ + OH^-$ changes from 10^{-14} at 25°C to 9.62 × 10^{-14} at 90°C. Choose correct option.

(1) pH of H_2O at 90°C is 6.51 and water become acidic

(2) pH of H_2O at 90°C is 6.51 and water remains neutral

(3) pH and neutrality of water both remains unchanged

(4) pH of H_2O at 90°C is 6.1 and water remains neutral

Q.6 The K_a values of acetic acid, hypochlorous acid and formic acid are 1.74 × 10^{-5}, 3.0 × 10^{-8} and 1.8 × 10^{-4} respectively. Which of the following orders of pH of 0.1 mol dm^{-3} solutions of these acids is correct?

(1) Acetic acid > Hypochlorous acid > Formic acid (2) Hypochlorous acid > Acetic acid > Formic acid

(3) Formic acid > Hypochlorous acid > Acetic acid (4) Formic acid > Acetic acid > Hypochlorous acid

Q.7 How much water should be added to 10.0 g of acetic acid to give a hydrogen-ion concentration equal to 1.0 × 10^{-3} M (given pK_a = 4.74)?

(1) 4 L (2) 6 L (3) 5 L (4) 3 L

ANSWER KEY

Que.	1	2	3	4	5	6	7
Ans.	3	2	4	2	2	2	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.

Acid + Base → Salt + Water

$HCl + NaOH \rightarrow NaCl\ (Salt) + H_2O \qquad \Delta H = -ve$

(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 ions or a hydroxyl (OH^-) group.

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

Types of general salts:

(i) Strong Acid Strong Base	(ii) Strong Acid Weak Base
(iii) Weak Acid Strong Base	(iv) Weak Acid Weak Base

(b) Acid salts :– Salt 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.

Ex. : $FeSO_4.(NH_4)_2SO_4.6H_2O$, (Ferrous ammonium sulphate) $K_2SO_4 . Al_2(SO_4)_3 . 24H_2O$ etc. (Potash alum)

(e) Complex salts:– These are formed by combination of simple salts or molecular compounds or complex ions are present in salt molecules. These are stable in solid state as well as in solutions.

$$\underbrace{FeSO_4 + 6KCN}_{\text{Simple salt}} \longrightarrow \underset{\text{Complex salt}}{K_4[Fe(CN)_6]} + K_2SO_4$$

$$K_4[Fe(CN)_6] \longrightarrow 4K^+ + [Fe(CN)_6]^{-4}$$

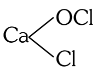
$$\underset{\text{Simple salt}}{CoSO_4} + \underset{\text{Molecular compound}}{6NH_3} \longrightarrow \underset{\text{Complex salt}}{[Co(NH_3)_6]SO_4}$$

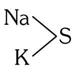
$$[Co(NH_3)_6]SO_4 \longrightarrow [Co(NH_3)_6]^{+2} + SO_4^{-2}$$

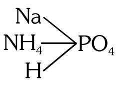
(f) Mixed salts :– The salts which furnishes more than one cation or more than one anion when dissolved in water is called a mixed salt.

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) 

(iii) 

(iv) 

(C) CONJUGATE ACID-BASE PAIR: Conjugate acid-base pairs means difference between two species of only one H^+ or OH^- ion.

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

Weak acid — Conjugate strong base

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

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

Weak base — Conjugate strong acid

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

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

Strong acid — Conjugate weak base (spectator ion)

($K_a \simeq \infty$) — Accepting tendency of $H^+ \simeq 0$

$$NaOH \rightleftharpoons Na^+ + OH^-$$

Strong base — Conjugate weak acid

($K_b \simeq \infty$) — (spectator ion)

(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 — CH_3COO^-

acid — conjugate base

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $\qquad$ $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} \quad \text{...... (i)} \qquad K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} \quad \text{.... (ii)}$$

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 $\boxed{pK_a + pK_b = 14}$

Note : These relation are applicable for only conjugate acid-base pairs.

HYDROLYSIS OF SALTS

Salt hydrolysis is defined as the process in which water reacts with salt to form acid and base. i.e. change in concentration of H^+ and OH^- ions of water.

Water + Salt $\rightleftharpoons$ Acid + Base

(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.

$NaCl + H_2O \rightleftharpoons NaOH + HCl$

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

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

(i) Hydrolysis of salt of [SA – SB] is not possible.

(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.

(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 $[K_h]$

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

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

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

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

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

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

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

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

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

(b) Degree of hydrolysis – Represented by h

	$NH_4^+ + H_2O \rightleftharpoons$	NH_4OH +	H^+
Initial concentration of salt	C	0	0
	C – Ch	Ch	Ch

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

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

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

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

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

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

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

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

taking – log on both sides

$$-\log [H^+] = -\log \sqrt{\frac{K_W \times C}{K_b}} \Rightarrow pH = -\log \left(\frac{K_W \times C}{K_b}\right)^{1/2}$$

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

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

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

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

(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.

$$CH_3COONa + H_2O \rightleftharpoons CH_3COOH + NaOH$$

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

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

(i) In this type of salt hydrolysis, anion reacts with water therefore called as anionic hydrolysis.

(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 $(1 - h) \approx 1$

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

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

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

(c) pH of the solution

$[OH^-] = Ch$

$[OH^-] = C\sqrt{\frac{K_W}{K_a \times C}}$ × or $[OH^-] = \sqrt{\frac{K_W \times C}{K_a}}$

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

$\therefore$ $\left[H^+\right] = \frac{K_W}{\left[OH^-\right]} = \frac{K_W}{\sqrt{\frac{K_W \times C}{K_a}}}$

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

taking – log on both sides

$-\log [H^+] = -\log \sqrt{\frac{K_W \times K_a}{C}}$

$pH = -\log \left(\frac{K_W \times K_a}{C}\right)^{1/2}$

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

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

$pH = \frac{1}{2} pK_W + \frac{1}{2} pKa + \frac{1}{2} \log C$

$pH = 7 + \frac{1}{2} pKa + \frac{1}{2} \log C$ ....(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.

$CH_3COONH_4 + H_2O \rightleftharpoons CH_3COOH + NH_4OH$

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

Solution is almost neutral but it may be acidic or basic depending upon the nature of acid and base and pH of the solution is near to 7.

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

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

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

For weak base

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

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

For weak acid

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

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

For water

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

$$K_W = [H^+][OH^-] \qquad \text{........(4)}$$

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

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

$K_h \times K_b \times K_a = K_W$

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

(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 $(1 - h) \approx 1$

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

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

(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)]

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

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

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

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

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

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

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

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

Note : Degree of hydrolysis of weak acid and weak base [WA – WB] type salt does not depend on the concentration of salt.

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

Summary :

Type of salts	$K_h = \frac{K_W}{\text{weak}}$ **At 25°C**	$h = \sqrt{\frac{K_h}{C}}$	**[H$^+$]**	**pH**
SA SB	N.A.	N.A.	10^{-7}	7
WA SB	$K_h = \frac{K_W}{K_a}$	$h = \sqrt{\frac{K_W}{K_a \times C}}$	$\sqrt{\frac{K_W \times K_a}{C}}$	$7 + \frac{1}{2} pKa + \frac{1}{2} \log C$
SA WB	$K_h = \frac{K_W}{K_b}$	$h = \sqrt{\frac{K_W}{K_b \times C}}$	$\sqrt{\frac{K_W \times C}{K_b}}$	$7 - \frac{1}{2} pK_b - \frac{1}{2} \log C$
WA WB	$K_h = \frac{K_W}{K_a \times K_b}$	$h = \sqrt{\frac{K_W}{K_a \times K_b}}$	$\sqrt{\frac{K_W \times K_a}{K_b}}$	$7 + \frac{1}{2} pKa - \frac{1}{2} pK_b$

If acid base are polybasic/polyacidic, then some more cases could arise.

(i) Solution containing polyvalent anions (or cations) or weak polyprotic acids (or bases)

(ii) The hydrolysis of these species with fake place in steps (just like dissociation of weak polyprotic acid)

(iii) Out of different steps, generally first step hydrolysis dominates mainly because of two reason.

(a) The hydrolysis constant of second & farther steps in generally negligible in comparison to first step hydrolysis constant

(b) The second & further step hydrolysis will be suppressed in presence of ions produced due to first step hydrolysis (common ion effect)

For polyprotic acids like (H_2S, H_3PO_4, H_2CO_3, H_2CO_3, $H_2C_2O_4$)

We already know that the dissociation always takeplace in steps eg. H_3PO_4

$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4$– $\qquad Ka_1 = \dfrac{[H^+][H_2PO_4^-]}{[H_3PO_4]}$ …… (1)

$H_2PO_4 \rightleftharpoons H^+ + HPO_4^{2-}$ $\qquad Ka_2 = \dfrac{[H^+][HPO_4^{2-}]}{[H_2PO_4]}$ …… (2)

$HPO_4^{2-} \rightleftharpoons H^+ + HPO_4^{3-}$ $\qquad Ka_3 = \dfrac{[H^+][HPO_4^{3-}]}{[HPO_4^{2-}]}$ …… (3)

For all acid we always have $Ka_1 >> Ka_2 >> Ka_3$

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

$C_1\alpha_1(1 - \alpha_2) \qquad C_1\alpha_1\alpha_2 \approx 0.1$

$10^{-14} = \dfrac{C_1\alpha_1\alpha_1 \times 0.1}{C_1\alpha_1(1-\alpha_2)}$

Degree of dissociation of $HS^- = \alpha_2 = 10^{-13}$

$[S^{2-}] = C_1\alpha_2\alpha_2 = 10^{-1} \times 10^{-6} \times 10^{-13} = 10^{-20}$ M

$[HS^-] \approx C_1\alpha_1 = 10^{-1} \times 10^{-6} = 10^{-7}$ M

H^+ ion can be calculate from 1st step only because (H^+) from 2nd & 3rd step can be neglected as

(a) $Ka_1 >> ka_2 >> Ka_3$

(b) $[H^+]$ from 1st dissociation will suppress the dissociation oof 2nd & 3rd step

Now from the hydrolysis of polyvalent ions (from salt like K_3PO_4, Na_2CO_3, $ZnSO_4$, $FeCl_3$,$(NH_4)_2$ C_2O_4 or Ions like PO_4^{3-}, CO_3^{2-}, Zn^{+2}, fe^{+3} etc)

Derivation	PO_4^{3-} +	H_2O	$\rightleftharpoons$	HPO_4^{2-}	+	PH^-
	t = 0	C		0		0
	t = eq	C(1 – h)		Ch		Ch

$Kh_1 = \dfrac{[OH^-][HPO_4^{2-}]}{PO_4^{3-}}$ ….. (4)

$HPO_4^{2-} + H_2O \rightleftharpoons H_2PO_4^{2-} + OH^-$

$Kh_2 = \dfrac{[OH^-][HPO_4^-]}{[HPO_4^{2-}]}$ ….. (5)

$H_2PO_4^- + H_2O \rightleftharpoons H_3PO_4 + OH^-$

$Kh_3 = \dfrac{[OH^-][H_3PO_4]}{[H_2PO_4^-]}$ ….. (6)

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

$Kw = [H^+][OH^-]$ ….. (7)

From above equations, we get

$Ka_1 \times Kh_3 = lw$; $Ka_2 \times Kh_2 = kw$; $Ka_3 \times Kh_1 = Kw$

Numerically $Kh_1 >> Kh_2 >> Kh_3$

Generally, P^H is calculated only using the first step hydrolysis

$$Kh_1 = \frac{Ch - Ch}{C(1-h)} = \frac{Ch^2}{1-h}$$

$$1 - h \approx 1$$

$$Kh_1 = Ch_2 \;\&\; h = \sqrt{\frac{Kh_1}{C}}$$

$$[OH^-] = Ch = \sqrt{Kh_1} \times C$$

$$[H^+] = \frac{Kw}{[OH^-]} = Kw\sqrt{\frac{Ka_3}{Kw \times C}} = \sqrt{\frac{Kw \times Ka_3}{C}}$$

$$\text{So, pH} = \frac{1}{2}\,[pK_w + pKa_3 + \log c]$$

If $h \geq 0.1$ we should solve quadratic equation

Solution containg amphiprotic anion (or cation)

Anion of $NaHCO_3$, NaHS etc can undergo ionization to from H^+ & can also undergo hydrolysis to form OH^-

(1) $HCO_3^- + H_2O \xrightleftharpoons{\text{ionisation}} Co_3^{2-} + H_3O^+ ; Ka_2$

(2) $HCO_3^- + H_2O \xrightleftharpoons{\text{hydrolysis}} H_2CO_3 + OH^- ; \frac{Kw}{Ka}$

Taking the assumption :- Degree of ionization (α) = Degree of hydrolysis (h) or $[CO_3^{2-}] = [H_2CO_3]$

Calculation of pH can be done by relation :-

$pH(HCO_3^-) = \left(\frac{pKa_1 + pKa_2}{2}\right)$ = Average of PKa values of parent acid involved in ionization & hydrolysis reaction

Similarly for $H_2PO_4^-$ & HPO_4^{2-} amphiprotic anions

$pH(H_2PO_4^-) = \left(\frac{pKa_1 + pKa_2}{2}\right)$ and

$pH(HPO_4^{2-}) = \left(\frac{pKa_2 + pKa_3}{2}\right)$

Where Ka_1, Ka_2 & Ka_3 are dissociation constant of H_3PO_4.

Ex.13 How many grams of NH_4Cl should be dissolved per litre of solution to have a pH of 5.13 ?

Sol. NH_4Cl is salt of SA + WB

$$pH = \frac{1}{2}[pK_w - \log c - pK_b]$$

$10.26 = 14 - \log C - 4.74$

$C = 10^{-1}$ M

$[NH_4Cl] = 10^{-1}$ M

$W_{NH4Cl} = 10^{-1} \times 53.5 \text{ g } L^{-1} = 5.35 \text{ g } L^{-1}$

Ex.14 Calculate the pH to degree of hydrolysis of 0.01 M solution of NaCN, Ka for HCN is 6.2 × 10^{-12}

Sol. NaCN = Salt of SB + w.A

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

$$k_h = \frac{[HCN][OH]}{[CN^-]} = \frac{kw}{Ka} = \frac{10^{-14}}{6.2\times10^{-12}}$$

Let x mole of salt undergo hydrolysis

$[CN^-] = 0.01 - x \approx 0.01$

$[HCN] = x$

$[OH^-] = x$

$$k_h = \frac{x\cdot x}{0.01} = 1.6\times10^{-3}$$

$x^2 = 1.6 \times 10^{-5} \Rightarrow x = 4 \times 10^{-3}$

$[OH^-] = x = 4 \times 10^{-3}$

$$[H_3O^+] = \frac{kw}{[OH^-]} = \frac{10^{-14}}{4\times10^{-3}} = 0.25\times10^{-11}$$

pH = – log (0.25 × 10^{-11}) = 11.602

$$\text{Degree of hydrolysis} = \frac{x}{0.01} = \frac{4\times10^{-3}}{0.01} = 4\times10^{-1}$$

PRACTICE SECTION-02

Q.1 Given, $HF + H_2O \xrightleftharpoons{K_a} H_3O^+ + F^-$

$F^- + H_2O \xrightleftharpoons{K_b} HF + OH^-$

Which relation is correct?

(1) $K_b = K_w$ (2) $K_b = \frac{1}{K_w}$ (3) $K_a \times K_b = K_w$ (4) $\frac{K_a}{K_b} = K_w$

Q.2 The correct statements are:

(1) CH_3COOH is a weak acid

(2) NH_4Cl gives an alkaline solution in water

(3) CH_3COONa gives an acidic solution in water

(4) Na_2SO_4 gives basic solution in water

Q.3 Calculate pH of 0.10 M KCN solution at 25°C. For HCN, $K_a = 6.2 \times 10^{-10}$.

(1) 3.9 (2) 11.1 (3) 7 (4) None

Q.4 Calculate the pH of 0.10 M solution of NH_4Cl. The dissociation constant (K_b) of NH_3 is 1.6×10^{-5}.

(1) 5.1 (2) 8.9 (3) 7 (4) 3

Q.5 A weak acid HA ($k_a = 10^{-5}$) combined with strong base to produce salt AB. Find out the equilibrium constant for this salt formation.

(1) 10^{-9} (2) 10^{9} (3) 10^{-5} (4) 10^{5}

Q.6 The pKa of acetic acid and pK_b of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the hydrolysis constant of ammonium acetate at 298 K

(1) 1.25×10^{-5} (2) 3.25×10^{-5} (3) 6×10^{-4} (4) 2×10^{-4}

Q.7 Which of the following salt represent maximum anionic water hydrolysis

(1) kCl (2) CH_3COONa (3) HCOONa (4) PhONa

ANSWER KEY

Que.	1	2	3	4	5	6	7
Ans.	3	1	2	1	2	2	4

SOLUBILITY AND SOLUBILITY PRODUCT (Ksp)

(A) SOLUBILITY

(a) Definition

(i) 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.

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

(ii) The maximum weight of solute which can be dissolved in a given amount of solvent is known as solubility.

(b) Important 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(Ksp) :

(a) 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 further dissolution of even a small amount of salt is not possible).

Ex. (i) Let solubility of AgCl	$AgCl \rightleftharpoons$	Ag^+	$+ Cl^-$
is S mol L^{-1}	a	0	0
After saturation	a – S	S	S

According to L.O.M.A.

Solubility product at saturation in terms of concentration of ions $K_{sp} = [Ag^+] [Cl^-]$

Solubility product in terms of solubility $K_{sp} = (S) (S)$

$$\boxed{K_{sp} = S^2}$$

Ex. (ii) Ksp for $CaCl_2$	$CaCl_2 \rightleftharpoons$	Ca^{+2}	$+ 2Cl^-$
Initially	a	0	0
Let solubility of $CaCl_2$ is S mol L^{-1}	a – S	S	2S

Solubility product in terms of concentration of ions Ksp = $[Ca^{+2}] [Cl^-]^2$

In terms of solubility $K_{sp} = (S) (2S)^2$

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

Ex. (iii) Ksp for $AlCl_3$	$AlCl_3 \rightleftharpoons$	Al^{+3}	$+ 3Cl^-$
Initially	a	0	0
Let solubility of $AlCl_3$ is S mol L^{-1}	a – S	S	3S

Solubility product in terms of concentration of ions $K_{sp} = [Al^{+3}] [Cl^-]^3$

In terms of solubility $K_{sp} = (S) (3S)^3$

$$\boxed{K_{sp} = 27S^4}$$

(b) General form $AxBy \rightleftharpoons xA^{+y} + yB^{-x}$

$a \qquad 0 \qquad 0$

$a - S \qquad xS \qquad yS$

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

$= [xs]^x \times [ys]^y = x^x.s^x.\ y^y.s^y$

$$\boxed{K_{sp} = x^x\,y^y\,S^{(x+y)}}$$

Ex. (i) $Al_2(SO_4)_3$

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

(ii) Na_2KPO_4

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

(iii) $NaKRbPO_4$

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

APPLICATION OF SOLUBILITY PRODUCT (Ksp)

(A) TO FIND OUT THE SOLUBILITY (S) :

(i) Ksp 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.

$AB(s) \rightleftharpoons A^+ + B^-$

$1 \qquad 0 \quad 0$

$1 \qquad s \quad s$

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

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

(ii) Ksp 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.

$AB_2(s) \rightleftharpoons A^{+2} + 2B^-$

$1 \qquad 0 \quad 0$

$1 \qquad s \quad 2s$

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

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

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

(iii) Ksp of AB_3 or A_3B (Mono-tri or tri-mono valency) type salt –

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

$AB_3(s) \rightleftharpoons A^{+3} + 3B^-$

$1 \qquad 0 \qquad 0$

$1 \qquad s \qquad 3s$

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

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

(iv) Ksp 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.

$$A_2B_3(s) \rightleftharpoons 2A^{+3} + 3B^{-2}$$

1	0	0
1	2s	3s

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

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

(B) COMMON ION EFFECT ON SOLUBILITY:

Important point:- Solubility of substances always decreases in the presence of common ion.

Ex. (i) Find out the solubility of AgCl in the presence of C NaCl solution?

$$AgCl \rightleftharpoons Ag^+ + Cl^-$$

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

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

$K_{sp} = S^2$

In NaCl solution

$$NaCl \longrightarrow Na^+ + Cl^-$$

C C C $[Cl^-] = C$

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

$$AgCl \rightleftharpoons Ag^+ + Cl^-$$

S' S' S'+C

According to L.O.M.A.

$K_{sp} = [Ag^+]'[Cl^-]'$

$K_{sp} = S'(S' + C) = S'^2 + S'C$ (Neglecting the higher power terms of S')

$K_{sp} = S'C$

$$\boxed{S' = \frac{K_{sp}}{C}}$$

Ex. (ii) Find out the solubility of $CaCl_2$ solution in the presence of C NaCl solution?

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

S S 2S [Let solubility of $CaCl_2$ is S mol L^{-1}]

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

$= 4S^3$

For NaCl solution

$$NaCl \longrightarrow Na^+ + Cl^-$$

C C C $[Cl^-] = C$

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

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

S' S' 2S'+C

According to L.O.M.A.

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

$K_{sp} = S'(2S' + C)^2 = S'(4S'^2 + 4S'C + C^2)$

$K_{sp} = 4S'^3 + 4S'^2C + S'C^2$ (Neglecting the higher power terms of S')

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

Ex. (iii) Find out the solubility of NaCl in the presence of C $CaCl_2$ solution?

$$\begin{array}{lcccc} CaCl_2 & \longrightarrow & Ca^{+2} & + & 2Cl^- \\ C & & C & & 2C \end{array}$$

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

$$\begin{array}{lcccc} NaCl & \rightleftharpoons & Na^+ & + & Cl^- \\ S' & & S' & & S'+2C \end{array}$$

According to L.O.M.A.

$K_{sp} = [Na^+]^1\,[Cl^-]^1$

$K_{sp} = S'\ (S' + 2C) = S'^2 + 2S'C$ (Neglecting the higher power terms of S')

$$\boxed{S' = \frac{K_{sp}}{2C}}$$

(C) GROUP PRECIPITATION :

(i) $Q_{sp} < K_{sp}$ ⇒ Unsaturated } No precipitation

(ii) $Q_{sp} = K_{sp}$ ⇒ Saturated }

(iii) $Q_{sp} > K_{sp}$ ⇒ Super saturated ⇒ precipitation

Group	Group reagent	Basic radical	Composition and colour of precipitate
Zero	NaOH or $Ca(OH)_2$, heat if required	NH_4^+	Ammonia gas is evolved
1	Dil HCl	Ag^+ Hg_2^{2+} Pb^{2+}	AgCl ; White Hg_2Cl_2 ; White $PbCl_2$; White
2 (A)	H_2S in presence of dil HCl (insoluble in YAS)	Hg^{2+} Pb^{2+} Bi^{3+} Cu^{2+} Cd^{2+}	HgS ; Black PbS ; Black Bi_2S_3 ; Black CuS ; Black CdS ; Yellow
2 (B)	H_2S in presence of dil HCl (Soluble in YAS)	As^{3+} Sb^{3+} Sn^{2+} Sn^{4+}	$As_2\,S_3$; Yellow Sb_2S_3 ; Orange SnS ; Brown SnS_2 ; Yellow
3	NH_4OH in presence of NH_4Cl	Fe^{3+} Cr^{3+} Al^{3+}	$Fe(OH)_3$; Reddish brown $Cr(OH)_3$; Green $Al(OH)_3$; Gelatinous white
4	H_2S in presence of NH_4OH and $NH_4\,Cl$	Zn^{2+} Mn^{2+} Co^{2+} Ni^{2+}	ZnS ; Dirty white MnS ; Buff (or light pink) CoS ; Blak NiS ; Black
5	$(NH_4)_2\ CO_3$ in presence of NH_4OH	Ba^{2+} Sr^{2+} Ca^{2+}	$BaCO_3$; White $SrCO_3$; White $CaCO_3$; White
6	$Na_2\ HPO_4$ in presence of NH_4OH	Mg^{2+}	$Mg(NH_4)\ PO_4$; White

Extra Key Points :

(a). Hg^{+1} (us) → Unstable in aqueous solution. [Stable in dimer form (Hg_2^{+2}) ⇒ Hg_2Cl_2]

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

S S 2S

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

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

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

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

S S 2S

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

(c). 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.

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

$Ksp_{II} < Ksp_{IV} <$ [radicals of IInd and IV group] $[S^{-2}]$

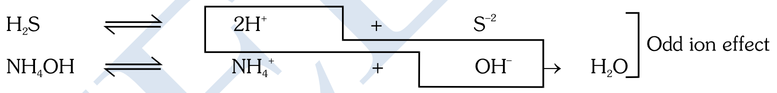
When H_2S gas pass in acidic medium $[H^+]$ then only IInd group radicals are precipitated.

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

$Ksp_{II} <$ [radicals of IInd and IV group] $[S^{-2}] < Ksp_{IV}$

(d). 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 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.

(e). 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 $[S^{-2}]\uparrow$

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

(f). In V group, order of K_{sp} and precipitation is –

$BaCO_3 < SrCO_3 < CaCO_3$

Additional Information

(A) 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

(B) Isohydric solution –

If different–different solution has same pH are called isohydric solution.

Note :– 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 $[H^+]_1 = [H^+]_2$

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

$$\boxed{Ka_1 C_1 = Ka_2 C_2} \text{ or } \boxed{\frac{Ka_1}{V_1} = \frac{Ka_2}{V_2}} \quad \because \boxed{C \propto \frac{1}{V}}$$

(C) Strength of Acids and Bases :–

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

If there are two weak acids HA_1 and HA_2 which has concentration C_1 and C_2, degree of ionisation α_1 and α_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 $C_1 = C_2$

$$\boxed{\frac{\text{Strength of weak acid } HA_1}{\text{Strength of weak acid } HA_2} = \sqrt{\frac{Ka_1}{Ka_2}}}$$

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

(D) Resultant concentration :– When two non reactive electrolytes AB and CD which has concentration C_1 and C_2 and volume V_1 and V_2 respectively are mixed in one container then total volume of container is $(V_1 + V_2)$ and resultant concentration of each electrolyte is–

Resultant concentration of AB $\Rightarrow$ $C'_1 = \dfrac{C_1V_1}{V_1 + V_2}$

Resultant concentration of CD $\Rightarrow$ $C'_2 = \dfrac{C_2V_2}{V_1 + V_2}$

If volume of both electrolytes are equal i.e. $V_1 = V_2 = V$

then resultant concentration of each electrolyte is becomes half of initial concentration.

$$C'_1 = \frac{C_1V}{V+V} = \frac{C_1}{2}, \qquad C'_2 = \frac{C_2V}{V+V} = \frac{C_2}{2}$$

$$\boxed{C'_1 = \frac{C_1}{2}} \text{ and } \boxed{C'_2 = \frac{C_2}{2}}$$

Note : 'n' different solutions of non reactive different electrolytes are mixed in equal volume then resultant concentration of each electrolyte is 1/n times.

Ex.15 The solubility of $BaSO_4$ in water is $1.07 \times 10^{-5}mol^{-3}$ estimate its solubility product

Sol. Solubility equilibrium for $BaSO_4$ is

$$BaSO_{4_{(s)}} \rightleftharpoons Ba^{+2}_{(eq)} + SO^{2-}_{4(eq)}$$

$S = 1.07 \times 10^{-5}m$

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

Ex 16 The solubility product of AgBr is 5.2×10^{-13}, calculate its solubility in mol dm^{-3} and g dm^{-3}.

Sol. $AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

$K_{sp} = (s)^2$

$$S = \sqrt{K_{sp}} = \sqrt{5.2 \times 10^{-13}} = 7.2 \times 10^{-7} \text{ mol dm}^{-3}$$

The solubility in g dm^{-3} = molar solubility × molar mass = $7.2 \times 10^{-7} \times 107.0 = 1.35 \times 10^{-4}$ g dm^{-3}

Ex. 17 Equal volume of 0.04M $CaCl_2$ & 0.0008M Na_2SO_4 are mixed. Will a precipitate form ? K_{sp} for $CaSO_4 = 2.4 \times 10^{-5}$

Sol.
$$\underset{0.04V}{CaCl_2} + \underset{0.0008V}{Na_2SO_4} \rightarrow \underset{0}{CaSO_4} + \underset{0}{2NaCl}$$

Suppose VmL of both are mixed

$$[Ca^{2+}] = \frac{0.04v}{2v}; [SO_4^{2-}] = \frac{0.0008M}{2v}$$

$$\therefore [Ca^{+2}][SO_4^{2-}] = \frac{0.04v}{2v} \times \frac{0.0008V}{2v} = 8 \times 10^{-6}$$

Thus $[Ca^{+2}][SO_4^{2-}]$ in solution < KsP

$\therefore$ $CaSo_4$ will not precipitate.

PRACTICE SECTION-03

Q.1 The relative strength of two weak acids HCOOH ($K_a = 2 \times 10^{-4}$) and HF ($K_a = 6.6 \times 10^{-4}$) is :

(1) 0.30 (2) 0.55 (3) 3.33 (4) 1.82

Q.2 The solubility product of AgCl is 4.0×10^{-10} at 298 K. The solubility of AgCl in 0.04m $CaCl_2$ will be?

(1) 2×10^{-5} m (2) 1×10^{-4} m (3) 5×10^{-9} m (4) 2.2×10^{-4} m

Q.3 On adding 0.1 M solution each of Ag^+, Ba^{2+}, Ca^{2+} ions in a Na_2SO_4 solution species first precipitated is
(K_{sp} $BaSO_4 = 10^{-11}$, K_{sp} $CaSO_4 = 10^{-6}$, K_{sp} $Ag_2SO_4 = 10^{-5}$)

(1) Ag_2SO_4 (2) $BaSO_4$ (3) $CaSO_4$ (4) all of these

Q.4 Equal volumes of the following Ca^{2+} and F^- solutions are mixed. In which of the solutions will precipitation occur? (K_{sp} of $CaF_2 = 1.7 \times 10^{-10}$)

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

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

Q.5 Solubility product constants (K_{sp}) of salts of types MX, MX_2 at M_3X at temperature T are 4.0×10^{-8}, 3.2×10^{-14} and 2.7×10^{-15} respectively. Solubility (mol dm^{-3}) of the salts at temperature T are in the order.

(1) $MX > MX_2 > M_3X$ (2) $M_3X > MX_2 > MX$ (3) $MX_2 > M_3X > MX$ (4) $MX > M_3X > MX_2$

Q.6 At 25°C, the solubility product of Hg_2Cl_2 in water is 3.2×10^{-17} mol^3 dm^{-9}. What is the solubility of Hg_2Cl_2 in water at 25°C?

(1) 1.2×10^{-12} M (2) 3.0×10^{-6} M (3) 2×10^{-6} M (4) 1.2×10^{-16} M

Q.7 How many times 1 M CH_3 COOH solution should be diluted so that pH of the solution is doubled?
(K_a of acetic acid = 1.8×10^{-5})

(1) 210 times (2) 200 times (3) 5.55×10^2 times (4) 5.55×10^4 times

ANSWER KEY

Que.	1	2	3	4	5	6	7
Ans.	2	3	2	4	4	3	4

pH

INTRODUCTION

pH of some important substance :-	
(i) Blood $\longrightarrow$	7.4
(ii) Tear $\longrightarrow$	7.4
(iii) Gastric Juice $\longrightarrow$	1 to 3
(iv) Soft drink $\longrightarrow$	2 to 4
(v) Acidic rain $\longrightarrow$	6
(vi) Milk $\longrightarrow$	6.3 to 6.6
(vii) Water $\longrightarrow$	7

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)

(v) pH of mixture of acid and base.

(a) pH of mixture of acids :

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

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

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

(b) pH of mixture of base :

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

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

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

(c) pH of mixture of acids and bases:

For acid:

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

For base:

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

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

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

$[H^+] = N$

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

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

$[OH^-] = N$

BUFFER SOLUTION

INTRODUCTION

A. Definition:

Those solution which have resistive nature for pH change are called Buffer Solution.

B. Properties of buffer solution:

(i) pH of buffer solution never depends on their concentration and dilution. (water add or escape).

(ii) When some amount of strong acid and strong base are mixed in buffer solution then pH of buffer solution is almost unchange.

Note :- Change in pH of a solution is due to change in concentration of present free H^+ or OH^- ions.

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.

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

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

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

Ex. CH_3COOH + CH_3COONa
WA WASB

- Weak acid + Strong base

1. When $N_1V_1 = N_2V_2 \Rightarrow$ Salt (WA SB)
2. When $N_1V_1 > N_2V_2 \Rightarrow$ Acidic buffer solution (WA + WASB salt)
   $[CH_3COOH + CH_3COONa]$
   $[CH_3COOH + CH_3COO^-]$ Na^+ (Spectator ion)

Case 1. When mixing of acid $[H^+]$

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

Case 2. When mixing of base $[OH^-]$

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

ie. in aqueous solution if weak acid and its conjugate are present then it is definitely acidic buffer solution.

(a) pH of acidic buffer solution:

CH_3COOH + CH_3COONa

Acid + Salt

$[CH_3COOH$ + $CH_3COO^-] + Na^+$

Acid Conjugate

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

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

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

taking –log on both sides

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

Henderson's equation:

$$\boxed{pH = pKa + \log\frac{[Salt]}{[Acid]}}$$

(b) pH range of acidic buffer solution:

CH_3COOH CH_3COONa

1 : 10 $pH = pKa + \log\frac{10}{1}$

$pH = pKa + 1$

10 : 1 $pH = pKa - 1$

So pH range

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

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

CH_3COOH CH_3COONa

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

$$\boxed{pH = pK_a}$$

(ii) Basic buffer solution :

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

Ex. $NH_4OH + NH_4Cl$ $[NH_4^+] = [NH_4Cl]$

$[NH_4OH + NH_4^+] + Cl^-$ [Spectator]

Case 1. When mixing of acid $[H^+]$

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

Case 2. When mixing of base $[OH^-]$

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

(a) pOH of basic buffer solution:

$[NH_4OH + NH_4Cl]$

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

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

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

Henderson's equation : $\boxed{pOH = pK_b + \log\frac{[Salt]}{[Base]}}$

(b) pOH range of basic buffer solution:

NH_4OH : NH_4Cl
1 : 10
10 : 1

$\boxed{pOH = pK_b \pm 1}$

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

NH_4OH : NH_4Cl
1 : 1

$\boxed{pOH = pK_b}$

BUFFER CAPACITY

Definition :- It is defined as the number of moles of acid (or base) require by one litre of a buffer solution for changing its pH by one unit.

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

Ex.18 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.

Sol. B.C. = $\frac{2}{0.5} = 4$

Ex.19 How much volume of 0.2 m solution of acid should be added to 100mL or 0.2m soln or sodium are take to prepare a buffer sol^n of pH = 6 ?
pKa or CH_3COOH = 4.74

Sol. $P^H = P^{ka} + \log\frac{Salt}{Acid}$

$$\log\frac{[Salt]}{[Acid]} = P^H - P^{ka} = 6 - 4.74 = 1.26$$

$$\frac{[Salt]}{[Acid]} = 18.2$$

Mole of CH_3COONa in solution $\frac{100\times0.2}{1000} = 0.02$

Let volume of 0.2 acetic Acid added $\frac{U\times0.2}{1000}$

$$\frac{0.2}{U\times\frac{0.2}{1000}} = 10.2$$

U = 5.49 mL

PRACTICE SECTION-04

Q.1 pH value of which of the following is NOT equal to one?

(1) 0.1 M CH_3COOH (2) 0.1 M NHO_3
(3) 0.05 M H_2SO_4 (4) 50 cm^3 0.4 M HCl + 50 cm^3 0.2 M NaOH

Q.2 Calculate the pH of a solution which contains 100 mL of 0.1 M HCl and 9.9 mL of 1.0 M NaOH.

(1) 3.04 (2) 6.08 (3) 11.96 (4) None

Q.3 Which one of the following will decrease the pH of 50 mL of 0.01 M hydrochloric acid?

(1) Addition of 50 mL of 0.01 M HCl (2) Addition of 50 mL of 0.002 M HCl
(3) Addition of metallic zinc (4) Addition of 5 mL of 1 M HCl

Q.4 The pH of a solution obtained by mixing 50 mL of 0.2 M HCl with 50 mL of 0.20 M CH_3COOH is

(1) 0.30 (2) 0.70 (3) 1.00 (4) 2.00

Q.5 40 ml of 0.1 M ammonia 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

Q.6 How many litres of water must be added to 1 litre of an aqueous solution of HCl with pH of 1 to create an aqueous solution of with pH of 2?

(1) 9.0 L (2) 0.1 L (3) 0.9 L (4) 2.0 L

Q.7 In what volume ratio NH_4Cl and NH_4OH solutions (each 1 M) should be mixed to get a buffer solution of pH 9.80? (pKb of NH_4OH = 4.74)

(1) 1 : 2.5 (2) 2.5 : 1 (3) 1 : 3.5 (4) 3.5 : 1

Q.8 A buffer solution contains 0.1 mole of sodium acetate in 1000 cm^3 of 0.1 M acetic acid. To the above buffer solution 0.1 mole of sodium acetate is further added and dissolved. The pH of the resulting buffer is equal to ....

(1) pKa – log 2 (2) pKa (3) pKa + 2 (4) pKa + log 2

ANSWER KEY

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**
Ans.	1	1	4	3	3	1	3	4

INDICATOR

INTRODUCTION

(A) Definition: - Those compound or substance which indicate the nature of solution by changes their self colour.

(B) Types of indicators:

(a). Acidic indicator (HIn)

Ex. Phenolphthalein (HPh)

$$HPh \rightleftharpoons H^+ + Ph^-$$

colourless pink

(b). Basic indicator (InOH)

Ex. Methyl orange (MeOH)

$$MeOH \rightleftharpoons Me^+ + OH^-$$

yellow orange red

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 works in opposite medium (it means if it is acidic then it will work in basic medium and vice versa) it means colour change in opposite medium due to their unionised form is change in ionised form.

Ex. Acidic indicator [HPh]

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

In highly acidic medium $[H^+]$ here :

$[HPh] >> [Ph^-]$

Colourless

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

In highly basic medium $[OH^-]$ here :

$[Ph^-] >> [HPh]$

Pink colour

(a) pH of acidic indicator -

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

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

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

Taking – log on both sides

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

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

$$\boxed{pH = pK_I + \log\frac{[\text{Ionised form}]}{[\text{Unionised form}]}}$$

(b) pOH of basic indicator -

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

$$K_I = \frac{[In^+][OH^-]}{[InOH]}$$

$$\boxed{pOH = pK_I + \log\frac{[\text{Ionised form}]}{[\text{Unionised form}]}}$$

(c) pH range of acidic indicator :

$[HIn]$: $[In^-]$

1 : 10

10 : 1

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

(d) pOH range of basic indicator :

[InOH] : $[In^+]$

1 : 10

10 : 1

$\boxed{pOH = pK_I \pm 1}$

(e) Neutral condition for indicators (No working condition)

For acidic indicator			For basic indicators		
[HIn]	:	$[In^-]$	[InOH]	:	$[In^+]$
1	:	1	1	:	1
	$\boxed{pH = pK_I}$			$\boxed{pOH = pK_I}$	

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

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) and methyl red
3.	WA/SB	7 – 11	Phenolphthalein (HPh)
4.	WA/WB	6.5 – 7.5	Phenol red

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.

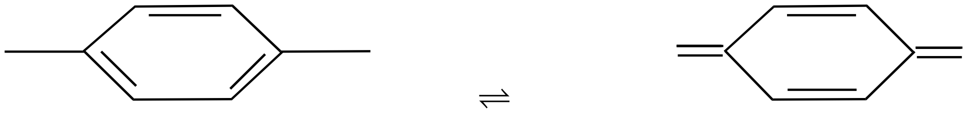
(iv). At the end point, nature of solution depends on the type of titration.

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

BENZENOID AND QUINONOID THEORY OF INDICATORS

(i). According to this theory, all indicators are aromatic compounds.

(ii). Unionised form of indicator is called benzenoid form and ionised form is called quinonoid form.

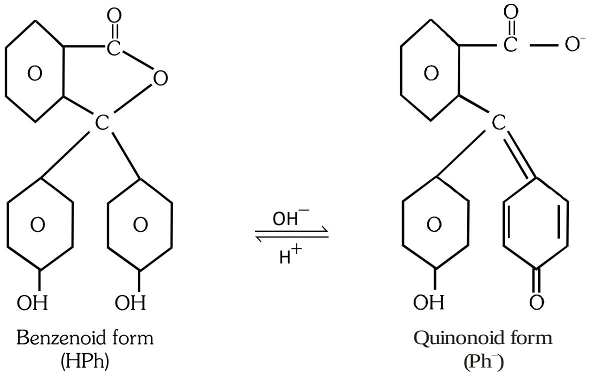

Benzenoid form Quinonoid form

(Unionised form) (Ionised form)

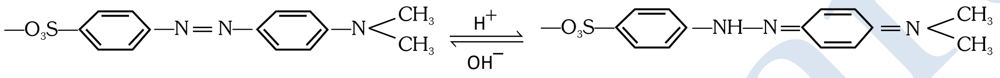
(iii) Every indicator has light colour in its benzenoid form and dark colour in quinonoid form.

(iv) Every indicator works in opposite medium due to their benzenoid form (unionised form) is changes in quinonoid form (ionised form).

(a)



(b)



Benzenoid form (MeOH)

Quinonoid form (MeOH)

Acidic solution (red)

ACID AND BASE CONCEPT

ARRHENIUS CONCEPT (1884)

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

Ex. HCl, HNO_3, H_2SO_4, H_3PO_4, H_2CO_3, H_2S, CH_3COOH etc.

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

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

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

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

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

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

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

$Al(OH)_3 + H_2O \longrightarrow Al^{+3} (aq.) + 3OH^- (aq.)$

(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 i.e. those electrolyte which have high degree of ionisation are strong electrolyte and those which have low degree of ionisation is weak electrolyte.

Ex. For strong electrolytes $\alpha \simeq 100\%$

For weak electrolytes $\alpha < 100\%$

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.

Ex. (i) HCl, HNO_3, H_2SO_4, H_2CO_3, H_2S, CH_3COOH, H_3PO_3 etc.

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

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

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

(a) $HCl\ (Acid) + H_2O\ (Solvent) \longrightarrow H_3O^+ + Cl^-$

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

(c) $NH_4^+\ (Acid) + H_2O\ (Solvent) \longrightarrow NH_3 + H_3O^+$

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

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

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

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

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

(iv) O^{-2}, SO_4^{-2}, CO_3^{-2}, Cl^-, Br^-, I^-, CN^- etc.

(a) $NaOH\ (Base) + H_2O\ (Solvent) \longrightarrow OH^- \xrightarrow[Unstable]{+NaOH_2^+} Na^+ + H_2O$

(b) $HS^-\ (Base) + H_2O\ (Solvent) \longrightarrow H_2S + OH^-$

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

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

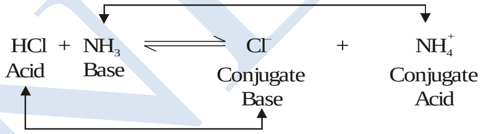
(c) **Nature of water:-** According to this concept nature of water is amphoteric or amphiprotic i.e. water can act as both acid and base.

(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.

Ex.



(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

Ex. (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

LEWIS CONCEPT (1939)

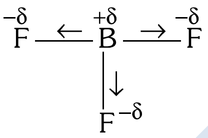
(a) Lewis acid :- According to this concept those species which have self tendency to accept the lone pair of electrons are called acids. i.e. Lewis acid is an electron pair acceptor (electrophilic).

$$\underset{\text{Acid}}{H^+} + \underset{\text{Base}}{:\ddot{\underset{\times}{O}}H^-} \longrightarrow H:\ddot{\underset{\times}{O}}H$$

Classification of Lewis acids :-

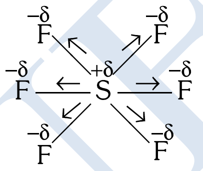
(i) Compounds whose central atom have an incomplete octect (electron deficient)

Ex. BF_3, BBr_3, BCl_3, BI_3, $B(CH_3)_3$, $B(OH)_3$, $AlCl_3$, $GeCl_3$ etc.

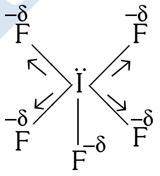


(ii) Compounds whose central atom have vacant d-orbitals and can accept one or more pair of lone electrons.

Ex. SF_4, SF_6, $SnCl_2$, $SnCl_4$, PX_3, PX_5, GeX_4, TeX_4, IF_5, IF_7, etc.



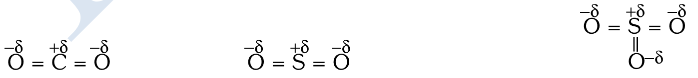
SF_6



IF_5

(iii) Molecules with a multiple bond between atoms of different electronegativities.

Ex. CO_2, SO_2, SO_3 etc.



$$\underset{\text{Lewis acid}}{O = C = O} + \underset{\text{Lewis base}}{OH^-} \longrightarrow {}^-O - \underset{\underset{OH}{|}}{C} = O \longrightarrow HCO_3^-$$

$$\underset{\text{Lewis acid}}{SO_3} + H_2O \longrightarrow \underset{\text{Lewis base}}{H_2SO_4}$$

(iv) Cations:-

Ex. Ag^+, Na^+, Li^+, Al^{+3}, Be^{+2}, Mg^{+2}, I^+, Cl^+, H^+ etc.

False cations :-

Ex. NH_4^+, H_3O^+, $CH_3COOH_2^+$, PH_4^+ etc.

Ex. $Ag^+ + 2 : NH_3 \rightleftharpoons [Ag(:NH_3)_2]^+$

$Cu^{+2} + 2(:H_2O)_2 \longrightarrow [Cu(:H_2O)_4]^{+2}$

(v) Elements which have six electrons in their outermost shell or valence shell.

O–Family (O, S, Se, Te) Po – Radioactive element

$$\underset{\text{Base}}{SO_3^{-2}} + \underset{\text{Acid}}{\ddot{\underset{\cdot\cdot}{O}}:} \longrightarrow [SO_3 \longrightarrow O]^{-2}$$

(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 – Ex.

$$\ddot{N}H_3,\ R-\ddot{N}H_2,\ R_2-\ddot{N}H,\ C_6H_5-\ddot{N}H_2,\ C_5H_5-\ddot{N},\ H_2\ddot{N}-\ddot{N}H_2,$$

$$H-\underset{\cdot\cdot}{\ddot{O}}-H,\ R-\underset{\cdot\cdot}{\ddot{O}}-H,\ R-\underset{\cdot\cdot}{\ddot{O}}-R,\ R-\underset{\cdot\cdot}{\ddot{S}}-R,\ R-\underset{\cdot\cdot}{\ddot{S}}-H, \text{ etc.}$$

Explain :- $\underset{+I}{R} \overset{\rightarrow}{\longrightarrow} \overset{-\delta}{O} \overset{\leftarrow}{\longleftarrow} \underset{+I}{R}$ (Electron density on O, increases due to +I effect)

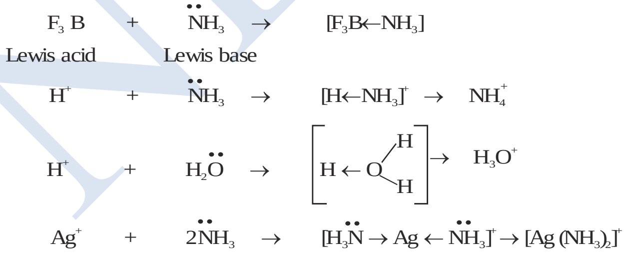
(ii) Anions –

Ex. 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.

Ex.



(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^-).

Ex. $SO_2 <$ SO_3 (Strong acid)

PRACTICE SECTION-05

Q.1 Which of the following statements about HCO_3^- are correct?

1. It is a Bronsted acid
2. It can ionize in water to form CO_3^{2-} (aq)
3. It does not exist in aqueous solution
4. It is a Bronsted base.

Select the correct answer using the codes given below :

(1) 1, 2 and 3 (2) 2, 3 and 4 (3) 1, 3 and and (4) 1, 2 and 4

Q.2 Which one of the following is the correct statement?

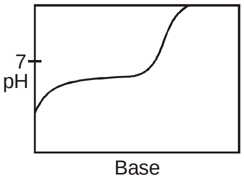
(1) H_2O is the conjugate base of OH^-

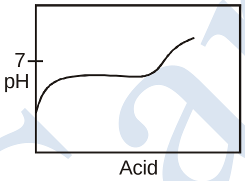
(2) NH_2^- is the conjugate acid of NH_3

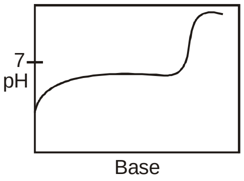
(3) H_2SO_4 is the conjugate acid of HSO_4^-

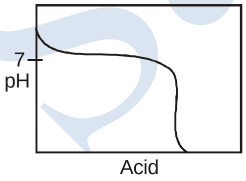
(4) NH_3 is the conjugate base of NH_2^-

Q.3 Which of the following indicates the titration of a weak base against a strong acaid?

(1)


(2)


(3)


(4)


Q.4 The pH range of an indicator is given by

(1) pH = pK + 1 (2) pH = pK – 1 (3) pH = pK (4) pH = pK ± 1

ANSWER KEY

Que.	1	2	3	4
Ans.	4	3	4	4

EXERCISE-I Topic wise Questions

Q.1 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

Q.2 Find out pH of solution having 2×10^{-3} moles of OH^- ion's in 2 litre solution :-

(1) pH = 3 (2) pH = 3 + log2

(3) pH = 3 – log2 (4) pH = 11

Q.3 The pH of a soft drink is 3.82. The hydrogen ion concentration will be:-

(1) 1.96×10^{-2}mol L^{-1} (2) 1.6×10^{-4}mol L^{-1}

(3) 1.96×10^{-5} mol L^{-1} (4) None of them

Q.4 Given $HF + H_2O \xrightarrow{K_a} H_3O^+ + F^-$; $F^- + H_2O \xrightarrow{K_B} HF + OH^-$.

Which realtion is correct ?

(1) $K_b + K_w$ (2) $K_b = \frac{1}{K_w}$

(3) $K_a \times K_b = K_w$ (4) $\frac{K_a}{K_b} = K_w$

OSTWALD'S DILUTION LAW

Q.5 Order of dissociation of 0.1 N CH_3COOH is:- (Dissociation constant = 1×10^{-5})

(1) 10^{-5} (2) 10^{-4}

(3) 10^{-3} (4) 10^{-2}

Q.6 Degree of dissociation of 0.1N CH_3COOH is: ($K_{acid} = 1 \times 10^{-5}$)

(1) 10^{-5} (2) 10^{-4} (3) 10^{-3} (4) 10^{-2}

Q.7 If α 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) $\frac{\alpha(y)^{-1}}{2}$ (2) $y(\alpha)^{-1}$

(3) $\frac{y(\alpha)^{-1}}{2}$ (4) None of them

Q.8 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

Q.9 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

Q.10 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

Q.11 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

Q.12 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

Q.13 If Ka of HCN = 4×10^{-10}, then the pH of 2.5×10^{-1} molar HCN (aq) is:-

(1) 4.2 (2) 4.7

(3) 0.47 (4) 5.0

Q.14 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

Q.15 Correct statement for HCN weak acid at 250C temperature.:-

(1) $\alpha = \frac{K_a}{[H^+]}$ (2) $\alpha = \frac{K_a \times [OH^-]}{K_w}$

(3) (1) & (2) both (4) $K_b = C\alpha^2$

Q.16 For which, dilution law is applicable :

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

(3) CH_3COONa(WASB) (4) None

BIODATA OF WATER

Q.17 At any temperature, the proton concentration of water is

(1) 10^{-14} M (2) Kw

(3) > 10^{-7} M (4) $\sqrt{K_w}$

Q.18 Ionic product of water will increase, if :-

(1) Decrease in pressure

(2) Add H^+

(3) Add OH^-

(4) Increase the temperature

Q.19 For water at 25° C, 2×10^{-7} moles per litre is the correct answer for which one of the followoing

(1) $[H^+] + [OH^-]$ (2) $[H^+]^2$

(3) $[OH^-]^2$ (4) $[H^+] - [OH^-]$

Q.20 The pH of 1 N H_2O is :-

(1) 7 (2) > 7 (3) < 7 (4) 0

Q.21 At 25^{0}C, the dissociation constant for pure water is given by :-

(1) $(55.4 \times 10^{14})^{-1}$ (2) 1×10^{-14}

(3) $\frac{1\times10^{-14}}{18}$ (4) None of these

Q.22 Ionic product of water is equal to :-

(1) Dissociation constant of water × $[H_2O]$

(2) Dissociation constant of water × $[H^+]$

(3) Product of $[H_2O]$ and $[H^+]$

(4) Product of $[OH^-]^2$ and $[H^+]$

Q.23 Addition of H^+ and OH^- ion's concentration at 90^{0}C

(1) 10^{-14} (2) 10^{-12}

(3) 2×10^{-6} (4) 2×10^{-7}

Q.24 At 90^{0}C, pure water has $[H_3O^+] = 10^{-6.7}$ mol L^{-1} what is the value of K_W at 90^{0}C:-

(1) 10^{-6} (2) 10^{-12}

(3) 10^{-67} (4) $10^{-13.4}$

Q.25 At 373 K, temp. the pH of pure H_2O can be:-

(1) < 7 (2) > 7 (3) = 7 (4) = 0

Q.26 Choose the correct relation:-

(1) $\frac{pH+pOH}{14} = 7$ (2) pH + pOH = 14

(3) pOH = 14 + pH (4) pH = 14 + pOH

Q.27 The common ion effect is shown by which of the following sets of solutions :-

(1) $BaCl_2 + BaNO_3$ (2) NaCl + HCl

(3) NH_4OH + NH4Cl (4) None

Q.28 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 6.023×10^{16}.

Q.29 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

Q.30 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) $10^{-7} \times (55.4)^{-1}$

SALTS, TYPES O F SALT & CONJUGATE THEORY

Q.31 Aqueous solution of ammonium sulphate

(1) Turns blue litmus to red

(2) Turns red litmus to blue

(3) Bleaches litmus

(4) Has no action on litmus

Q.32 Which of the following is not an acidic salt :-

(1) $NaHSO_4$ (2) HCOONa

(3) NaH_2PO_3 (4) None of them

Q.33 Salt hydrolysis in water is due to following nature of water

(1) Neutral nature (2) Acidic nature

(3) Basic nature (4) Amphiprotic nature

Q.34 Which is a basic salt :–

(1) PbS

(2) $PbCO_3$

(3) $PbSO_4$

(4) $2PbCO_3 . Pb(OH)_2$

Q.35 A salt 'X' is dissolved in water of pH = 7. The salt is made resulting solution becomes alkaline in nature. The salt is made

(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

Q.36 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

HYDROLYSIS OF SALTS

Q.37 When salt NH_4Cl is hydrolysed at 25°C, the pH is

(1) 7 (2) <7 (3) >7 (4) 0

Q.38 Degree of hydrolysis of 0.25 M CH3COONa is 10% what will be the degree of hydrolysis if concentration of CH_3COONa is made 0.01 M

(1) 10% (2) 1% (3) 50% (4) 75%

Q.39 Which salt will not show hydrolysis :-

(1) KCl (2) Na_2SO_4 (3) NaCl (4) All

Q.40 $HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$ is related:-

(1) $h = \sqrt{K_h}$ (2) $h = \sqrt{\dfrac{K_h}{C}}$

(3) $h = \sqrt{\dfrac{K_h}{V}}$ (4) $K_h = \sqrt{hc}$

Q.41 Four solution of NH_4Cl are taken with concentration 1 M, 0.1 M, 0.01 M & 0.001 M. their degree of hydrolysis are h_1, h_2 & h_3, h_4. What is the gradation of degree of hydrolysis ?

(1) $h_1 > h_2 > h_3 > h_4$ (2) $h_1 = h_2 = h_3 = h_4$
(3) $h_4 > h_3 > h_2 > h_1$ (4) None of these

Q.42 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

Q.43 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

Q.44 pH of K_2S solution is:-

(1) 7 (2) Less than 7
(3) More than 7 (4) 0

Q.45 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) $pH = \frac{1}{2} pK_W + \frac{1}{2} pKa - \frac{1}{2} pK_b$

(3) $pH = \frac{1}{2} pK_W + \frac{1}{2} pKa + \frac{1}{2} \log c$

(4) None of above

Q.46 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}

Q.47 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) $[OH^-] = \left(\dfrac{C.K_w}{K_a}\right)^{1/2}$

(4) $[OH^-] = C.\ K_a.\ K_W$.

Q.48 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

Q.49 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)

Q.50 Degree of Hydrolysis of $\dfrac{N}{100}$ solution of KCN is

(Given $Ka = 1.4 \times 10^{-9}$)

(1) 2.7×10^{-3} (2) 2.7×10^{-2}
(3) 2.7×10^{-4} (4) 2.7×10^{-5}

SOLUBILITY & SOLUBILITY PRODUCT (K_{sp})

Q.51 Solubility of AgBr in water is S_1, in 0.01 M NaBr is S_3 and in 0.05 M $AgNO_3$ is S_4. The correct order of these solubilities is :

(1) $S_1 > S_2 > S_3 > S_4$ (2) $S_1 > S_2 = S_3 > S_4$

(3) $S_4 > S_3 > S_2 > S_1$ (4) $S_1 > S_3 > S_2 > S_4$

Q.52 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

Q.53 The solubility of AgCl in 0.1 M NaCl is (K_{sp} of AgCl = 1.2×10^{-10})

(1) 0.05 (2) 1.2×10^{-6}

(3) 2×10^{-5} (4) 1.2×10^{-9}

Q.54 If the solubility of Ag_2SO_4 in 10^{-2} M Na_2SO_4 solution be 2×10^{-6} M then Kap of Ag_2SO_4 will be

(1) 32×10^{-24} (2) 16×10^{-18}

(3) 32×10^{-18} (4) 16×10^{-24}

Q.55 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) $[Hg^{2+}]^2 \times [I^-]^2$

Q.56 At 25°C, the K_{sp} value of AgCl is 1.8×10^{-10}. If 10^{-5} moles of Ag^+ are added to solution then K_{sp} will be :-

(1) 1.8×10^{-15} (2) 1.8×10^{-10}

(3) 1.8×10^{-5} (4) $18 \times 10^{+10}$

Q.57 50 mL of a solution containing 10^{-3} mole of Ag^+ is mixed with 50 mL of a 0.1 M HCl solution. How much Ag^+ remains in solution ? (Kap of AgCl = 1.0×10^{-10})

(1) 2.5×10^{-6} (2) 2.5×10^{-7}

(3) 2.5×10^{-4} (4) 2.5×10^{-10}

Q.58 If the solubility of Ag_2SO_4 in 10^{-2} M Na_2SO_4 solution be 2×10^{-8} M then Kap of Ag_2SO_4 will be

(1) 10^{-6} (2) 10^{-7} (3) 2×10^{-6} (4) 2×10^{-7}

Q.59 Concentration of Ag^+ ions in saturated solution of Ag_2CrO_4 at 20°C is 1.5×10^{-4} mol L^{-1}. At 20°C, the solubility product of Ag_2CrO_4 is :-

(1) 3.3750×10^{-12} (2) 1.6875×10^{-10}

(3) 1.68×10^{-12} (4) 1.6875×10^{-11}

Q.60 If the concentration of CrO_4^{2-} ion in a saturated solution of silver chromate will be 2×10^{-4} M, solubility product of silver chromate will be -

(1) 4×10^{-8} (2) 8×10^{-12}

(3) 32×10^{-12} (4) 6×10^{-12}

Q.61 If the solubility of AgCl (formula mass=143) in water at 25°C is 1.43×10^{-4} g/100 mL of solution then the value of K_{sp} will be:–

(1) 1×10^{-5} (2) 2×10^{-5}

(3) 1×10^{-10} (4) 2×10^{-10}

Q.62 The solubility product of As2S3 is given by the expression:-

(1) $Ksp = [As^{3+}] \times [S^{-2}]$

(2) $Ksp = [As^{3+}]^1\ [S^{-2}]^1$

(3) $Ksp = [As^{3+}]^3\ [S^{-2}]^2$

(4) $Ksp = [As^{3+}]^2\ [S^{-2}]^3$

Q.63 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) $2916\ S^8$

Q.64 If solubility product of the base $M(OH)_3$ is 2.7×10^{-11}, the concentration of OH^{-1} will be

(1) 3×10^{-3} (2) 3×10^{-4}

(3) 10^{-3} (4) 10^{-11}

Q.65 If the solubility of Ag_2CrO_4 is S mole/litre. Its solubility product is:

(1) S^2 (2) S^3 (3) $4S^3$ (4) $2\ S^3$

APPLICATION OF SOLUBILITY PRODUCT (K_{sp})

Q.66 At 30°C, In which of the one litre solution, the solubility of Ag_2CO_3 (solubility product = 8×10^{-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

Q.67 The addition of KCl to AgCl decreases the solubility of AgCl, because
(1) K_{sp} of AgCl decreases
(2) K_{sP} of AgCl increases
(3) Solution becomes unsaturated
(4) Ionic product exceeds the k_{sp} value

Q.68 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 $AgNO_3$

Q.69 Solubility product of a sparingly soluble salt AX_2 is 3.2×10^{-11}. Its solubility in mol/litre is:
(1) 5.6×10^{-6} (2) 3.1×10^{-4}
(3) 2×10^{-4} (4) 4×10^{-4}

Q.70 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) $[Ag^+] [SCN^-]^2 < K_{sp}$

Q.71 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) $s = \frac{1}{2}S$

Q.72 It a solution contains 10^{-6} M each of X^-, Y^{2-} and Z^{3-} ions, then upon addition of $AgNO_3(s)$ slowly to the above solution with stirring : (Given : $K_{ap}(AgX) = 9 \times 10^{-14}$, $K_{ap}(Ag_2Y) = 4.9 \times 10^{-21}$, $K_{ap}(Ag_3Z) = 5.12 \times 10–^{28}$)
(1) Ag_3Z will be the first one to precipitate out.
(2) Ag_2Y will be the first one to precipitate out.
(3) AgX will be the first one to precipitate out.
(4) Nothing can be said with certainity.

Q.73 Which of the following metal sulphide has maximum solubility in water?
(1) CdS ($K_{sp} = 36 \times 10^{-30}$)
(2) FeS ($K_{sp} = 11 \times 10^{-20}$)
(3) HgS ($K_{sp} = 36 \times 10^{-54}$)
(4) ZnS ($K_{sp} = 11 \times 10^{-22}$)

Q.74 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) 4×10^{-3} M (2) 0.4×10^{-4} M
(3) 4×10^{-2} M (4) 4×10^{-4} M

Q.75 Which of the following has maximum solubility (Ksp value is given in brackets) :-
(1) HgS (1.6×10^{-54})
(2) PbSO4 (1.3×10^{-8})
(3) ZnS (7.0×10^{-26})
(4) AgCl (1.7×10^{-10})

Q.76 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

Q.77 The solubility product of three sparingly soluble salts are given below :

No.	**Formula**	**Solubility product**
1	PQ	4.0×10^{-20}
2	PQ2	3.2×10^{-14}
3	PQ3	2.7×10^{-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

Q.78 The Ksp value for $Gd(OH)_3$ is 2.8×10^{-23}, the pH at which $Gd(OH)_3$ begins to precipitate is:-
(1) 6.08 (2) 5.08 (3) 8.47 (4) 4.08

Q.79 If the solubility product of $AgBrO_3$ and Ag_2SO_4 are 5.5×10^{-5} and 2×10^{-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) $sAgBrO_3 = sAgSO_4$

Q.80 Solubility product of $Mg(OH)_2$ is 1×10^{-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

Q.81 Consider (1) $Zn(OH)_2$ (2) $Cr(OH)_3$ (3) $Mg(OH)_2$ (4) $Al(OH)_3$ which hydroxide is precipitated by NH_4OH containing NH_4Cl :-

(1) 1, 2 (2) 2, 4
(3) Only 4 (4) 1, 2, 3 and 4

Q.82 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

Q.83 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

Q.84 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

Q.85 The solubility product of hydroxides of Mg^{+2}, Zn^{+2}, and Fe^{+3} decreases as
Ksp $Mg(OH)_2$ > Ksp $Zn(OH)_2$ > Ksp $Fe(OH)_3$ 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) $Zn(OH)_2$, $Mg(OH)_2$, $Fe(OH)_3$

Q.86 What is the molar concentration of chloride ion for 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

pH

Q.87 10^{-2} mole of NaOH was added to 10 litre of water. The pH wil chage by

(1) 4 (2) 3 (3) 11 (4) 7

Q.88 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

Q.89 Find out (OH^-) concentration in pH = 3.28 solution

(1) 5.3×10^{-4} (2) 5.3×10^{-10}
(3) 1.8×10^{-10} (4) 1.8×10^{-11}

Q.90 Calculate the pH of resulting solution obtained by mixing 50 mL of 0.6 N HCl and 50 ml of 0.3 N NaOH

(1) 0.1 (2) 0.8 (3) 2.1 (4) 4

Q.91 The pH of HCl is 5. It is diluted by 1000 times Its pH will be

(1) 5 (2) 8 (3) 2 (4) 6 – 7

Q.92 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

Q.93 What is H^+ ion concentration of 5×10^{-3} M H_2CO_3 solution having a 10% dissociation:-

(1) 10^{-3} (2) 10^{-2}
(3) 10^{-1} (4) 5×10^{-2}

Q.94 pH of 3×10^{-3} M solution of H_3X will be Assuming $\alpha_1 = 1/3$, α_3 = negligible

(1) 2.40 (2) 3.0 (3) 3.4771 (4) 4.0

Q.95 What will be th pH and % α (degree of hydrolysis) respectively for the salt BA of 0.1 M concentration ?
Given : K_a for HA = 10^{-6} and K_b for BOH = 10^{-6}

(1) 5.1 % (2) 7.10 %
(3)9.0.01 % (4) 7.0.01 %

Q.96 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

Q.97 10^{-6} M HCl is diluted to 100 times. Its pH is :-
(1) 6.0 (2) 8.0 (3) 6.95 (4) 9.5

Q.98 pH of an aqueous solution of HCl is 5. If 1 c.c. of this solution is diluted to 1000 times. The pH will become
(1) 8 (2) 5 (3) 6.9 (4) None

Q.99 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

Q.100 Which relation is wrong
(1) $10^{-pH} + 10^{-pOH} = 10^{-14}$
(2) pH $\alpha \frac{1}{\left[H^+\right]}$
(3) $K_w \alpha T$
(4) dissociation constant of water K = 1.8×10^{-16}

Q.101 Find the pH of 0.1 M $NaHCO_3$.
(1) 3.7 (2) 8.4
(3) 9.6 (4) None of these

Q.102 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

Q.103 Which wil have pH closer to 1 :
(1) 100 ml N/10 HCl + 100 ml N/10 NaOH
(2) 55 ml N/10 HCl + 45 ml N/10 NaOH
(3) 10 ml N/10 HCl + 90 ml N/10 NaOH
(4) 75 ml N/5 HCl + 25 ml N/5 NaOH

Q.104 An aqueous solution whose pH = 0 is :-
(1) Basic (2) Acidic
(3) Neutral (4) Amphoteric

Q.105 The pH of a 10^{-10} M NaOH solution is nearest to :-
(1) 10 (2) 7 (3) 4 (4) – 10

Q.106 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

Q.107 Which would decrease the pH of 25 cm^3 of a 0.01 M solution of hydrochloric acid:-
(1) The addition of 25 cm^3, 0.005 M hydrochloric acid
(2) The addition of 25 cm^3, 0.02 M hydrochloric acid
(3) The addition of magnesium metal
(4) None

Q.108 The pH of a 0.02 M ammonia solution which is 5% ionised will be :-
(1) 2 (2) 11 (3) 5 (4) 7

Q.109 Pure water is kept in a vessel and it remains exposed to atmospheric CO_2 which is absorbed. Then the pH will be :-
(1) Greater than 7
(2) Less than 7
(3) 7
(4) Depends on ionic product of water

Q.110 pH of 10^{-1} M formic acid is:-
(1) 1 (2) > 1 (3) < 1 (4) 13

Q.111 For $\frac{N}{10}$ H_2SO_4, pH value is :-
(1) 1 (2) 0.586 (3) 0.856 (4) None

Q.112 How many moles of NaOH must be removed from 1 litre of aqueous solution to change its pH from 12 to 11.
(1) 0.009 (2) 0.01 (3) 0.02 (4) 0.1

Q.113 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

Q.114 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

Q.115 What is the ΔpH (final – initial) for 1/3 & 2/3 stages of neutralizaition of 0.1 M CH_3COOH with 0.1 M NaOH :
(1) +2log 2 (2) –2 log 3
(3) 2 log 1/4 (4) 2 log 2/3

Q.116 One litre solution contains 1M HOCl [Ka= 10^{-8}] and 1 M NaOH. What is the pH of the solution:-
(1) 8 (2) 11 (3) 5 (4) 2

Q.117 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

Q.118 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 : - [Kw = 1×10^{-14}]
(1) 13 (2) 3.3 (3) 11 (4) 9.8

Q.119 0.1 M H_2S has $K_1 = 10^{-6}$ & $K_2 = 1.5 \times 10^{-12}$. What will be ht concentration of S^{-2} in the solution.
(1) $\approx 10^{-6}$ (2) $\approx 10^{-9}$
(3) $\approx 1.5 \times 10^{-12}$ (4) 1.2×10^{-13}

Q.120 Four acids HA, HB, HC and HD form salts with NaOH of pH 7,8,9 and 10 respectively when each solution was 0.1 M, the strongest acid is :
(1) HA (2) HB (3) HC (4) HD

Q.121 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.

Q.122 What is the pH of 1 M CH_3COONa solution? Ka of acetic acid = 1.8×10^{-5}, $K_w = 10^{-14}$ mol^2 L^{-2}
(1) 2.4 (2) 3.6 (3) 4.8 (4) 9.4

BUFFER SOLUTION & INDICATORS

Q.123 The volume of 0.2 M NaOH needed to prepare a buffer of pH 4.74 with mL of 0.2 M acetic acid (pK_b of CH_3COO^- = 9.26) is :
(1) 50 mL (2) 25 mL (3) 20 mL (4) 10mL

Q.124 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 (Ka = 1.34×10^{-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

Q.125 $\frac{N}{10}$ 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 :- [Ka = 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

Q.126 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.

Q.127 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 K_b for NH_3 equals 1.8×10^{-5}. What is the pH of this solution?
(1) 8.73 (2) 9.08
(3) 9.44 (4) 11.72

Q.128 pKb 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

Q.129 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

Q.130 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

Q.131 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

Q.132 The pink colour of phenolphthalein in alkaline medium is due to -
(1) Negative ion (2) Positive ion
(3) OH^- ions (4) Neutral form

Q.133 An acid indicator is represented by HI_n.($K_a = 10^{-6}$). The range of chage of colour for the indicator is

(1) 3 – 5 (2) 4 – 6 (3) 5 – 7 (4) 6 – 8

Q.134 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

Q.135 100 ml. of 0.1, M NaOH solution is titrated with 100 mL of 0.05 M H_2SO_4 solution. The pH of the resulting solution is : (For H_2SO_4, $K_{a1} = \infty$, $K_{a2} = 10^{-2}$)

(1) 7 (2) 7.2 (3) 7.4 (4) 6.8

Q.136 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) $CH_3COOH + NH_4OH$

Q.137 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

Q.138 A certain acidic buffer solution contains equal concentration of X^- and HX. The Kb for X^- is 10^{-10}. The pH of the buffer is :-

(1) 4 (2) 7

(3) 10 (4) 14

Q.139 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

Q.140 Which solution is not a buffer solution ?

(1) NaCN (2 mole) + HCl (1 mole) in 5 L

(2) NaCN (1 mole) + HCl (1 Mole) in 5 L

(3) NH_3 (2 mole) + HCl (1 mole) in 5 L

(4) CH_3COOH (2 mole) + KOH (1 mole) in 5L

Q.141 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

Q.142 Hunderson equation pH – pKa = 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

Q.143 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 : $Kb(NH_4OH) = 1.8 \times 10^{-5}$

(1) 3.0×10^{-3} (2) 9.0×10^{-4}

(3) 9.0×10^{-3} (4) 3.0×10^{-4}

Q.144 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

Q.145 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. Kb for $NH_3 = 1.8 \times 10^{-5}$:-

(1) 8.1187 (2) 9.8117

(3) 8.8117 (4) None of these

Q.146 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

Q.147 Which of the following is not buffer solution?

(1) CH_3COOH/CH_3COONa

(2) HCl/NaCl

(3) HCOOH/HCOONa

(4) NH_4OH/NH_4Cl

Q.148 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

Q.149 What is the suitable indicator for titration of NaOH and oxalic acid: -

(1) Methyl orange (2) Methyl red
(3) Phenolphthalein (4) Starch solution

Q.150 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

Q.151 A Certain buffer solution contains equal concentration of X^- and HX. The K_a HX is 10^{-8}. The pH of buffer is:

(1) 3 (2) 8
(3) 11 (4) 14

ACID AND BASE

Q.152 The conjugated acid of O^{-2} ion's is :-

(1) O_2^+ (2) H^+
(3) H_3O^+ (4) OH^-

Q.153 Amphoteric oxide is:-

(1) NO_2 (2) CO_2
(3) Al_2O_3 (4) (1) & (3) both

Q.154 The strongest base among the following is

(1) Cl^- (2) CH_3COO^-
(3) HSO_4^- (4) NO_3^-

Q.155 Ionization constant of AOH and BOH base one 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

Q.156 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

Q.157 An example of Lewis acid is:-

(1) CaO (2) CH_3NH_2
(3) SO_3 (4) None of these

Q.158 In the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ water behaves as :-

(1) Acid
(2) Base
(3) Neutral
(4) Both acid & Base

Q.159 Which acts as Lewis base in the reaction

$BCl_3 +: PH_3 \longrightarrow Cl_3B \leftarrow PH_3$

(1) PH_3 (2) BCl_3
(3) Both 1 & 2 (4) None

Q.160 In the dissociation, $H_2A \xrightleftharpoons{K_1} H^+ + HA^-$

$$HA^- \xrightleftharpoons{K_2} H^+ + A^{-2}$$

(1) K1 is equal to K2
(2) K1 is smaller than K2
(3) K1 is greater than K2
(4) K1 is negligible

Q.161 Which is the strongest Lewis base :–
(1) SbH_3 (2) AsH_3 (3) PH_3 (4) NH_3

Q.162 Which of the following can act as both Bronsted acid and a Bronsted base?
(i) $HCOO^-$ (ii) NH_3
(iii) O^{-2} (iv) HSO_4^-
(1) (i) and (ii) (2) (ii) and (iii)
(3) (ii) and (iv) (4) (i) and (iv)

Q.163 Conjugate base of hydrazoic acid is :-

(1) HN_3^- (2) N_3^- (3) N3– (4) N_2^-

Q.164 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

Q.165 In the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, the conjugate acid-base pair is

(1) NH_3 and H_2O (2) NH_3 and OH^-
(3) H_2O and NH_4^+ (4) NH_4^+ and NH_3

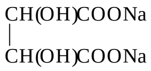
Q.166 When ammonia is added to water it decreases the concentration of which of the following ion

(1) OH^- (2) H_3O^+ (3) NH_4^+ (4) None

Q.167 The strongest acid among the following is -

(1) $ClO_3(OH)$ (2) $ClO_2(OH)$
(3) $SO(OH)_2$ (4) $SO_2(OH)_2$

Q.168 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

Q.169 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

Q.170 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

Q.171 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^-

Q.172 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-}

Q.173 Which one of the following is the weakest base:-

(1) NaOH (2) NH_4OH
(3) $Ca(OH)_2$ (4) $Ba(OH)_2$

Q.174 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

Q.175 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)

Q.176 Which of the following acts as Lewis acid

(1) Cu^{2+} (2) $AlCl_3$
(3) CO_2 (4) All the above

Q.177 CH_3COO^- ion is a :-

(1) Weak conjugate base
(2) Strong conjugate base
(3) Weak conjugate acid
(4) Strong conjugate acid

Q.178 Which of the following is strongest conjugate base

(1) ClO_4^- (2) HCO_3^-
(3) F^- (4) HSO_4^-

Q.179 Which of the following is not a Lewis base :-

(1) CN^- (2) ROH
(3) NH_3 (4) $AlCl_3$

Q.180 Which of the following is acid :-

(1) NH_4Cl (2) $MgCl_2$
(3) CO_2 (4) H_2O

Q.181 Aluminium chloride is :-

(1) Bronsted Lowry acid
(2) Arrhenius acid
(3) Lewis acid
(4) Lewis base

Q.182 Water is a :-

(1) Protogenic solvent
(2) Protophilic solvent
(3) Amphiprotic solvent
(4) Aprotic solvent

Q.183 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

Q.184 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

Q.185 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

EXERCISE-II Analytical Questions

Q.1 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

Q.2 What will be the value of pH of 0.01 mol dm^{-3} CH_3COOH ($K_a = 1.74 \times 10^{-5}$)?

(1) 3.4 (2) 3.6

(3) 3.9 (4) 3.0

Q.3 At 90°C pure water has $[H_3O^+] = 10^{-6}$ mol L^{-1}. The value of KW at 90°C is :-

(1) 10^{-6} (2) 10^{-12}

(3) 10^{-14} (4) 10^{-8}

Q.4 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) $[Sn^{4+}][2S^{2-}]^2$

Q.5 K_a for CH_3COOH is 1.8×10^{-5} and K_b for NH_4OH is 1.8×10^{-5}. The pH of ammonium acetate will be:

(1) 7.005 (2) 4.75

(3) 7.0 (4) Between 6 and 7

Q.6 If the solubility product Ksp of a sparingly soluble salt MX_2 at 25°C is 1.0×10^{-11}, the solubility of the salt in mol L^{-1} at this temperature will be :

(1) 2.46×10^{14} (2) 1.36×10^{-4}

(3) 2.60×10^{-7} (4) 1.20×10^{-10}

Q.7 A solution of $FeCl_3$ in water acts as acidic due to:-

(1) Acidic impurities

(2) Ionisation

(3) Hydrolysis of Fe^{3+}

(4) Dissociation

Q.8 Addition of conc. HCl to saturated $BaCl_2$ solution precipitates $BaCl_2$, because: -

(1) It follows from Le Chatelier's principle

(2) of common-ion-effect

(3) ionic product $[Ba^{++}][Cl^-]$ remains constant in saturated solution

(4) At constant temperature, the product $[Ba^{2+}][Cl^-]^2$ remains constant in a saturated solution.

Q.9 At 25^0 C, OH^- ion concentration in 10^{-5} M HCl(aq.) will be :-

(1) Zero (2) 10^{-9} M

(3) 10^{-5}M (4) 10^{-2}M

Q.10 The correct order of increasing solubility of AgCl in (A) water (B) 0.1 M NaCl

(C) 0.1 $BaCl_2$ (D) 0.1 M NH_3 is

(1) D > A > B > C (2) D > C > B > A

(3) B > A > D > C (2) A > D > B > C

Q.11 The dissociation constant for water is 1×10^{-14} mol^2L^{-2}, what is the pH of a 0.001 M KOH solution

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

Q.12 The solubility of A_2X_3 is y mol dm^{-3}. Its solubility product is :-

(1) $6y^4$ (2) $64y^4$ (3) $36y^5$ (4) $108y^5$

Q.13 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

Q.14 If solubility product of AgCl at 25° C is 5×10^{-13} then solubility of its :

(1) 5×10^{-13} (2) 7.1×10^{-7}

(3) 2.5×10^{-13} (4) 2.5×10^{-6}

Q.15 The precipitate of CaF_2 ($K_{sp} = 1.7 \times 10^{-10}$) is obtained when equal volumes of the following are mixed?

(1) 10^{-4} M Ca^{2+} + 10^{-4} M F^-

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

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

(4) 10^{-3} M Ca^{2+} + 10^{-5} M F^-

Q.16 Which statement is incorrect for solubility product (Ksp):-
(1) Constant at constant temperature
(2) It behave as equilibrium constant
(3) It is more than zero for soluble salt
(4) It's unit is M^2 always

Q.17 When add 0.01 M HCl in aqueous solution of acetic acid
(1) CH_3COO^- molar conc. is decrease
(2) CH_3COOH molar conc. is decrease
(3) CH_3COO^- molar conc. is increase
(4) NOT

Q.18 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.

Q.19 At 298 K, the solubility of $PbCl_2$ is 2×10^{-2} mol L^{-1} then K_{sp} = ? :-
(1) 1×10^{-7} (2) 3.2×10^{-7}
(3) 1×10^{-5} (4) 3.2×10^{-5}

Q.20 The pH of saturated aqueous solution of $Ba(OH)_2$ is 10. If the K_{sp} of $Ba(OH)_2$ is 5×10^{-13}, then the concentration of Ba^{2+} ions in the solution is
(1) 1×10^{-5} (2) 1×10^{-3}
(3) 5×10^{-5} (4)1×10^{-2}

Q.21 A solution of $MgCl_2$ in water has pH :
(1) < 7 (2) > 7 (3) 7 (4) 14.2

Q.22 If solubility product of AB_2 is 4×10^{-12} then solubility is :-
(1) 4×10^{-12} (2) 10^{-12}
(3) 1×10^{-4} (4) 2×10^{-4}

Q.23 Which of the following salt undergo hydrolysis
(1) CH_3COONa (2) KNO_3
(3) NaCl (4) K_2SO_4

Q.24 The minimum concentration of silver ions that is required to start the precipitation of Ag_2S ($K_{sp} = 10^{-51}$) in a 0.1 M solution of $S^{2-}$is:
(1) 10^{-49} M (2) 10^{-50} M
(3) 10^{-26} M (4) 10^{-25} M

Q.25 The pH of 1M NH_4Cl(aq) at 25° C should be :
(1) zero (2) between 6 and 7
(3) 7 (4) above 7

Q.26 Ionic product of water increases with :–
(1) Decreasing pressure
(2) Increasing H^+
(3) By mixing OH^-
(4) Increasing temperature

Q.27 In III group what will be done to differentiate between Fe^{+3} and Cr^{+3} :–
(1) Increase the concentration of NH_4OH
(2) Increase the concentration of NH_4^+ ion
(3) Decrease the concentration of OH^-
(4) both (2) and (3)

Q.28 Correct relation is :-
(1) $pKa \times pK_b = pK_w$ (2) $pKa + pK_b = pK_w$
(3) $pKa/pK_b = pK_w$ (4) $pKa - pK_b = pK_w$

Q.29 If k_{sp} of CaF_2 in pure water is 1.70×10^{-10}, then find the solubility of CaF_2 in 0.10M NaF solution :-
(1) 1.70×10^{-10} (2) 1.70×10^{-9}
(3) 1.70×10^{-8} (4) 0.10 M

Q.30 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

Q.31 Which of the following salts has maximum solubility?
(1) HgS, $K_{sp} = 1.6 \times 10^{-54}$
(2) $PbSO_4$, $K_{sp} = 1.3 \times 10^{-8}$
(3) ZnS, $K_{sp} = 7.0 \times 10^{-26}$
(4) AgCl, $K_{sp} = 1.7 \times 10^{-10}$

Q.32 Ka for CH_3COOH is 1.8 × 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

Q.33 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

Q.34 Concentration of a weak acid is 0.1 N and $K_a = 10^{-5}$ then pH will be :-

(1) 4 (2) 3 (3) 2 (4) 5

Q.35 pH of tomato juice is 4.4. Then concentration of H_3O^+ will be :-

(1) 39×10^{-4} (2) 3.9×10^{-5}

(3) 3.9×10^{-4} (4) 3.9×10^{5}

Q.36 On adding NH_4Cl to NH_4OH solution :-

(1) Conc. of OH^- increase

(2) Conc. of NH_4^+ decreases

(3) Conc. of OH^- decreases

(4) Dissociation of NH_4OH increases

Q.37 A salt is precipitated from a solution when :-

(1) Solution is saturated

(2) Ionic product > Solubility product

(3) Ionic product < Solubility product

(4) Solution is unsaturated

Q.38 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

Q.39 The Ksp for $Cr(OH)_3$ is 1.6 × 10^{-30}. The molar solubility of this compound in water is :-

(1) $\sqrt[2]{1.6\times10^{-30}}$ (2) $\sqrt[4]{1.6\times10^{-30}}$

(3) $\sqrt[4]{1.6\times10^{-30}/27}$ (4) 1.6 × 10–30/27

Q.40 An acid HA ionises as

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

The pH of 1.0 M solution is 5. Its dissociation constant would be :-

(1) 1×10^{-10} (2) 5

(3) 5×10^{-8} (4) 1×10^{-5}

Q.41 Equal volumes of tree acid solutions of pH 3, 4 and 5 are mixed in a vessel. What will be the H^+ ion concentration in the mixture?

(1) 1.11×10^{-4} M (2) 3.7×10^{-4} M

(3) 3.7×10^{-3} M (4) 1.11×10^{-3} M

Q.42 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) 1×10^{-7} (2) 3×10^{-7}

(3) 1×10^{-3} (4) 1×10^{-5}

Q.43 The ionic product for water is 1×10^{-14} mol^2L^{-2}, what is the pH of a 0.001M KOH solution :

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

Q.44 Which buffer solution out of the following will have pH > 7 :–

(1) $CH_3COOH + CH_3COONa$

(2) HCOOH + HCOOK

(3) CH_3COONH_4

(4) $NH_4OH + NH_4Cl$

Q.45 pH of 0.01N H_2SO_4 is :-

(1) 1.7 (2) 2.0 (3) 2.3 (4) 2.7

Q.46 Which of the following is the strongest acid :

(1) $Ka = 1 \times 10^{-7}$ (2) $Ka = 1 \times 10^{-5}$

(3) pKa = 9 (4) pKa = 3

Q.47 The concentration of $Ca(OH)_2$ is 0.05 M then pH

(1) 13.0 (2) 12.7 (3) 11.2 (4) 1

Q.48 0.015 g moles of NH_4OH and 0.025g moles of NH_4Cl are present in a solution then pH of mixture is :-

(1) 9.0335 (2) 8.0335 (3) 9.665 (4) 8.665

Q.49 What is the molar solubility of $Fe(OH)_3$ if $K_{sp} = 1.0 \times 10^{-38}$?

(1) 3.16×10^{-10} (2) 1.386×10^{-10}

(3) 1.45×10^{-9} (4) 1.12×10^{-11}

Q.50 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) $0{\cdot}5 \times 10^{-10}$ (2) 2×10^{-10}

(3) 10^{-6} (4) 10^{-8}

Q.51 Which of the following is right for diprotic acid :

(1) $Ka_2 > Ka_1$ (2) $Ka_1 > Ka_2$

(3) $Ka_2 > \frac{1}{Ka_1}$ (4) $Ka_2 = Ka_1$

Q.52 The pKa of a weak acid (HA) is 4.5. The pOH of an aqueous buffer solution of HA in which 50% of the acid is ionized is :-

(1) 2.5 (2) 9.5

(3) 7.0 (4) 4.5

Q.53 pH of solution of 0.0005M $Ca(OH)_2$ is :-

(1) 3.0 (2) –3.0

(3) 11.0 (4) 14.0

Q.54 If pK_b of a base is 7.0, then K_a of its conjugate acid will be

(1) 7 (2) 10^7

(3) 10^{-7} (4) 10^{-14}

Q.55 pKa of Quinoline base is 4.88. What will be the pKa of 0.01M solution of it

(1) 4.88 (2) 0.01 (3) 9.12 (4) 14

Q.56 The pH of an aqueous solution of a 1×10^{-7} M solution of HCl will be :-

(1) 7

(2) slightly less than 7

(3) slightly greater than 7

(4) 1

Q.57 H_3A is weak triprotic acid

$\left(K_{a_1}=10^{-5}, K_{a_2}=10^{-9}, K_{a_3}=10^{-13}\right)$

What is the value of pX of 0.1 M H_3A (aq) solution?

where pX = - log X and X = $\frac{\left[A^{3-}\right]}{\left[HA^{2-}\right]}$

(1) 7 (2) 8 (3) 9 (4) 10

Q.58 Wha tis the hydronium ion concentration of a 0.02 M solution of Cu^{2+} solution of copper (II) perchlorate? The acidity constant of the following reaction is 5 × 10^{-9}.

$$Cu^{2+}(aq.) + 2H_2O(l) \rightleftharpoons Cu(OH)^{+}(aq.) + H_3O^{+}(aq.)$$

(1) 1×10^{-5} (2) 7×10^{-4}

(3) 5×10^{-4} (4) 1×10^{-4}

Q.59 A 1 L solution contains 0.2 M NH_4OH and 0.2 M NH_2Cl. If 1.0 mL of 0.001 M HCl is added to it

(1) 2×10^{-5} (2) 5×10^{-10}

(3) 2×10^{-3} (4) None of these

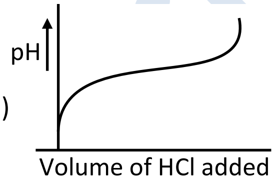
Q.60 A buffer solution is made up of acitec acid [pKa = 5] having conc. = 1.5 and sodium acetate having conc. = 0.15 M. What is the number of OH^- ions present in 1 litre solution?

(1) 10^{-10} N_A (2) 10^{-4} N_A

(3) 10^{-3} N_A (4) 10^{-6} N_A

Q.61 When 100 mL of 0.1 M NaCN solution is titrated with 0.1 M HCl solution the variation of pH of solution with volume of HCl added will be :

(1)



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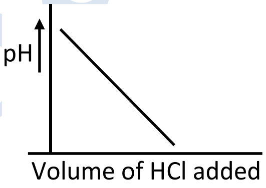

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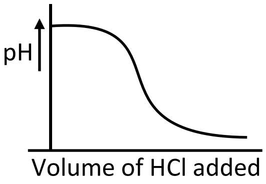

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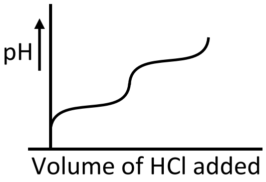


Q.62 1.0 L solution is prepared by mixing 61 gm benzoic acid (pKa = 4.2) with 72 gm of soldier benzoate and then 300 mL M HBr solution was added. The pH of final solution is :

(1) 3.6 (2) 3.8 (3) 4.2 (4) 4.8

Q.63 At 90°C, pure water has $[H^+] = 10^{-6}$ M, if 100 ml of 0.2 M HNO_3 is added to 20 ml of 1 M NaOH at 90°C then pH of the resulting solution will be

(1) 5 (2) 6

(3) 7 (4) None of these

Q.64 When 0.02 moles of NaOH are added to a litre of buffer solution, its pH chages from 5.75 to 5.80. What is its buffer capacity :-

(1) 0.4 (2) 0.05 (3) –0.05 (4) 2.5

Q.65 What amount sodium propanoate should be added to one litre of an aqueous solution containing 0.02 mole of propanoic acid ($K_a = 3 \times 10^{-5}$ at 25°C) to obtain a buffer solution of pH 4.7

(1) 4.52×10^{-2} mol (2) 3.52×10^{-2} mol

(3) 2.52×10^{-2} mol (4) 3×10^{-2} mol

Q.66 In a buffer solution the ration 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) None

Q.67 A 20.0 mL sample of a 0.20 M solution of the weak diprotic acid H_2A is titrated with 0.250 M NaOH. The solution of the second equilvalent point is :

(1) 0.10 M NaHA (2) 0.153 M Na_2A

(3) 0.10 M Na_2A (4) 0.0769 M Na_2A

Q.68 During the titration of a weak diprotic acid (H_2A) against a strong base (NaOH), the pH of the solution half-way to the first equivalent point and that at the first equivalent point are given respectively by :

(1) pK_{a_1} and $pK_{a_1} + pK_{a_2}$

(2) $\sqrt{pK_{a_1}}$ and $\frac{pK_{a_1} + pK_{a_2}}{2}$

(3) pK_{a_1} and $\frac{pK_{a_1} + pK_{a_2}}{2}$

(4) pK_{a_1} and pK_{a_2}

Q.69 Phenolphthalein does not act as an indicator for the titration between :-

(1) KOH and H_2SO_4

(2) NaOH and CH_3COOH

(3) Oxalica acd and $KMnO_4$

(4) $Ba(OH)_2$ and HCl

Q.70 One litre of saturated of $CaCO_3$ is evaporated to dryness, 7.0 g of residue is left. The solubility product for $CaCO_3$ is :-

(1) 4.9×10^{-3} (2) 4.9×10^{-5}

(3) 4.9×10^{-9} (4) 4.9×10^{-7}

Q.71 A_3B_2 is a sparingly solution salt of molar mass M (g mol^{-1}) and solubility x g lit^{-1}. The ration of the molar concentration of B^{3-} to the solubility product of the salt is

(1) $108 \frac{x^5}{M^5}$ (2) $\frac{1}{108} \frac{M^4}{x^4}$

(3) $\frac{1}{54} \frac{M^4}{x^4}$ (4) None

Q.72 What will happn if the pH of the solution of 0.001 M $Mg(NO_3)_2$ solution is adjusted to pH = 9 (K_{sp} of $Mg(OH)_2 = 8.9 \times 10^{-12}$)

(1) ppt will take place

(2) ppt will not take place

(3) Solution will be saturated

(4) None of these

Q.73 Na_3PO_4 which should be added in 10 L of 1.0×10^{-5} M – $BaCl_2$ Solution without any precipitation

(1) 2×10^{-4} (2) 0.328 gm

(3) 0.164 gm (4) 0.82 gm

EXERCISE-III Previous Year Questions

JEE-MAIN

Q.1 The solubility in water of a sparingly soluble salt AB_2 is 1.0×10^{-5} mol L^{-1}. Its solubility product will be **[AIEEE-2003]**

(1) 1×10^{-15} (2) 1×10^{-10}
(3) 4×10^{-15} (4) 4×10^{-10}

Q.2 The solubility of $Mg(OH)_2$ is x mole/lit. then its solubility product is – **[AIIEEE-2002]**

(1) x^3 (2) $5x^3$ (3) $4x^3$ (4) $2x^2$

Q.3 The molar solubility in mol L^{-1} of a sparingly soluble salt MX_4 is 's'. The corresponding solubility procuct is K_{SP} 's' given in terms of K_{SP} by relation : **[AIEEE-2004]**

(1) $s = (K_{SP} / 128)^{1/4}$ (2) $s = (128K_{SP})^{1/4}$
(3) $s = (256K_{SP})^{1/5}$ (4) $s = (K_{SP} / 256)^{1/5}$

Q.4 The solubility product of a salt gaving general formula MX_2, in water is : 4×10^{-12}. The concentration of M^{2+} ions in the aqueous solution of the salt is - **[AIEEE-2005]**

(1) 1.0×10^{-4} M (2) 2.0×10^{-6} M
(3) 4.0×10^{-10} M (4) 1.6×10^{-4} M

Q.5 Hydrogen ion concentration in mol/L in a solution of pH = 5.4 will be - **[AIEEE-2005]**

(1) 3.88×10^{6} (2) 3.98×10^{8}
(3) 3.98×10^{-6} (4) 3.68×10^{-6}

Q.6 In a saturated solution of the sparingly soluble strong electrolyte $AgIO_3$ (molecular mass = 283) the equilibrium which sets in is – **[AIEEE-2005]**

$$AgIO_3 \rightleftharpoons Ag^+_{(aq)} + IO^-_{3(aq)}$$

If the solubility product constant K_{sp} of $AgIO_3$ at a at a given temperature is 1.0×10^{-8}, what is the mass of $AgIO_3$ contained in 100 ml of its saturated solution ?

(1) 28.3×10^{-2} g (2) 2.83×10^{-3} g
(3) 1.0×10^{-7} g (4) 1.0×10^{-4} g

Q.7 The pKa of a weak acid, HA, is 4.80. the pK_b of a weak base, BOH, is 4.78. The pH of an aqueous solution of the corresponding salt, BA, will be - **[AIEEE-2008]**

(1) 9.58 (2) 4.79 (3) 7.01 (4) 9.22

Q.8 Solid $Ba(NO_3)_2$ is gradually dissolved in a 1.0×10^{-4} M Na_2CO_3 solution. At what concentration of Ba^{2+} will a precipitate begin to form? **[AIEEE-2009]**

(1) 8.1×10^{-8} M (2) 8.1×10^{-7} M
(3) 4.1×10^{-5} M (4) 5.1×10^{-5} M

Q.9 Solubility product of silve bromide is 5.0×10^{-13}. The quantity of potassium bromide (molar mass taken as 120 g mol^{-1}) to be added to 1 litre of 0.05 M solution silver mitrate to start the precipitation of AgBr is :- **[AIEEE-2010]**

(1) 5.0×10^{-8} g (2) 1.2×10^{-10} g
(3) 1.2×10^{-9} g (4) 6.2×10^{-5} g

Q.10 In aqueous solution the ionization constants for carbonic acid are
$K_1 = 4.2 \times 10^{-7}$ and $K_2 = 4.2 \times 10^{-11}$ **[AIEEE-2010]**
Select the correct statement for a saturated 0.034 M solution of the carbonic acid :-
(1) the concentration of H^+ is double that of CO_3^{2-}
(2) The concentration of CO_3^{2-} is 0.034 M
(3) The concentration of CO_3^{2-} is greater than that of HCO_3^-
(4) The concentration of H^+ and HCO_3^- are approximately equal

Q.11 At 25°C, the solubility product of $Mg(OH)_2$ is 1.0×10^{-11}. At which pH, will Mg^{2+} ions start precipitating in the form of $Mg(OH)_2$ from a solution of 0.001 M Mg^{2+} ions ions ? **[AIEEE-2010]**

(1) 8 (2) 9 (3) 10 (4) 11

Q.12 The K_{sp} for $Cr(OH)_3$ is 1.6×10^{-30}. The molar solubility of this compound in water is :- **[AIEEE-2011]**

(1) $\sqrt[2]{1.6\times10^{-30}}$ (2) $\sqrt[4]{1.6\times10^{-30}}$
(3) $\sqrt[4]{1.6\times10^{-30} / 27}$ (4) $1.6 \times 10^{-30}/27$

Q.13 An acid HA ionizes as $HA \rightleftharpoons H^+ + A^-$
The pH of 1.0 M solution is 5. Its dissociation constant would be :- **[AIEEE-2011]**
(1) 1×10^{-10} (2) 5
(3) 5×10^{-8} (4) 1×10^{-5}

Q.14 If K_{sp} of CaF_2 at 25°Cis 1.7×10^{-10}, the combination amongst the following which gives a precipitate of CaF_2 is :- **[JEE-MAIN(online)-2012]**

(1) 1×10^{-2} M Ca^{2+} and 1×10^{-5} M F^-
(2) 1×10^{-4} M Ca^{2+} and 1×10^{-4} M F^-
(3) 1×10^{-3} M Ca^{2+} and 1×10^{-5} M F^-
(4) 1×10^{-2} M Ca^{2+} and 1×10^{-3} M F^-

Q.15 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 :- **[AIEEE-2012]**

(1) 1×10^{-7} (2) 3×10^{-7}
(3) 1×10^{-3} (4) 1×10^{-5}

Q.16 How many litres of wter 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 ? **[AIEEE-2013]**

(1) 0.1 L (2) 0.7 L (3) 2.0 L (4) 9.0 L

Q.17 Solid $Ba(NO_3)_2$ is gradually dissolved in a 1.0×10^{-4} M Na_2CO_3 solution. At which concentration of Ba^{2+}, precipitate of $BaCO_3$ begins begins to from ? (K_{sp} for $BaCO_3 = 5.1 \times 10^{-9}$) **[JEE-MAIN(online)-2013]**

(1) 5.1×10^{-5} M (2) 8.1×10^{-7} M
(3) 4.1×10^{-5} M (4) 7.1×10^{-8} M

Q.18 NaOH is a strong base. What will be pH of 5.0×10^{-2} M NaOH solution ? (log2 = 0.3) **[JEE-MAIN(online)-2013]**

(1) 13.70 (2) 13.00 (3) 14.00 (4) 12.70

Q.19 Which of of the following arrangements repredents the correct order of solubilituies of spartingly soluble salts Hg_2Cl_{2+}, $Cr_2(SO_4)_3$, $BaSO_4$ and $CrCl_3$ respectively ? **[JEE-MAIN(online)-2013]**

(1) $\left(\frac{K_{SP}}{4}\right)^{\frac{1}{3}}, \left(\frac{K_{SP}}{108}\right)^{\frac{1}{5}}, (K_{SP})^{\frac{1}{2}}, \left(\frac{K_{SP}}{27}\right)^{\frac{1}{4}}$

(2) $(K_{SP})^{\frac{1}{2}}, \left(\frac{K_{SP}}{4}\right)^{\frac{1}{3}}, \left(\frac{K_{SP}}{27}\right)^{\frac{1}{4}}, \left(\frac{K_{SP}}{108}\right)^{\frac{1}{5}}$

(3) $(K_{SP})^{\frac{1}{2}}, \left(\frac{K_{SP}}{108}\right)^{\frac{1}{5}}, \left(\frac{K_{SP}}{27}\right)^{\frac{1}{4}}, \left(\frac{K_{SP}}{4}\right)^{\frac{1}{3}}$

(4) $\left(\frac{K_{SP}}{108}\right)^{\frac{1}{5}}, \left(\frac{K_{SP}}{27}\right)^{\frac{1}{4}}, (K_{SP})^{\frac{1}{2}}, \left(\frac{K_{SP}}{4}\right)^{\frac{1}{3}}$

Q.20 What would be the pH of a solution obtained by mixing 5 g of acetic acid and 7.5 g of sodium acetate and making the volume equal to 500 mL? **[JEE-MAIN(online)-2013]**

(ka = 1.75×10^{-5}, pKa = 4.76)

(1) 4.76 < pH < 5.0
(2) pH < 4.70
(3) pH of solution will be equal to pH of acetic acid
(4) pH = 4.70

Q.21 In some solution, the concentration of H_3O^+ remains constant even when small amounts of strong acid or strong base are added to them. Tehse solutios are known as: **[JEE-MAIN(online)-2014]**

(1) Colloidal solution (2) True solutions
(3) Ideal solutions (4) Buffer soutions

Q.22 Zirconium phosphate $[Zr_3(PO_4)_4]$ dissociates into three zirconium cations of charge +4 and four phosphate anions of charge –3. If molar solubility of zirconium phosphate is denoted by S and its solubility prduct by K_{sp} then which of the following relationship between S and K_{sp} is correct ? **[JEE-MAIN(online)-2014]**

(1) $S = \{K_{sp}/144\}^{1/7}$ (2) $S = \{K_{sp}/(6912)^{1/7}\}$
(3) $S = \{K_{sp}/(6912)\}^{1/7}$ (4) $S = \{K_{sp}/6912\}^{7}$

Q.23 pKa of a weak acid (HA) and pK_b of a weak base (BOH) are 3.2 and 3.4, respectively. The pH of their salt (AB) solution is **[JEE-MAIN(online)-2017]**

(1) 7.2 (2) 6.9 (3) 7.0 (4) 1.0

Q.24 Addition of sodium hydroxide solution to a weak acid (HA) results in a buffer of pH 6. If ionization constant of HA is 10^{-5}. The ration o fsalt to acid concentration in the buffer solution will be : **[JEE-MAIN(online)-2017]**

(1) 4 : 5 (2) 1 : 10 (3) 10 : 1 (4) 5 : 4

Q.25 50 mL of 0.2 M ammonia solution is treated with 25 mL of 0.2 M HCl. If pK_b of ammonia solution is 4.75, the pH of the mixture will be :-

(1) 8.25 (2) 4.75 (3) 9.25 (4) 3.75

Q.26 Which of the following salts is the most basic in aqueous solution ? **[JEE-MAIN(online)-2018]**

(1) CH_3COOK (2) $FeCl_3$
(3) $Pb(CH_3COO)_2$ (4) $Al(CN)_3$

Q.27 An alkali is titrated against an acid wihth methyl orange as indicator, which of the following is a correct combination? **[JEE-MAIN(online)-2018]**

	Base	Acid	End point
(1)	Strong	Strong	Pinkinsh red to yellow
(2)	Weak	Strong	Yellow to Pinkish red
(3)	Strong	Strong	Pink to colourless
(4)	Weak	Strong	Colourless to pink

Q.28 An aqueous solution contains 0.10 M H_2S and 0.20 M HCl. If equilibrium constants for the formation of HS^- from H_2S is 1.0×10^{-7} and that of S^{2-} from HS^- ions is 1.2×10^{-13} then the concentration of S^{2-} ions in aqueous solution is :- **[JEE-MAIN(offline)-2018]**

(1) 3×10^{-20}
(2) 6×10^{-21}
(3) 5×10^{-19}
(4) 5×10^{-8}

Q.29 A aqueous solution contains an unknown concentration of Ba^{2+}. When 50 mL of a 1 M solution of Na_2SO_4 is added, $BaSO_4$ just begins to precipitate. The final volume is 500 mL. the solubility product of $BaSO_4$ is 1×10^{-10}. What is the original concentration of Ba^{2+} ? **[JEE-MAIN(offline)-2018]**

(1) 2×10^{-9} M
(2) 1.1×10^{-9} M
(3) 1.0×10^{-10} M
(4) 5×10^{-9} M

Q.30 Following four solutions are prepared by mixing different volumes of NaOH and H?Cl of different concentrations, pH of which one of them will be equl to 1 ? **[JEE-MAIN(online)-2018]**

(1) $75\text{mL}\,\frac{M}{5}$ HCl + $25\text{mL}\,\frac{M}{5}$ NaOH

(2) $100\text{mL}\,\frac{M}{10}$ HCl + $100\text{mL}\,\frac{M}{10}$ NaOH

(3) $55\text{mL}\,\frac{M}{10}$ HCl + $45\text{mL}\,\frac{M}{10}$ NaOH

(4) $60\text{mL}\,\frac{M}{10}$ HCl + $40\text{mL}\,\frac{M}{10}$ NaOH

Q.31 The minimum volume of wqter required to dissovbe 0.1 g lead (II) chloride to get a saturated solution (K_{sp} of $PbCl_2 = 3.2 \times 10^{-8}$; atomis mass of Pb = 207 u) is : **[JEE- MAIN(online)-2018]**

(1) 0.36 L
(2) 0.18 L
(3) 17.98 L
(4) 1.798 L

Q.32 If K_{sp} of Ag_2CO_3 is 8×10^{-12}, the molar solubilituy of Ag_2CO_3 in 0.1 M $AgNO_3$ is : **[JEE-MAIN(online)-2019]**

(1) 8×10^{-12} M
(2) 8×10^{-10} M
(3) 8×10^{-11} M
(4) 8×10^{-13} M

Q.33 25 mL of the given HCl solution requires 30 mL of 0.1 M sodium carbonate soluion. What is the volume of this HCl solution required to titrate 30 mL of 0.2 M aqueous NaOh solution ? **[JEE-MAIN(online)-2019]**

(1) 25 mL
(2) 50 mL
(3) 12.5 mL
(4) 75 mL

Q.34 A mixture of 10 m mol of $Ca(OH)_2$ and 2g of soldium sulphate was dissolved in water and the volume ws made up to 100 mL. The mass of calcium salphate formed and the concentration fo OH^- in resulting solution, respectively, are : (Molar mass of $Ca(OH)_2$, Na_2SO_4 and $CaSO_4$ are 74, 143 and 136 g mol^{-1}, respectively; K_{sp} of $Ca(OH)_2$ is 5.5×10^{-6})

(1) 1.9 g, 0.14 mol L^{-1}
(2) 13.6 g, 0.14 mol L^{-1}
(3) 1.9 g, 0.28 mol L^{-1}
(4) 13.6 g, 0.28 mol L^{-1}

Q.35 The pH of rain water, is approximately : **[JEE-MAIN(online)-2019]**

(1) 6.5
(2) 7.5
(3) 5.6
(4) 7.0

Q.36 20 mL of 0.1 M H_2SO_4 solution is added to 30 mL of 0.2 M NH_4OH solution. The pH of the resulatant mixture is : [pK_b of NH_4OH = 4.7]. **[JEE-MAIN(online)-2019]**

(1) 9.4
(2) 5.0
(3) 9.0
(4) 5.2

Q.37 If solubility product of $Zr_3(PO_4)_4$ is denoted by K_{sp} and its molar solubility is denoted by S, then which of the following relation between S and k_{sp} is correct

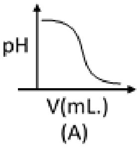
(1) $S = \left(\frac{K_{sp}}{929}\right)^{\frac{1}{9}}$

(2) $S = \left(\frac{K_{sp}}{216}\right)^{\frac{1}{7}}$

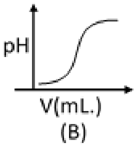
(3) $S = \left(\frac{K_{sp}}{144}\right)^{\frac{1}{6}}$

(4) $S = \left(\frac{K_{sp}}{6912}\right)^{\frac{1}{7}}$

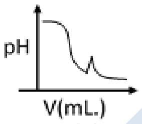
Q.38 In an acid-base titration, 0.1 M HCl solution was added to the NaOH solution fo unknow strength. Which of the following correctly shows the chage of pH of the titrction mixture in this experiment?



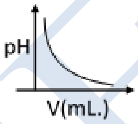
(A)



(B)



(C)



(D)

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

Q.39 The pH of a 0.02M NH_4Cl solution will be

[JEE-MAIN(ONLINE)-2019]

[given K_b(NHOH) = 10^{-5} and log2 = 0.301]

(1) 4.65 (2) 5.35 (3) 4.35 (4) 2.65

Q.40 The molar solubility of $Cd(OH)_2$ is 1.84×10^{-5} M in water. The expected solubility of $Cd(OH)_2$ in a buffer solution of pH = 12 is :

[JEE-MAIN(online)-2019]

(1) 6.23×10^{-11} M (2) 1.84×10^{-9} M

(3) $\frac{2.49}{1.84} \times 10^{-9}$ M (4) 2.49×10^{-10} M

Q.41 Two solutions A and B, each of 100 L was made by dissobing 4g of NaOh and 9.8 g of H_2SO_4 in water, respectively. The pH of the resultant solutions obtained from mixing 40 L of solution A and 10 L of solution B is ____.

[JEE-MAIN(online)-2020]

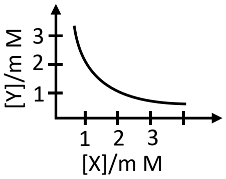
Q.42 3g of acetic acid is added to 250 mL of 0.1 M HCl and the solution made up to 500 mL. To 20 mL of theis solution $\frac{1}{2}$ mL of 5 M NaOh is added. The pH of the solution is

[JEE-MAIN(online)-2020]

[Given : pKa of acetic acid = 4.75, molar mass of acetic acid = 60g/mol, log3 = 0.4771] neglect any chagen in volume

Q.43 The stoichiometry and solubility of a salt with the solubility curve below is, respectively

[JEE-MAIN(online)-2020]



(1) X_2Y, 2×10^{-9} M^3 (2) XY_2, 1×10^{-9} M^3

(3) XY_2, 4×10^{-9} M^3 (4) XY, 2×10^{-6} M^3

Q.44 For the following Assertion and Reason, the correct option is : **[JEE-MAIN(online)-2020]**

Assertion : The pH of water increases with increase in temperature.

Reason : The dissociation of water into H^+ and OH^- is an exothermic reaction.

(1) both assertion and reason are true, but the reason is not the correct explanation fotr the assertion.

(2) Both assertion and reason are false

(3) Assertion is not true, but reason is true

(4) both assertion and reason are true, and the reason is the correct explanation for the assertion.

Q.45 The Ksp for the following dissociation is 1.6×10^{-5} **[JEE-MAIN(online)-2020]**

$$PbCl_{2(s)} \rightleftharpoons Pb^{2}_{(aq)} + 2Cl^{-}_{(aq)}$$

Which of the following choices is correct for a mixture of 30 mL 0.134 M $Pb(NO_3)_2$ and 100 mL 0.4 M NaCl ?

(1) $Q < K_{sp}$

(2) $Q > K_{sp}$

(3) $Q = K_{sp}$

(4) Not enough data provided

Q.46 The solubility of $Cr(OH)_3$ At 298 K is 6.0×10^{-31}. The concentration of hydroxide ions in a saturated solution of $Cr(OH)_3$ will be : **[JEE-MAIN(online)-2020]**

(1) $(18 \times 10^{-31})^{1/4}$ (2) $(2.22 \times 10^{-31})^{1/4}$

(3) $(4.86 \times 10^{-29})^{1/4}$ (4) $(18 \times 10^{-31})^{1/2}$

JEE-ADVANCE

Q.47 What will be the resultant pH when 200 ml of an aqueous solution of HCl (pH = 2.0) is mixed with 300 ml of an aqueous solution of NaOH (pH = 12.0)? **[JEE-1998]**

Q.48 The pH of 0.1 M solution of the following salts increases in the order

(1) NaCl < NH_4Cl < NaCN < HCl

(2) HCl < NH_4Cl < NaCl < NaCN

(3) NaCN < NH_4Cl < NaCl < HCl

(4) HCl < NaCl < NaCN < NH_4Cl

Q.49 A buffer solution can be prepared from a mixture of

(1) Sodium acetate and acetic acid in water

(2) Sodium acetate and hydrochloric acid in water

(3) ammonia and ammonium chloride in water

(4) ammonia and sodium hydroxide in water.

Q.50 The solubility of $Pb(OH)_2$ in water is 6.7×10^{-6} M. Calculate the solubility of $Pb(OH)_2$ in in a buffer solution of pH = 8. **[JEE-1999]**

Q.51 The average concentration of SO_2 in the atmosphere over a city on a certain day is 10 ppm, when the average temperature is 298 K. Given that the solubility of SO_2 in water at 298 K is 1.3653 moles litre^{-1} and the pKa of H_2SO_4 is 1.92, extimate the pH of rain on that day. **[JEE 2000]**

[Given : $10^{-1.92} = 1.2 \times 10^{-2}$, $\sqrt{5.5678} = 2.5627$, log(1.2213) = 0.08668]

Q.52 For sparingly soluble salt ApBq. The relationship of its soubility product (L_s) with its solubiity (S) is – **[JEE 2001]**

(1) $L_s = S^{p+q}.p^p.q^q$

(2) $L_s = S^{p+q}.p^p.q^p$

(3) $L_s = S^{pq}.p^p.q^q$

(4) $L_s = S^{pq}.(p.q)^{p+q}$

Q.53 500 ml of 0.2 M aqueous solution of acetuc acd is mixed with 500 mL of 0.2 M HCl at 25°C.

(a) Calculate the degree of dissociation of acetic acid in the resulting solution and pH of the solution.

(b) If 6 g of NaOH is added to the above solution, determine final pH. Assume there si no chage in volume on mixing. K_a of acetic acid is 1.75×10^{-5} M. **[JEE-2002]**

Q.54 A solution which is 10^{-3} M each in Mn^{2+}, Fe^{2+}, Zn^{2+} and Hg^{2+} is treated with 10^{-16} M sulphide ion. If K_{sp}, MnS, FeS, ZnS and HgS are 10^{-15}, 10^{-23}, 10^{-23}, and 10^{-54} respectively. Which one will precipitate first ? **[JEE 2003]**

(1) FeS (2) MnS

(3) HgS (4) ZnS

Q.55 Will the pH of water be same at 4°C and 25°C? Explain **[JEE 2003]**

Q.56 0.1 M of HA is titrated with 0.1 M NaOH, calculate the pH at end point. Given Ka(HA) = 5 × 10^{-6} and $\alpha << 1$. **[JEE 2004]**

Q.57 HX is a weak acid ($K_a = 10^{-5}$). It forms a salt NaX (0.1 M) on reacting with caustic soda. The degree of hydrolysisof NaX is **[JEE 2004]**

(1) 0.01% (2) 0.0001%

(3) 0.1% (4) 0.5%

Q.58 Ch_3NH_2(0.1 mole, $K_b = 5 \times 10^{-4}$) is added to 0.08 moles of HCl and the solution is diluted to one oitrem resulting hydrogen ion concentration is

[JEE 2005]

(1) 1.6×10^{-11} (2) 8×10^{-11}

(3) 5×10^{-5} (4) 2×10^{-2}

Q.59 if $Ag^+ + NH_3 \rightleftharpoons [Ag(NH_3)]^+$; $K_1 = 1.6 \times 10^3$ and

$[Ag(NH_3)]^3 + NH_3 \rightleftharpoons [Ag(NH_3)_2]^+$; $K_2 = 6.8 \times 10^3$

The formation constant of $[Ag(NH_3)_2]^+$ is :

[JEE 2006]

(1) 6.08×10^{-6} (2) 6.8×10^{-6}

(3) 1.6×10^3 (4) 1.088×10^7

Q.60 The species present in solution when CO_2 is dissoved in water :

(1) CO_2, H_2CO_3, HCO_3^-, CO_3^{2-}

(2) H_2CO_3, CO_3^{2-}

(3) CO_3^{2-}, HCO_3^-

(4) CO_2, H_2CO_3

Q.61 2.5 mL of $\frac{2}{5}$ M weak monoacidic base ($K_b = 1 \times 10^{-12}$ at 25°C) is titrated with $\frac{2}{15}$ M HCl in water at 25°C. The concentration of H^+ at equivalence point is **[JEE 2008]**

($K_w = 1 \times 10^{-14}$ at 25°C)

(1) 3.7×10^{-13} M

(2) 3.2×10^{-7} M

(3) 3.2×10^{-2} M

(4) 2.7×10^{-2} M

Q.62 Solubility product constants (K_{SP}) of salts of types MX, MX_2 and M_3X at temperature 'T' ar 4.0×10^{-8}, 3.2×10^{-14} and 2.7×10^{-15}, respectively. Solubilities (mol dm^{-3}) of the salts at temperature 'T' are in the order : **[JEE - 2008]**

(1) $MX > MX_2 > M_3X$

(2) $M_3X > MX_2 > MX$

(3) $M_2X > M_3X > MX$

(4) $MX > M_3X > MX_2$

Q.63 The dissociation constant of a substituted benzoic acid at 25°C is 1.0×10^{-4}. The pH of a 0.01 M solution of its sodium salt is

[JEE- 2009]

Q.64 Aqueous solutions of HNO_3, KOH_4 CH_3COOH and CH_3COONa of identical concentrations are provided. The pair(s) of solutions which form a buffer upon mixing is(are) -

(1) HNO_3 and CH_3COOH

(2) KOH and CH_3COONa

(3) HNO_3 and CH_3COONa

(4) CH_3COOH and CH_3COONa

Q.65 In 1 L saturated solution of AgCl [K_{sp}(AgCl) = 1.6 × 10^{-10}], 0.1 mol of CuCl [K_{sp}(CuCl) = 1.0×10^{-6}] is added. The resultant concentration of Ag^+ in the solution is 1.6×10^{-8}. The value of 'x' is.

[JEE - 2011]

Q.66 The initial rate of hydrolysis of methyl acetate (1M) by a weak acid (Ha, 1M) is 1/100th of that of a strong acid (HX, 1M), at 25°C. the K_a of HA is

[JEE 2013]

(1) 1×10^{-4} (2) 1×10^{-5}

(3) 1×10^{-6} (4) 1×10^{-3}

Q.67 The K_{sp} of Ag_2CrO_4 is is 1.1×10^{-12} at 298 K. the solubility (in mol/L) of Ag_2CrO_4 in 0.1 M $AgNO_3$ solution is **[JEE 2013]**

(1) 1.1×10^{-11} (2) 1.1×10^{-10}

(3) 1.1×10^{-12} (4) 1.1×10^{-9}

Paragraph For Questions 68 and 69

When 100 mL of 1.0 M HCl was mixed with 100 mL of 1.0 M NaOH in an insulated beaker at constant pressure, a temperature increase of 5.7°C was measured for the beaker and its contents.

(Expt-1). Because the enthalpy of neutralization of a strong acid with a strong base is a constant (-57.0 kJmol^{-1}), this experiment could be used to measure the calorimeter constant. In a second experiment (Expt-2), 100 mL of 2.0 M acetic acid ($K_a = 2.0 \times 10^{-5}$) was mixed with 100 mL of 1.0 M NaOH (under identical conditions to (Expt-1)) where a temperature rise of 5.6°C was measured.

(Consider heat capacity of all solutions as 4.2 Jg^{-1}K^{-1} and density of all solutions as 1.0 g mL^{-1})

Q.68 Enthalpy of dissociation (in kJ mol^{-1}) of acetic acid obtained from the Expt-2 is **[JEE2015]**

(1) 1.0 (2) 10.0 (3) 24.5 (4) 51.4

Q.69 The pH of the solution after Expt-2

(1) 2.8 (2) 4.7 (3) 5.0 (4) 7.0

Q.70 The solubility of a salt of weak acid (AB) t pH 3 is $Y \times 10^{-3}$ mol L^{-1}. The value of Y is ____ .

(Given that the value of soubility product of AB(K_{sp}) = 2×10^{-10} and the value of ionization constant of HB(K_a) = 1×10^{-8}) **[JEE - 2018]**

ANSWER KEY

EXERCISE-I

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**
Ans.	2	4	2	3	4	4	3	1	2	4	1	2	4	4	3
Que.	**16**	**17**	**18**	**19**	**20**	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**
Ans.	4	4	4	1	1	1	1	3	4	1	2	3	4	1	1
Que.	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**	**41**	**42**	**43**	**44**	**45**
Ans.	1	2	1	4	4	4	2	3	4	2	3	3	3	3	3
Que.	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	2	3	3	4	2	4	4	4	2	3	2	1	1	3	3
Que.	**61**	**62**	**63**	**64**	**65**	**66**	**67**	**68**	**69**	**70**	**71**	**72**	**73**	**74**	**75**
Ans.	3	4	4	1	3	4	1	2	3	2	3	2	2	4	2
Que.	**76**	**77**	**78**	**79**	**80**	**81**	**82**	**83**	**84**	**85**	**86**	**87**	**88**	**89**	**90**
Ans.	2	4	3	3	1	2	2	3	4	1	1	1	4	4	2
Que.	**91**	**92**	**93**	**94**	**95**	**96**	**97**	**98**	**99**	**100**	**101**	**102**	**103**	**104**	**105**
Ans.	4	3	1	1	2	2	3	3	3	1	2	2	4	2	2
Que.	**106**	**107**	**108**	**109**	**110**	**111**	**112**	**113**	**114**	**115**	**116**	**117**	**118**	**119**	**120**
Ans	1	2	2	2	2	1	1	3	2	1	2	3	1	3	1
Que.	**121**	**122**	**123**	**124**	**125**	**126**	**127**	**128**	**129**	**130**	**131**	**132**	**133**	**134**	**135**
Ans.	4	4	2	2	1	2	3	4	3	3	1	1	2	3	2
Que.	**136**	**137**	**138**	**139**	**140**	**141**	**142**	**143**	**144**	**145**	**146**	**147**	**148**	**149**	**150**
Ans.	3	1	1	2	2	1	2	2	1	3	3	2	4	3	3
Que.	**151**	**152**	**153**	**154**	**155**	**156**	**157**	**158**	**159**	**160**	**161**	**162**	**163**	**164**	**165**
Ans.	2	4	3	2	2	1	3	1	1	3	4	3	2	1	4
Que.	**166**	**167**	**168**	**169**	**170**	**171**	**172**	**173**	**174**	**175**	**176**	**177**	**178**	**179**	**180**
Ans.	2	1	2	2	4	3	3	2	3	3	4	2	2	4	3
Que.	**181**	**182**	**183**	**184**	**185**										
Ans.	3	3	1	1	2										

EXERCISE-II

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**
Ans.	1	1	2	1	3	2	3	2	2	1	4	4	2	2	2
Que.	**16**	**17**	**18**	**19**	**20**	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**
Ans.	4	1	3	4	3	1	3	1	4	2	4	4	2	3	4
Que.	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**	**41**	**42**	**43**	**44**	**45**
Ans.	2	1	3	2	2	3	2	2	3	1	2	4	4	4	2
Que.	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	4	1	1	2	2	2	2	3	3	1	2	4	1	3	1
Que.	**61**	**62**	**63**	**64**	**65**	**66**	**67**	**68**	**69**	**70**	**71**	**72**	**73**		
Ans.	3	1	2	1	4	2	4	3	3	1	3	2	2		

EXERCISE-III

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**
Ans.	3	3	4	1	3	2	3	4	3	4	3	3	1	4	4
Que.	**16**	**17**	**18**	**19**	**20**	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**
Ans.	4	1	4	1	1	4	3	2	3	3	1	2	1	2	1
Quo.	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**	**41**	**43**	**44**	**45**	**46**
Ans.	2	2	1	3	3	3	4	1	2	4	10.6	3	2	2	1
Que.	**48**	**49**	**52**	**54**	**56**	**57**	**58**	**59**	**60**	**61**	**62**	**63**	**64**	**65**	**66**
Ans.	2	1,2,3	1	3	9	1	2	4	1	4	4	8	3, 4	7	1
Que.	**67**	**68**	**69**	**70**											
Ans.	2	1	2	4.47											

Q.42. 5.22 to 5.24

Q.47 Ph = 11.3010

Q.50 S = 1.203×10^{-3} M

Q.51 pH = 0.91325

Q.53 (a) = 0.0175% (b) = 4.757

Q.55 No, it will be > 7 at 0°C.

JEE Module Details

(Total = 24)

CLASS - XII : 12 MODULES

PHYSICS

Module - 1

Ch. No.	Chapter Name
1.	Electrostatics
2.	Capacitor & R-C Circuit
3.	Current Electricity

Module - 2

Ch. No.	Chapter Name
1.	MEC
2.	Magnetic Materials
3.	Bar Magnets & Earth Magnetism
4.	EMI
5.	AC
6.	EMW

Module - 3

Ch. No.	Chapter Name
1.	Ray Optics
2.	Wave Optics

Module - 4

Ch. No.	Chapter Name
1.	Modern Physics
2.	Nuclear Physics
3.	Electronics - Semiconductor
4.	Principles of Communication System

CHEMISTRY

Module -1 (Physical)

Ch. No.	Chapter Name
1.	The Solid State
2.	Solutions
3.	Electrochemistry
4.	Chemical Kinetics
5.	Surface Chemistry

Module -2 (Inorganic)

Ch. No.	Chapter Name
1.	The p -Block Elements
2.	General Principles and Processes of Isolation of Elements (Metallurgy)
3.	The d - and f Block Elements
4.	Coordination Compounds

Module -3 (Organic)

Ch. No.	Chapter Name
1.	Halogen Derivatives
2.	Oxygen Containing Compound
3.	Nitrogen Containing Compound
4.	Biomolecules, Polymers & Chemistry Every Day Life

MATHEMATICS

Module - 1

Ch. No.	Chapter Name
1.	Functions
2.	Inverse Trigonometric Functions
3.	Matrix
4.	Determinants

Module - 2

Ch. No.	Chapter Name
1.	Limit
2.	Continuity & Differentiability
3.	MOD
4.	AOD

Module - 3

Ch. No.	Chapter Name
1.	Integration
2.	Area Under Curve
3.	Differential Equations

Module - 4

Ch. No.	Chapter Name
1.	Vectors
2.	3 - Dimensional Geometry
3.	Probability

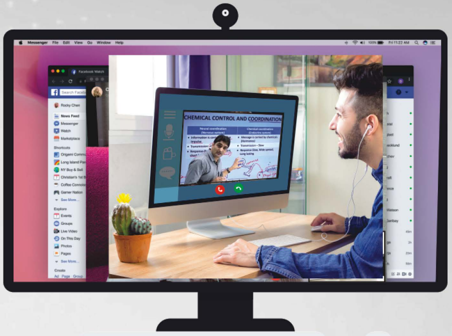
Module - 5

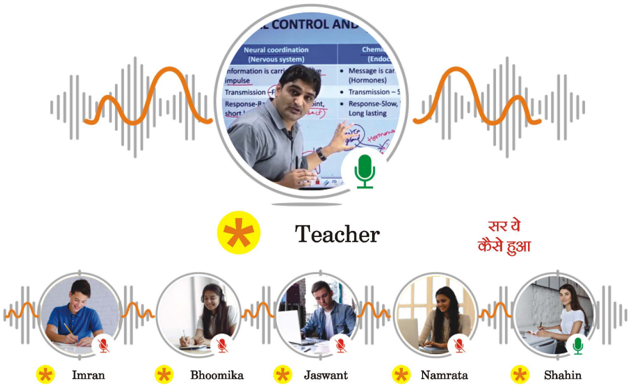
Ch. No.	Chapter Name
1.	H & D
2.	M. Reasoning
3.	Linear Programing
4.	Statistics

Live Classes for NEET, JEE & NTSE

Face to Face

Interaction with Teacher





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