

CHEMICAL EQUILIBRIUM

IRREVERSIBLE REACTION

Reactions which proceed in such a way that reactants are completely converted (i. e. 100%) in to products are called irreversible reaction.

Characteristics of Irreversible Reaction :

- **(i)** Reactions moves in one direction i.e. in forward direction only i.e. from left to right.
- **(ii)** Products do not react together to produce reactants again i.e. no backward reaction.
- **(iii)** Always carried out in open vessel.
- (iv) In such reactions arrow (\rightarrow) is placed between reactant and product which shows direction of chemical change. [Reactant \rightarrow Products]

Examples of Irreversible Reaction :

(i) Thermal decomposition reaction :

$$
2 KClO3(s) \xrightarrow{\text{MnO}_2} 2 KCl(s) + 3O2(g)
$$

(decomposition of potassium chlorate)

(ii) Precipitation reaction :

$$
AgNO3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO3(aq)
$$

white opt.

(iii) Neutralisation reaction : Reaction between strong acid and a strong base to give salt and water is called neutralisation reactions.

$$
H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O
$$
 (ii)
Strong Acid Strong Base Salt Water

(iv) Redox and combustion reactions :
\n
$$
SnCl2 + 2FeCl3 \rightarrow SnCl4 + 2FeCl2
$$
\n
$$
2Mg + O2 \longrightarrow 2MgO
$$

REVERSIBLE REACTION

The chemical reactions which take place in both directions
i.e. in the forward direction as well as in backward direction (B) i.e. in the forward direction as well as in backward direction under similar conditions are called reversible reactions.

Forward direction \implies Left to right

Backward direction \Rightarrow Right to left $\left(\rightleftarrow)$

Reactions which proceed in both the directions and does not reach to completion are called reversible reactions.

Characteristics of Reversible Reaction :

- **(i)** In these reactions arrow (\rightarrow) is replaced by two half arrows (\rightleftharpoons) pointing the reaction in both the directions.
	- Reactants \rightleftharpoons Products
- **(ii)** Always carried out in closed vessel.
- **(iii)** The reaction never reaches to completion.
- **(iv)** Characterised by the establishment of equilibrium.
- **(v)** Reactions occurs in both the directions.

Examples of Reversible Reactions :

- **(i) Reaction between gases in closed vessel :** $2\,\mathrm{SO}_2(g)+\mathrm{O}_2(g) \rightleftharpoons 2\,\mathrm{SO}_3(g)+x\,\mathrm{k.cal}$ $N_2(g) + 3 H_2(g) \rightleftharpoons 2NH_3(g) + x k.$ cal **(ii) Thermal decomposition reactions :**
	- $CaCO_3^2(s) \rightleftharpoons CaO(s) + CO_2^2(s) + x k.$ cal $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$

TYPES OF REVERSIBLE REACTIONS

(A) Homogeneous Reactions : The reversible reaction in which only one phase is present i.e. all the reactants and the products are in same physical state. **Example :** $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $CH_3COOH(\ell) + C_2H_5OH(\ell) \rightleftharpoons CH_3COOC_2H_5(\ell) + H_2O(\ell)$

Homogeneous reactions are of three types.

(i) First Type : When there is no change in number of molecules ($\Delta n = 0$) Example: $(g) + I_2(g) \rightleftharpoons 2HI(g)$

$$
2\,\text{NO}(g) \rightleftharpoons \text{N}_2(g) + \text{O}_2(g)
$$

(ii) Second Type : When there is increase in number of molecules. $(\Delta n > 0)$ Example: (g)

$$
PCl5(g) \rightleftharpoons PCl3(g) + Cl2(g)
$$

2NH₃(g) \rightleftharpoons N₂(g) + 3H₂(g)

(iii) Third Type : When there is decrease in number of molecules. $(\Delta n < 0)$ Example :

$$
N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)
$$

2SO₂(g) + O₂(g) \rightleftharpoons 2SO₃(g)

Heterogeneous Reactions : The reversible reaction in which more than one phases is present.

Example:
$$
MgCO_3(s) \rightleftharpoons MgO(s) + CO_2(g)
$$

3Fe(s) + 4 H₂O(l) \rightleftharpoons Fe₃O₄(s) + 4 H₂(g)

EQUILIBRIUM

Equilibrium is the state at which the concentration of reactants and products do not change with time i.e., concentrations of reactants and products become constant. At equilibrium state:

- **(i)** Rate of forward reaction (R_f) =Rate of backward reaction (R_b)
- **(ii)** The stage of the reversible reaction at which the concentrations of the reactants and products do not change with the time.

Characteristics of Equilibrium State :

- * Equilibrium state can only be achieved if a reversible reaction is carried out in closed vessel.
- Free energy change is zero $\Delta G^{\circ} = 0$.
- Equilibrium is dynamic in nature.
- Chemical Equilibrium can be attained from either of the side i.e. from side of reactants or products.

$$
2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)
$$

or
$$
H_2(g) + I_2(g) \rightleftharpoons 2 \text{ HI}(g)
$$

Types of equilibrium :

Equilibrium in a system implies the existence of the following types of equilibrium simultaneously.

- **1. Thermal equilibrium:** There is no flow of heat from one part to another i.e. $T = constant$.
- **2. Mechanical equilibrium:** There is no flow of matter from one part to another i.e. $P = constant$.
- **3. Physical equilibrium:** It is a state of equilibrium between the same chemical species in different phases (solid, liquid and gaseous)
- **4. Chemical equilibrium:** There is no change in composition of any part of the system with time.

Physical equilibrium :

Thevarious equilibrium which can exist in any physical system are,

 $Solid \rightleftharpoons Liquid$

- Liquid \rightleftharpoons Vapour
- Solid \rightleftharpoons Gas (vapour)

Solid \rightleftharpoons Saturated solution of solid in a liquid

Gas (vapour) \rightleftharpoons Saturated solution of gas in a liquid

(i) Solid-liquid equilibrium :

 $H_2O(s) \rightleftharpoons H_2O (liquid)$

Rate of transfer of molecules from ice to water

= Rate of transfer of molecules from water to ice Rate of melting of ice= Rate of freezing of water

(ii) Liquid-vapour equilibrium : When vapour of a liquid exists in equilibrium with the liquid, then

Rate of vaporisation = Rate of condensation, H_2O (liquid) $\rightleftharpoons H_2O$ (vapour)

Conditions necessary for a liquid-vapour equilibrium

- **(a)** The system must be a closed system i.e., the amount of matter in the system must remain constant.
- **(b)** The system must be at a constant temperature.

- **(c)** The visible properties of the system should not change with time.
- **(iii) Solid-vapour equilibrium:** Certain solid substances on heating get converted directly into vapour without passing through the liquid phase. This process is called sublimation. The vapour when cooled, gives back the solid, it is called disposition.

Solid \rightleftharpoons Vapour

The substances which undergo sublimation are camphor, iodine, ammonium chloride etc.

For example, Ammonium chloride when heated sublimes.

NH₄Cl (solid)
$$
\frac{\text{Heat}}{\text{Cool}}
$$
 NH₄Cl (vapour)

(iv) Equilibrium between a solid and its solution:

When a saturated solution is in contact with the the solid solute, there exists a dynamic equilibrium between the solid and the solution phase.

Solid substance \rightleftharpoons Solution of the substance Example: Sugar and sugar solution. In a saturated solution, a dynamic equilibrium is established between dissolved sugar and solid sugar.

Sugar (solid) \rightleftharpoons Sugar (aqueous)

At the equilibrium state, the number of sugar molecules going into the solution from the solid sugar is equal to the number of molecules precipitating out from the solution, i.e., at equilibrium,

Rate of dissolution of solid sugar

= Rate of precipitation of sugar from the solution.

(v) Equilibrium between a gas and its solution in a liquid:

Gases dissolve in liquids. The solubility of a gas in any liquid depends upon the,

- (a) Nature of the gas and liquid.
- (b) Temperature of the liquid.
- (c) Pressure of the gas over the surface of the solution.

CHEMICAL EQUILIBRIUM

LAW OF CHEMICAL EQUILIBRIUM

Law of Mass Action : At a given temperature, the rate of reaction at a particular instant is proportional to the product of the active masses of the reactants at that instant raised to power which are numerically equal to their respective molecules in stoichiometric equation describing the reaction. **LIBRIUM**
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 EXECUTE: $K_P = \frac{R_1}{R_2} = \frac{R_f}{R_b} = \frac{K_f}{K_b}$

(equilibrium constant of a reaction is 20.0. At
 Reaction is α **CHALLET TO ALLET AND ACTES**
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For the reaction : $nA + mB \rightleftharpoons zC + wD$ According to law of mass action

$$
R_f \propto [A]^n [B]^m; \qquad R_b \propto [C]^z [D]^w
$$

\n
$$
R_f = K_1 [A]^n [B]^m; \qquad R_b \propto [C]^z [D]^w
$$

\nAt equilibrium : $R_f = R_b$
\n $K_1 [A]^n [B]^m = K_2 [C]^z [D]^w$
\n $K_1 [C]^z [D]^w$ [Product]

$$
\frac{R_1}{K_2} = \frac{[C_1 \quad [B_1]]}{[A]^n [B]^m} = \frac{[P_1 \quad [B_2]]}{[Reactant]}
$$

Finally K_c (equilibrium constant) =
$$
\frac{K_1}{K_2} = \frac{R_f}{R_b} = \frac{K_f}{K_b}
$$

Example 1 :

The equilibrium constant of a reaction is 20.0. At equilibrium, the rate constant of forward reaction is 10.0. Find the rate constant for backward reaction.

Sol. $K_c = 20$ $K_f = 10, K_b = ?$

$$
K_c = \frac{K_f}{K_b}
$$
 ; $20 = \frac{10}{K_b}$: $K_b = \frac{10}{20} = 0.5$

EXPRESSION OF K_c
Let 'a' moles of PCl₅ be taken in a closed vessel of volume 'v' litre. Suppose 'x' mole is dissociated at equilibrium.

According to law of mass action

$$
K_c = \frac{[Product]}{[Readerant]} \quad ; \quad K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} \qquad \qquad \text{Sol. (}
$$
\n
$$
K_c = \frac{\frac{x}{v} \times \frac{x}{v}}{\frac{a - x}{v}} = \frac{x^2}{(a - x)v} \qquad \qquad \text{Exam}
$$

when $a = 1$, x becomes degree of dissociation (fraction of one mole dissociated into the products)

 $K_c = \frac{x^2}{(1-x)v}$.

according to ostwald dilution law $v \propto 1/C$

Hence : $K_c = \frac{\alpha^2 C}{(1-\alpha)}$

EQUILIBRIUM CONSTANT INTERMS OF PRESSURE

It is denoted by
$$
K_{p}
$$

Partial pressure $=$ $\frac{\text{no. of moles}}{\text{Total moles}} \times \text{Total pressure}$

SPONADVANCED LEARNING

TANT IN TERMS OF PRESSURE

P

no.of moles

Total moles

PCl₅ \rightleftharpoons PCl₃ + Cl₂

1 0 0

m 1-x x x

x + x = 1 + x **Example:** $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ Initial moles no. of 1 0 0 Moles of equilibrium $1-x$ x x Total moles $1-x+x+x=1+x$

2 b b K R K K R K PCl5 1 x P P 1 x ; Cl2 ^x P P 1 x ; PCl3 ^x P P 1 x K^P ⁼ PCl Cl 3 2 PCl5 P .P P

Unit of equilibrium constant :

Unit of \overline{K}_c = (Concentration)^{Δn} Unit of $K_p = (atm)^{\Delta n}$
 $\Delta n =$ no. of moles of gaseous product – no. of moles of gaseous reactant

\mathbf{R} elation between $\mathbf{K}_{\mathbf{p}}$ & $\mathbf{K}_{\mathbf{c}}$

$$
Hy Kc (equilibrium constant) = $\frac{K_1}{K_2} = \frac{R_f}{R_b} = \frac{K_f}{K_b}$
\n
$$
Hy Kc (equilibrium constant of a reaction is 20.0. At\nequilibrium constant of a reaction is 20.0. At\nh = 0. of moles of gaseous product\n
$$
H = -\frac{10}{1000} + \frac{100}{1000} + \
$$
$$
$$

x **Example 2 :**

 $\frac{x}{v}$ $\frac{x}{v}$ **Example 2:**
The unit of K_C for the reaction : $2\text{SO}_3 \rightleftharpoons 2\text{SO}_2 + \text{O}_2$ (A) mole² lit⁻²

- (B) mole per litre (C) mole x litre (D) litre per mole
- 51 **Sol. (B).** Unit fo $K_c = [M]^{\Delta n}$, where $M = \text{mole/litre}$ Δn = Number of gaseous mole of product – Number of gaseous mole of reactant Here $\Delta n = (2 + 1) - 2 = 1$
	- \therefore unit of K_c = [mole/litre]¹ = mole per litre.

Example 3 :

 $\frac{a-x}{v}$ $\frac{a-x}{v}$ $\frac{x}{v}$ $\frac{x}{v}$ In which of the following reaction $K_p > K_c$ (A) N₂(g) + 3H₂(g) \rightleftharpoons 2NH₃(g) $(B) H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $(C) PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$ $(D) 2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ **Sol.** (D). Value of K_p is greater then K_c for those reaction in

which $\Delta n > 0$ Here for the reaction

2SO₃(g)
$$
\rightleftharpoons
$$
 2SO₂(g) + O₂(g); $\Delta n = 3 - 2 = 1 > 0$
Therefore K_p > K_c for this reaction

Example 4 :

At 1000K, the value of K_p for the reaction : $\text{PCI}_5 \rightleftharpoons \text{PCI}_3 + \text{CI}_2$ is 20R, then find K_c. . **Sol.** $K_p = 20R$ (given) for the reaction $\Delta n = (2 - 1) = 1$ $K_p = K_c (RT)^{\Delta n}$ $20 R = K_c (RT)^1$ $\therefore K_c = \frac{2046}{R \times 1000} = 0.02$ STUDY!

STUDY!

OK, the value of K_p for the reaction :
 $R_{C_1} = \frac{[H1]^2}{[H_2][I_2]}$
 $R = K_c (RT)^{\Delta n}$

For the equation of above reaction $\Delta n = (2-1) = 1$
 $R = K_c (RT)^{\Delta n}$
 $= \frac{20 R}{R \times 1000} = 0.02$
 EXISTICS OF EQUILIBRIUM C STUDY

STUDY

the value of K_p for the reaction :
 $\Rightarrow PCI_3 + CI_2$ is 20R, then find K_c.

(given)
 $K_{C_1} = \frac{[H1]^2}{[H_2][I_2]}$
 $K_e (RT)^{\Delta n}$

If the equation of above reaction $\Delta n = (2 - 1) = 1$
 $K_e (RT)^{\Delta n}$
 $\Rightarrow K_{\text{c}}(RT)^$

CHARACTERISTICS OF EQUILIBRIUM CONSTANT

- **(i)** The expression for equilibrium constant, K is applicable only when concentrations of the reactants and products have attained their equilibrium values and do not change with time.
- **(ii)** The value of equilibrium constant is independent of initial concentration of the reactants and product.
- **(iii)** Equilibrium constant has one unique value for a particular reaction represented by a balanced equation at a given temperature. Ibirium constant has one unique value for a particular

following reaction : A + B = C + D

ion represented by a balanced equation at a given

erature.

erature.

 $\frac{1}{2}$ A + $\frac{1}{n}$ B = C + D

erature.

 $\frac{1}{2$
- **(iv)** If reaction is performed in multiple steps $A \rightleftharpoons B$; overall reaction

A
$$
\rightleftharpoons
$$
 B ; overall reaction
\nStep-1 : A (g) \rightleftharpoons C (g) + D (g)
\nStep-2 : C (g) \rightleftharpoons E (g)
\nStep-3 : D (g) + E (g) \rightleftharpoons B (g)
\nA (g) \rightleftharpoons B (g) K_p then K_p = K_{p1} · K_{p2} · K_{p3}
\n**Example 5:**
\nA (g) \rightleftharpoons B (g) K_p then K_p = K_{p1} · K_{p2} · K_{p3}
\n**Example 5:**
\nA (g) \rightleftharpoons B (g) K_p then K_p = K_{p1} · K_{p2} · K_{p3}
\n**Example 5:**
\nA a given temperature, the equi
\nreactions : NO (g) + ½ O₂ (g) \rightleftharpoons 2 NO₂ (g) \rightleftharpoons NO₂ (g) \rightleftharpoons NO₂ (g)
\nIf K₁ is 4 × 10⁻³, then find K₂.
\nIf we take reaction 2HI \rightleftharpoons H₂ + I₂
\nThen, we write the value of equilibrium constant K_{C1},
\nK₁ = $\frac{[NO_2]}{[NO_2]^{\frac{1}{2}}}$ = 4 × 10⁻³
\nNow for reaction 2NO₂ (g) \rightleftharpoons 2N
\nNow, if we take

Step-3 : D (g) + E (g) \Rightarrow B (g) h

$$
A(g) \rightleftharpoons B(g) \quad K_p \text{ then } K_p = K_{p_1} \cdot K_{p_2} \cdot K_p
$$

Factors which influence the equilibrium constant:

(a) Mode of representation of reaction : If we take reaction $2HI \rightleftharpoons H_2 + I_2$

> Then, we write the value of equilibrium constant K_{C_1} , K_{C_2} for the above reaction as following :

$$
K_{C_1} = \frac{[H_2][I_2]}{[HI]^2}
$$
(1)

Now, if we take reaction $H_2 + I_2 \rightleftharpoons 2HI$

Then, we write the value of equilibrium constant $K_{C₂}$ for (c) above reaction as following

$$
K_{C_2} = \frac{[H1]^2}{[H_2][I_2]} = \frac{1}{K_{C_1}}
$$
(2)

(b) Stoichiometry of the reaction :

Method of writing the equation of the reversible reaction is called as stoichiometry of the reaction.

Now, we consider the formation of HI(g) by the combination of $H_2(g)$ and $I_2(g)$.

$$
H_2 + I_2 \xrightarrow{\Delta} 2HI
$$

The expression of its equilibrium constant is-

$$
K_{C_1} = \frac{[H1]^2}{[H_2][I_2]}
$$

If the equation of above reaction is written by following

STUDYMATERIAL: CHEMISTRY
\n
$$
K_{C_1} = \frac{[H1]^2}{[H_2][I_2]}
$$
\nIf the equation of above reaction is written by following\n
$$
\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons H1(g)
$$
\nThe expression for the equilibrium constant is -\n
$$
K_{C_2} = \frac{[H1]}{[H_2]^{1/2}[I_2]^{1/2}}
$$
\nOn the basis of comparing both the equilibrium constant equation. $K_{C_2} = \sqrt{K_{C_1}}$ or $(K_{C_1})^{1/2}$
\n**Note :** When we divide a reaction by a factor 'n' in the

The expression for the equilibrium constant is -

$$
K_{C_2} = \frac{[H1]}{[H_2]^{1/2} [I_2]^{1/2}}
$$

On the basis of comparing both the equilibrium constant

equation. $K_{C_2} = \sqrt{K_{C_1}}$ or $(K_{C_1})^{1/2}$

on the reactants and products

in constant is independent of initial

in constant is independent of initial

actants and do not change
 $\begin{array}{lll}\n\text{equation.} & \text{Re}_2 = \sqrt{K_1} & \text{or } (K_1) \\
\text{equation.} & \text{equation.} & \text{the value of new equilibrium constant} \\
\text{equation.} & \text{for the value$ **EXECUALLY CONSTANT**

The expression for equinorium constant is $K_{C_2} = \frac{[H]}{[H_2]^{1/2}} [1/2]^{1/2}$

the reactants and products

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 Note : W **STUDYMATERIAL: CHEMISTRY**
 $\frac{[HI]^2}{[H_2][I_2]}$

of above reaction is written by following

(g) + $\frac{1}{2}I_2(g) \rightleftharpoons HI(g)$

for the equilibrium constant is -
 $\frac{[HI]}{2I^{1/2}[I_2]^{1/2}}$

comparing both the equilibrium const **STUDYMATERIAL: CHEMISTRY**
 $C_1 = \frac{[H1]^2}{[H_2][I_2]}$

ation of above reaction is written by following
 $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \implies HI(g)$

ssion for the equilibrium constant is -
 $= \frac{[H1]}{[H_2]^{1/2}[I_2]^{1/2}}$

sis of com **STUDYMATERIAL: CHEMISTRY**
 $1 = \frac{[H1]^2}{[H_2][I_2]}$

on of above reaction is written by following
 $H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons HI(g)$

ion for the equilibrium constant is -
 $\frac{[H1]}{[H_2]^{1/2}[I_2]^{1/2}}$

of comparing both the **Note :** When we divide a reaction by a factor 'n' in the equation, the value of new equilibrium constant is equal to the root of n of the previous equilibrium constant. For Example, Suppose, the equilibrium constant for the following reaction : $A + B \rightleftharpoons C + D$ is K_1 then [HI]²

¹2][1₂]

above reaction is written by following
 $+\frac{1}{2}I_2(g) \rightleftharpoons HI(g)$
 $\frac{1}{2}I_2(1g) \rightleftharpoons HI(g)$
 $\frac{1}{2}I_2(1g)^{1/2}$

omparing both the equilibrium constant
 $=\sqrt{K_{C_1}}$ or $(K_{C_1})^{1/2}$

divide a react $\frac{1}{12}$
 $\frac{1}{2}$

(IHI]
 $\frac{1}{2}$
 $\frac{1}{2}$

(IHI]
 $\frac{1}{2}$

(IHI]
 $\frac{1}{2}$

(I₁)^{1/2}

divide a reaction by a factor 'n $\frac{1}{12}$
 $\frac{1}{12}$

we reaction is written by following
 $\frac{1}{2}(g) \rightleftharpoons H1(g)$

equilibrium constant is -
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$

aring both the equilibrium constant
 $\frac{1}{K_{C_1}}$ or $(K_{C_1})^{1/2}$

de a rea $\frac{1}{2}$ H₂ (g) + $\frac{1}{2}$ L₂(g) \rightleftharpoons HI (g)

The expression for the equilibrium constant is -

IET expression for the equilibrium constant is -
 $K_{C_2} = \frac{[H_1]}{[H_2]^{1/2}[I_2]^{1/2}}$

On the basis of comparing (g) + $\frac{1}{2}I_2(g) \rightleftharpoons H1(g)$

(g) + $\frac{1}{2}I_2(g) \rightleftharpoons H1(g)$

for the equilibrium constant is -
 $\frac{[HI]}{[I_2]^{1/2}}$

f comparing both the equilibrium constant
 $C_2 = \sqrt{K_{C_1}}$ or $(K_{C_1})^{1/2}$

f comparing both the equ When we divide a reaction by a factor 'n' in the

the value of new equilibrium constant is equal

of of or the previous equilibrium constant.

to for of the previous equilibrium constant.

mple, Suppose, the equilibrium c

for the reaction :
$$
\frac{1}{n}A + \frac{1}{n}B \rightleftharpoons \frac{1}{n}C + \frac{1}{n}D
$$

The value of the equilibrium constant K_2 is equal to

$$
n\sqrt{K_1}
$$
 or $(K_1)^{1/n}$; $K_2 = K_1^{1/r}$

2 **Example 5 :**

3 K_{p_2} At a given temperature, the equilibrium constant for the reactions : NO (g) + $\frac{1}{2}$ O₂ (g) \rightleftharpoons NO₂ (g) and $2 N O_2(g) \rightleftharpoons 2 N O(g) + O_2(g)$ are K_1 and K_2 respectively. If K₁ is 4×10^{-3} , then find K₂. ion, the value of new equilibrium constant is equal
root of n of the previous equilibrium constant.
xample, Suppose, the equilibrium constant for the
ving reaction : $\frac{1}{n}A + \frac{1}{n}B \rightleftharpoons C + D$ is K_1 then
reaction :

Sol. NO(g) +
$$
\frac{1}{2}
$$
 O₂(g) \rightleftharpoons NO₂(g)

$$
K_1 = \frac{[NO_2]}{[NO][O_2]^{1/2}} = 4 \times 10^{-3}
$$

Now for reaction $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$

$$
K_2 = \frac{[NO]^2[O_2]}{[NO_2]^2} = \frac{1}{K_1^2} = \frac{1}{(4 \times 10^{-3})^2} = 6.25 \times 10^4
$$

nece the equilibrium constant:

sentation of reaction :

sol. NO (g) + $\frac{1}{2}$ O₂(g) \rightleftharpoons Ne

sentation as following :

action as following :
 $K_1 = \frac{[NO_2]}{[NO_1 O_2]}^{1/2} = 4 \times$
 $K_1 = \frac{[NO_2]}{[NO_1 O_2]}^{1/2} = 4 \times$
 2: $C(g) \rightleftharpoons E(g)$
 $A(g) \rightleftharpoons B(g)$
 $A(g) \rightleftharpoons$ (3.41)
 (4.41)
 (2) take reaction $2H1 \Leftrightarrow H_2 + 1$

and we write the value of equilibrium constant K_{C_1} ,
 $K_1 = \frac{[NO_2]}{[NO][O_2]^{1/2}} = 4 \times 10^{-3}$
 $K_{C_1} = \frac{[NO_1][O_2]}{[HO][O_2]^2} = \frac{1}{K_1^2} = \frac{1}{(4 \times 10^{-3} \text{ C})^2}$
 $K_{C_2} = \frac{[HI]^2}{[HI]^2}$: $K_2 = K_1^{2.2}$

ature, the equilibrium constant for the
 $+\frac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$ and
 $O(g)+O_2(g)$ are K_1 and K_2 respectively.

hen find K_2 .
 $\rightleftharpoons NO_2(g)$
 $\left[\frac{1}{2} = 4 \times 10^{-3}\right]$
 $SO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$
 erature, the equilibrium constant for the

2) + $\frac{1}{2}$ O₂ (g) \rightleftharpoons NO₂ (g) and

NO (g) + O₂ (g) are K₁ and K₂ respectively.

then find K₂.

g) \rightleftharpoons NO₂(g)
 $\frac{1}{\sqrt{2}} = 4 \times 10^{-3}$
 $\frac{2NO_2(g)}{k_1$ reaction: $A + B \rightleftharpoons C + D1sK_1$ then

tetion: $\frac{1}{n}A + \frac{1}{n}B \rightleftharpoons \frac{1}{n}C + \frac{1}{n}D$

of the equilibrium constant K_2 is equal to
 $(K_1)^{1/n}$; $K_2 = K_1^{1/n}$

1 temperature, the equilibrium constant for the
 $NO(g) + \frac{1$ tion: $\frac{1}{n}A + \frac{1}{n}B \rightleftharpoons \frac{1}{n}C + \frac{1}{n}D$

of the equilibrium constant K_2 is equal to
 K_1)^{1/n}; $K_2 = K_1^{1/n}$

temperature, the equilibrium constant for the

NO (g) + $\frac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$ and
 \Rightarrow 2 NO $K_2 = K_1^{3/4}$

e, the equilibrium constant for the
 $O_2(g) \rightleftharpoons NO_2(g)$ and
 $+ O_2(g)$ are K_1 and K_2 respectively.

find K_2 .
 $NO_2(g)$
 4×10^{-3}
 $2(g) \rightleftharpoons 2NO(g) + O_2(g)$
 $= \frac{1}{K_1^2} = \frac{1}{(4 \times 10^{-3})^2} = 6.25 \times$ $\Rightarrow \frac{1}{n}C + \frac{1}{n}D$

anstant K₂ is equal to
 $\Rightarrow NO_2(g)$ and

are K₁ and K₂ respectively.
 $NO(g) + O_2(g)$
 $\frac{1}{(4 \times 10^{-3})^2} = 6.25 \times 10^4$

demperature favours the

sase in temperature favours the

forward reaction s **Temperature :** Increase in temperature favours the endothermic reaction and decrease in temperature favours the exothermic reaction for the forward reaction so for exothermic reactions, the value of K_c and K_p decrease with rise in temperature while for endothermic reactions, the value of K_c and K_p increases with rise in temperature. This type of variation in equilibrium constant with temperature given by Van't Hoff equation as follows : 55:

sagiven temperature, the equilibrium constant for the

agiven temperature, the equilibrium constant for the

tections: NO (g) + ½ O₂ (g) \rightleftharpoons NO₂ (g) and

K₁ is 4 × 10⁻³, then find K₂.

9(g) + $\frac{1}{2}$ $K_1^2 = (4 \times 10^{-3})^2 = 0.23 \times 10$
crease in temperature favours the
n and decrease in temperature favours
tion for the forward reaction so for
s, the value of K_c and K_p decrease
ture while for endothermic reactions,
 K_p 0^{-3}
 $\Rightarrow 2NO(g) + O_2(g)$
 $\frac{1}{2} = \frac{1}{(4 \times 10^{-3})^2} = 6.25 \times 10^4$

in temperature favours the

ecrease in temperature favours

or the forward reaction so for

value of K_c and K_p decrease

reases with rise in temperat (g) \Rightarrow 2NO(g) + O₂(g)
 $\frac{1}{K_1^2} = \frac{1}{(4 \times 10^{-3})^2} = 6.25 \times 10^4$

aase in temperature favours the

n for the forward reaction so for

the value of K_c and K_p decrease

e while for endothermic reactions,

increa 10⁻³
 \Rightarrow 2NO(g) + O₂(g)
 $\frac{1}{\zeta_1^2} = \frac{1}{(4 \times 10^{-3})^2} = 6.25 \times 10^4$

ie in temperature favours the

decrease in temperature favours

for the forward reaction so for

e value of K_c and K_p decrease

while for NO(g) + O₂(g)
 $\frac{1}{(4 \times 10^{-3})^2}$ = 6.25 × 10⁴

temperature favours the

ase in temperature favours

e forward reaction so for

e of K_c and K_p decrease

for endothermic reactions,

s with rise in temperature.

u NO(g) + O₂(g)
 $\frac{1}{(4 \times 10^{-3})^2}$ = 6.25 × 10⁴

temperature favours the

ase in temperature favours

e forward reaction so for

e of K_c and K_p decrease

for endothermic reactions,

s with rise in temperature.

u

$$
\frac{\Delta H}{RT^2} = \frac{d \left(\ln K \right)}{dT}
$$

After integration,

$$
\log K_2 - \log K_1 = \frac{\Delta H}{2.303R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]
$$

v

or
$$
\log \frac{K_2}{K_1} = \frac{\Delta H}{2.303R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]
$$

EQUILIBRIUM
 $\frac{2}{1} = \frac{\Delta H}{2.303R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$

(b) Intercept 'c' = $\log_{10} A = 1$

(c) $\log K = 10 - \frac{9.5745}{2.303 \times 8.31}$

(e) $\log K = 10 - \frac{9.5745}{2.303 \times 8.31}$

(e) $\log K = 10 - \frac{9.5745}{2.303 R}$

(e) $\log K = 1$ CAL EQUILIBRIUM
 $\log \frac{K_2}{K_1} = \frac{\Delta H}{2.303R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$

(b) Intercept 'c' = $\log_{10} A = 10$.: A = 10^{10}

(c) $\log K = 10 - \frac{9.574}{2.303 \times 8.314 \times 298}$; K = 9

re, K₂ = equilibrium constant at temperatur **UILIBRIUM**
 $= \frac{\Delta H}{2.303R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$

(b) Intercept 'c' = $\log_{10} A = 10$.: A = 10^{10}

(c) $\log K = 10 - \frac{9.574}{2.303 \times 8.314 \times 298}$; K = 9.9

equilibrium constant at temperature T₂

(d) $\log \left(\frac{K_2}{K$ Where, K_2 = equilibrium constant at temperature T₂ K_1 = equilibrium constant at temperature T₁ ΔH = Energy of reaction of constant temperature $R =$ Molar gas constant

Slope =
$$
-\frac{\Delta H^{\circ}}{2.303 \text{ R}}
$$

According to the temperature, reaction are of three types.

- (i) Non-thermic reaction means $\Delta H = 0$ $\log K_2 - \log K_1 = 0$; $\log K_2 = \log K_1$
- There is no effect of temperature on this type of reaction. (ii) Endothermic reaction $\Rightarrow \Delta H = (+)$ ve
- $\log K_2 \log K_1 = (+)$ ve, means $K_2 > K_1$ On increasing of temperature, equilibrium constant will also increase for this type of reaction.
- (iii) Exothermic reaction $\Rightarrow \Delta H = (-)$ ve $\log K_2 - \log K_1 = (-) \text{ ve}, \text{ means } K_2 \le K_1$ On the increase of temperature equilibrium constant will decreases for exothermic reaction.

Example 6 :

Variation of equilibrium constant K with temperature T is

given by van't Hoff equation, $\log K = \log A - \frac{1}{2.202B}$.

A graph between log K and T^{-1} was a straight line as shown in the figure and having $\theta = \tan^{-1}(0.5)$ and OP = 10. Calcualte

-
- (b) A (pre-exponential factor)
- (c) Equilibrium constant K, at 300 K,
- (d) K at 900K if ΔH° is independent of temperature.

Sol. (a)
$$
\log_{10} K = \log_{10} A - \frac{\Delta H^{\circ}}{2.303RT}
$$

It is an equation of a straight line of the type $y = c + mx$

It is an equation of a straight line of the type $y = c + mx$

Slope m = tan
$$
\theta = \frac{\Delta H^{\circ}}{2.303R}
$$

\n
$$
0.5 = \frac{\Delta H^{\circ}}{2.303 \times 8.314}; \ \Delta H^{\circ} = 9.574 \text{ J mol}^{-1}
$$
\nK =

(b) Intercept 'c' = $log_{10}A = 10$: $A = 10^{10}$

(b) Intercept 'c' =
$$
\log_{10} A = 10
$$
 $\therefore A = 10^{10}$
\n(c) $\log K = 10 - \frac{9.574}{2.303 \times 8.314 \times 298}$ $\therefore K = 9.96 \times 10^9$
\n(d) $\log \left(\frac{K_2}{K_1}\right) = \frac{\Delta H}{2.303R} \left\{ \frac{1}{T_1} - \frac{1}{T_2} \right\}$
\n $\log \frac{K_2}{9.96 \times 10^9} = \frac{9.574}{2.303 \times 8.314} \left\{ \frac{1}{298} - \frac{1}{798} \right\}$

(d)
$$
\log \left(\frac{K_2}{K_1}\right) = \frac{\Delta H}{2.303R} \left\{ \frac{1}{T_1} - \frac{1}{T_2} \right\}
$$

Intercept 'c' =
$$
log_{10}A = 10
$$
 \therefore A = 10¹⁰
\n $log K = 10 - \frac{9.574}{2.303 \times 8.314 \times 298}$; K = 9.96 × 10⁹
\n $log(\frac{K_2}{K_1}) = \frac{\Delta H}{2.303R} \left\{ \frac{1}{T_1} - \frac{1}{T_2} \right\}$
\n $log \frac{K_2}{9.96 \times 10^9} = \frac{9.574}{2.303R} \left\{ \frac{1}{298} - \frac{1}{798} \right\}$
\nOn solving, K₂ = 9.98 × 10⁹
\nBRUMCONSTANT EXPRESSIONS FOR REACTIONS
\nInmogeneous gaseous equilibria when (Δ n = 0)
\nintesis of hydrogen iodide
\nlculation of K : Suppose the initial concentration of

EQUILIBRIUM CONSTANT EXPRESSIONS FOR REACTIONS

Case I : Homogeneous gaseous equilibria when $(\Delta n = 0)$ **Synthesis of hydrogen iodide**

(a) Calculation of K^c : Suppose the initial concentration of H_2 and I_2 is 'a' and 'b' respectively, x is degree of dissociation and 'v' is volume in litre of container.

(c)
$$
\log K = 10 - \frac{9.5}{2.303 \times 8.314 \times 298}
$$
; $K = 9.96 \times 10^9$
\n T_2
\n(d) $\log \left(\frac{K_2}{K_1}\right) = \frac{\Delta H}{2.303R} \left\{ \frac{1}{T_1} - \frac{1}{T_2} \right\}$
\n $\log \frac{K_2}{9.96 \times 10^9} = \frac{9.574}{2.303 \times 8.314} \left\{ \frac{1}{298} - \frac{1}{798} \right\}$
\nOn solving, $K_2 = 9.98 \times 10^9$
\nEQUILIBRIUMCONSTANT EXPRESSIONS FOR REACTIONS
\nCase I : Homogeneous gaseous equilibria when ($\Delta n = 0$)
\nSynthesis of hydrogen iodide
\n(a) Calculation of K_c : Suppose the initial concentration of
\n H_2 and I_2 is 'a' and 'b' respectively, x is degree of
\ndissociation and 'v' is volume in litre of container.
\n $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
\neaction.
\nIndles at equilibrium $a - x$ b-x 2x
\nstart will
\nActive mass $\frac{a - x}{v} \frac{b - x}{v} \frac{2x}{v}$
\nAccording to law of mass action.
\nonstant $K_c = \frac{[HI]^2}{[H_2][I_2]}$; $K_c = \frac{[2x/v]^2}{[a-x][b-x]} = \frac{4x^2}{(a-x)(b-x)}$
\n(b) Calculation of K_p : The equilibrium constant K_p , can also
\nbe calculated considering partial pressure of reactants and products at equilibrium.
\n $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
\nand ΔH^o
\n ΔH^o
\nIn trial concentration a b b
\n $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
\n $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
\n ΔH^o
\n ΔH^o
\n ΔH^o
\n ΔH^o
\n

According to law of mass action.

$$
E = \frac{1}{k_1} \left(\frac{1}{2.303R} \right) \left(\frac{1}{11} - \frac{1}{12} \right)
$$
\n
$$
\log \frac{K_2}{9.96 \times 10^9} = \frac{9.574}{2.303 \times 8.314} \left\{ \frac{1}{298} - \frac{1}{798} \right\}
$$
\nOn solving, $K_2 = 9.98 \times 10^9$
\n**LIBRIUM CONSTANT EXPRESSIONS FOR REACTIONS**
\n1: Homogeneous gaseous equilibria when ($\Delta n = 0$)
\n**Synthesis of hydrogen iodide**
\n1: Homogeneous gaseous equilibria when ($\Delta n = 0$)
\nSynthesis of hydrogen iodide
\nH₂ and I₂ is 'a' and 'b' respectively, x is degree of
\ndissociation and 'v' is volume in litre of container.
\nInitial conc.
\n1: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
\nMoles at equilibrium $a - x b - x 2x$
\nActive mass
\n
$$
\frac{a - x}{v} \quad \frac{b - x}{v} \quad \frac{2x}{v}
$$
\nAccording to law of mass action.
\n
$$
K_c = \frac{[HI]^2}{[H_2][I_2]} \, ; K_c = \frac{[2x/v]^2}{\left[\frac{a - x}{v}\right] \left[\frac{b - x}{v}\right]} = \frac{4x^2}{(a - x)(b - x)}
$$
\n**Calculation of** K_p : The equilibrium constant K_p , can also
\nbe calculated considering partial pressure of reactants and
\nproducts at equilibrium.
\n
$$
H_2(g) + I_2(g) \rightleftharpoons 2HI(g)
$$
\nInitial concentration a b 0
\nMoles at equilibrium $a - x b - x 2x$
\nTotal number of moles at equilibrium
\n= (a - x) + (b - x) + 2x = (a + b)
\nLet the total pressure of the system at equilibrium be P
\natmosphere.
\nPartial pressure of H_2
\n
$$
P_{H_2} = \frac{(a - x)}{(a + b)} p
$$
\nPartial pressure of H_2
\n
$$
P_{H_1} = \left(\frac{2x}{a
$$

(b) Calculation of K_p **:** The equilibrium constant K_p , can also be calculated considering partial pressure of reactants and products at equilibrium.

 $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $2.303RT$ Initial concentration a b 0 Moles at equilibrium $a-x$ b – x 2x Total number of moles at equilibrium

> $= (a-x) + (b-x) + 2x = (a+b)$ Let the total pressure of the system at equilibrium be P atmosphere.

Partial pressure of H₂ $P_{\text{H}_2} = \frac{(\alpha - \lambda)}{(\alpha + \lambda)} P$

$$
P_{H_2} = \frac{(a-x)}{(a+b)} P
$$

Partial pressure of I_2

$$
P_{I_2} = \left(\frac{b-x}{a+b}\right)P
$$

Partial pressure of HI

$$
P_{\rm HI} = \left(\frac{2x}{a+b}\right) P
$$

the increase of temperature equilibrium constant
\ndecreases for exothermic reaction.
\nTo equilibrium constant K with temperature T is
\ndecreases for exothermic reaction.
\nTo the equation,
$$
\log K = \log A - \frac{\Delta H^{\circ}}{2.303KT}
$$
.
\nTo the figure and having $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo the figure and having $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo the figure and having $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo find the figure and having $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo find the figure and having $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo find the figure of the system at equilibrium $\theta = \tan^{-1} (0.5)$ and $OP = 10$.
\nTo find the difference of the system at equilibrium $\theta = \tan^{-1} (\theta - \theta - \theta)$ and $\theta = \tan^{-1} (\theta - \theta - \theta)$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium $\theta = \theta$ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the difference of the system at equilibrium θ and $\theta = \theta$ and $\theta = \theta$.
\nTo find the system of the system of the system of the system at equilibrium θ and $\theta = \theta$ and $\$

$$
K_p = \frac{4x^2}{(a-x)(b-x)}
$$

This also follows the relation $p = K_c (RT)^{\Delta n}$

Here, $\Delta n = 0$; $K_p = K_c$ i.e. for all gaseous reaction of this type, K_p and K_c are equal and have no units.

- **(i) Effect of pressure :** $K_c \propto p$ pressure⁰ As the equation does not include volume term, K_c is independent of pressure.
- **(ii) Effect of adding substances :** (Reactant and products) On adding reactant rate of forward reaction increase on adding product rate of backward reaction increase Concentration of reactant \uparrow Rate of forward reaction \uparrow Concentration of product \uparrow Rate of backward reaction \uparrow
- **(iii) Effect of temperature :** The formation of HI from H_2 and I_2 is an exothermic reaction i.e. by increasing temperature the value of K_c decreases. As $T \uparrow$, $K_c \downarrow$ (exothermic reaction)

Case II : Homogeneous gaseous reactions when $\Delta n > 0$ **Dissociation of PCl⁵**

(a) **Calculation of** K_c **: Let 'a' moles of PCl₅ be taken in a** closed vessel of volume 'v' litre. Suppose 'x' mole is dissociated at equilibrium.

According to law of mass action

$$
K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} \; ; \; K_c = \frac{\frac{x}{v} \times \frac{x}{v}}{\frac{a - x}{v}} = \frac{x^2}{(a - x)v}
$$

when $a = 1$, x becomes degree of dissociation

$$
K_c = \frac{x^2}{(1-x)v}
$$

(b) Calculation of K_p **:** Let the total pressure at equilibrium be P. Then total number of moles at equilibrium

$$
= (a-x) + x + x = a + x
$$

The partial pressure of each reactant and product

$$
p_{PCl_5} = \left(\frac{a-x}{a+x}\right)P; p_{Cl_2} = \left(\frac{x}{a+x}\right)P; p_{PCl_3} = \left(\frac{x}{a+x}\right)P \qquad \textbf{(b)} \qquad \textbf{Calc}
$$

Applying law of mass action

At equilibrium

\n7. A
$$
a - x
$$

\n7. A $a - x$

\n8. A $a - x$

\n9. A $a - x$

\n10. A $a - x$

\n21. A $a - x$

\n33. A $a - x$

\n4. A $a - x$

\n5. A $a - x$

\n6. A $a - x$

\n7. A $a - x$

\n8. A $a - x$

\n9. A $a - x$

\n10. A $a - x$

\n11. A $a - x$

\n12. A $a - x$

\n13. A $a - x$

\n14. A $a - x$

\n15. A $a - x$

\n16. B $a - x$

\n17. A $a - x$

\n18. B $a - x$

\n19. A $a - x$

\n10. B $a - x$

\n11. B $a - x$

\n12. A $a - x$

\n13. A $a - x$

\n14. A $a - x$

\n15. A $a - x$

\n16. A $a - x$

\n17. A $a - x$

\n18. A $a - x$

\n19. A $a - x$

\n10. A $a - x$

\n11. A $a - x$

\n12. B $a - x$

\n13. A <math display="inline</p>

when $a = 1$, $K_p = \frac{x^2 P}{1 r^2}$ 2 (i) **Effect of pressure :** $K_p = \frac{x^2 P}{1 + x^2}$ 2

Suppose $x \ll 1$ then $1 - x^2 = 1$

STUDYMATERIAL: CHEMISTRY
\nEffect of pressure :
$$
K_p = \frac{x^2 P}{1 - x^2}
$$

\nSuppose x <<1 then $1 - x^2 = 1$
\n $K_p = x^2 P$; $x^2 \propto \frac{1}{P}$; $x \propto \frac{1}{\sqrt{P}}$ $P \downarrow$, $x \uparrow$
\nDegree of dissociation of PCl is inversely

MATERIAL: CHEMISTRY
 $\frac{x^2 P}{1-x^2}$
 $x^2 = 1$
 $\frac{1}{\sqrt{P}} P \downarrow, x \uparrow$

n of PCl₅ is inversely

t of pressure so decrease in Degree of dissociation of PCl_5 is inversely proportional to square root of pressure so decrease in pressure increase dissociation of PCl_5

(ii) Effect of concentration :

Concentration of PCl₅ \uparrow Rate of forward reaction \uparrow Concentration of PCl₃ & Cl₂ TRate of backward reaction T

(iii) **Effect of temperature :** The dissociation of PCl_5 is endothermic reaction. The value of equilibrium constant increase with increase in temperature $K \alpha T$ As $T \uparrow$, $K \uparrow$

Case III : Homogeneous gaseous reaction $\Delta n < 0$ **Synthesis of Ammonia**

(a) **Calculation of** K_c **:**

$$
\frac{x}{x} \qquad \qquad \text{Active mass} \quad \frac{a-x}{v} \qquad \qquad \frac{b-3x}{v} \qquad \qquad \frac{2x}{v}
$$

v v Applying law of mass action

On adding reactant rate of forward reaction increase
\non adding product rate of backward reaction increase
\nConcentration of reactant. Take of forward reaction
\nConcentration of reactant. The formation of H1 from H₂
\nConcentration of PCI₃ (a) 1
\nConcentration of PCI₃ (b) 1
\nConcentration of PCI₃ (c) 1
\nConcentration of PCI₃ (d) 1
\nand L₃ is an cothermic reaction
\nand L₃ is an cothermic reaction
\nand L₃ is an cothermic reaction
\n
$$
As T T, K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 & 0 \\ 0 & 0 & 0 \\ 0 & 0 & 0 \end{pmatrix}
$$
\n
$$
K_{\varphi} = \begin{pmatrix} 1 & 0 &
$$

If a = 1 and b = 3 then K_c =
$$
\frac{4x^2v^2}{27(1-x)^4}
$$

If x << 1 then $(1-x)^4 = 1$

$$
K_c = \frac{4x^2v^2}{27} \; ; \; x^2 \propto \frac{1}{v^2} \; ; \; x \propto \frac{1}{v^2}
$$

; $p_{\text{PCl}_3} = \left(\frac{P}{a + x}\right)P$ Let total pressure at equilibrium be P. $\left(\frac{\mathbf{x}}{\mathbf{p}}\right)$ **b Calculation of Kp**: Total no. of moles at equilibrium

$$
= (a-x) + (b-3x) + 2x = (a+b-2x)
$$

3

$$
p_{NH_3} = \frac{2x}{(a+b-2x)} \times P \ ; \ p_{N_2} = \frac{a-x}{(a+b-2x)} \times P
$$

$$
\frac{12_{2}}{1}; K_{c} = \frac{\frac{v}{x} \times \frac{v}{v}}{\frac{a - x}{v}} = \frac{x^{2}}{(a - x)v}
$$
\n
$$
K_{c} = \frac{[2x/v]^{2}}{[a - x]^{2}} = \frac{4x^{2}/v^{2}}{[a - x]^{3}}
$$
\n
$$
K_{c} = \frac{[2x/v]^{2}}{[a - x]^{2}} = \frac{4x^{2}/v^{2}}{[a - x]^{3}}
$$
\n
$$
Ha = 1 \text{ and } b = 3 \text{ then } K_{c} = \frac{4x^{2}v^{2}}{27(1 - x)^{4}}
$$
\n
$$
= \text{or f each reactant and product}
$$
\n
$$
= (a - x) + x + x = a + x
$$
\n
$$
K_{c} = \frac{4x^{2}v^{2}}{27}; x^{2} \propto \frac{1}{v^{2}}; x \propto \frac{1}{v^{2}}
$$
\n
$$
PC_{12} = \left(\frac{x}{a + x}\right)P; P_{P}C_{13} = \left(\frac{x}{a + x}\right)P
$$
\n
$$
= \text{for a ch reactant and product}
$$
\n
$$
K_{c} = \frac{4x^{2}v^{2}}{27}; x^{2} \propto \frac{1}{v^{2}}; x \propto \frac{1}{v^{2}}
$$
\n
$$
= \text{for a ch reactant and product}
$$
\n
$$
K_{c} = \frac{4x^{2}v^{2}}{27}; x \propto \frac{1}{v^{2}}; x \propto \frac{1}{v^{2}}
$$
\n
$$
= \text{for a ch. vector, } \text{or f }
$$

If a = 1 and b = 3 then
$$
K_p = \frac{16x^2(2-x)^2}{27(1-x)^4p^2}
$$

Suppose $x \ll 1$ then $(1-x)^4 = 1$; $(2-x)^2 = 4$

$$
K_p = \frac{64x^2}{27P^2}
$$
; $x^2 \propto P^2$; $x \propto P$

- **(i) Effect of pressure :** With increase in pressure the volume term decrease therefore x increase. $P \uparrow V \downarrow x \uparrow$
- **(ii) Effect of temperature :** On increasing temperature value of equilibrium constant decreases.

 $T \uparrow K \downarrow$ exothermic reaction

Case IV : Homogeneous liquid system :

Formation of ethyl acetate $CH_3COOH(\ell) + C_2H_5OH(\ell) \rightleftharpoons CH_3COOC_2H_5(\ell) + H_2O(\ell)$ Initial conc.

$$
\frac{a-x}{v} \qquad \qquad \frac{b-x}{v} \qquad \qquad \frac{x}{v} \qquad \qquad \frac{x}{v} \qquad \qquad \text{Sol.}
$$

Applying law of mass action

Here $V =$ total volume

value of equilibrium constant decreases.
\n
$$
T \uparrow K \downarrow
$$
 exothermic reaction
\n**Homogeneous liquid system :**
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; A}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; A}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; B}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; B}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; C}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}; [B] = 2 \text{ mole; D}$
\n $\therefore [B] = 2 \text{ mole; D}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [A] = 4 - 2 = 2 \text{ mole}$
\n $\therefore [B] = 2 \text{ MeV}$
\n $\therefore [C] = 2 \text{ MeV}$
\n $\therefore [C] = 2 \text{$

$$
K_c = \frac{x^2}{(a-x)(b-x)}
$$
. If $a = b = 1$ then $K_c = \frac{x^2}{(1-x)^2}$ **Example 10:**
One mole of ethyl alcohol

This does not involve volume term thus K_c is not affected by change in volume $K_c \propto V^0$

Case V : Heterogeneous Equilibrium

The decomposition of solid $NaHCO₃$ to produce solid Na_2CO_3 , CO_2 gas and H_2O . 2 NaHCO₃ (s) \rightleftharpoons Na₂CO₃(s) + CO₂(g) + H₂O(g)

Applying law of mass action

$$
K_c' = \frac{[\text{Na}_2\text{CO}_3][\text{H}_2\text{O}][\text{CO}_2]}{[\text{NaHCO}_3]^2}
$$

It involves two pure solids, $Na₂CO₃$ and $NaHCO₃$. It is customary not to include concentration of pure solid in equilibrium expression. bes not involve volume term thus K_e is not affected

age in volume $K_e \propto V^0$

composition of spidrolysis of

composition of spidrolysis of

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conposition of spidrolysis of

composition of s 2 3 3 and H₂O₃ 1

2 3 as and H₂O₃ 1

2 3 as and H₂O₃ (3) + CO₂(g) + H₂O(g)
 $\frac{1}{2}$ as and H₂O₃(s) + CO₂(g) + H₂O(g)
 $\frac{1}{2}$ as action
 $\frac{1}{2}$ (NaHCO₃¹)²

2 as action
 $\frac{1}{2}$ (NaH

$$
\frac{K_c [NaHCO_3]^2}{[Na_2CO_3]} = [CO_2][H_2O]
$$

\n
$$
K_c = [CO_2][H_2O]
$$

Example 7 :

For the reaction N₂O₄ \rightleftharpoons 2NO, degree of dissociation is S₆ α . Find the number of moles at equilibrium.

Sol.
 $=\frac{16x^2(2-x)^2}{27(1-x)^4P^2}$
 $=-x)^4=1$; $(2-x)^2=4$

Sol.

Sol.
 $\begin{array}{ccc}\n\text{Initially} & \frac{N_2O_4}{1} \rightleftharpoons 2NO \\
\text{Intially} & 1 & 0 \\
\text{Moles at equilibrium } 1-\alpha & 2\alpha \\
\therefore \text{ Total number of moles at equilibrium} \\
 = 1-\alpha+2\alpha=1+\alpha\n\end{array}$ Sol.
 $\frac{1}{4p^2}$
 $\frac{1}{4p^2}$
 $\frac{1}{4p^2}$
 $\frac{1}{4p^2}$
 $\frac{1}{2p^2}$
 $\frac{1$ **Sol.** $N_2O_4 \rightleftharpoons 2NO$
 $27(1-x)^4P^2$
 $4 = 1$; $(2-x)^2 = 4$
 Example 8:
 Example 8:
 Example 8:
 Example 6 A are mixed with 4 mole of B when 2 mole of G **Sol.** $\begin{array}{r} \text{Sol.} \\ \hline \text{DOMADVANIGED HERR} \\ 27(1-x)^4 P^2 \\ 4 = 1; (2-x)^2 = 4 \end{array}$ Sol. $\begin{array}{r} \text{N}_2O_4 \rightleftharpoons 2\text{NO} \\ \text{Initially} \\ \text{Moles at equilibrium } 1 - \alpha \\ \therefore \text{Total number of moles at equilibrium} \\ = 1 - \alpha + 2\alpha = 1 + \alpha \end{array}$
 $\times \propto P$ Example 8:
Total number of moles at equilibr Sol.
Initially $(-x)^4 P^2$ Moles at equilibrium $1 - \alpha$ 2 α **Sol.** $N_2O_4 \rightleftharpoons 2NO$ Initially 1 0 \therefore Total number of moles at equilibrium $= 1 - \alpha + 2\alpha = 1 + \alpha$

Example 8 :

4 mole of A are mixed with 4 mole of B, when 2 mole of C are formed at equilibrium according to the reaction, $A + B \rightleftharpoons C + D$. Find the equilibrium constant. **EXERCISE AND SURVE THE SURVEY OF A SURVEY ON A SURVEY OF A SURVE** N₂O₄ \rightleftharpoons 2NO
 $\frac{1}{1}$ 0

quilibrium 1 – α 2α
 $\alpha + 2\alpha = 1 + \alpha$

Aare mixed with 4 mole of B, when 2 mole of C are

equilibrium according to the reaction,

C+D. Find the equilibrium constant.

A + B C + D

4 4

Sol. A + B C + D Initially $4 \t 4 \t 0 \t 0$ At equilibrium $4-\alpha$ $4-\alpha$ α α Given at equilibrium mole of C = 2 (\therefore α = 2)

 \therefore [A] = 4 – 2 = 2 mole; [B] = 4 – 2 = 2 mole $[D] = \alpha = 2$ mole

$$
K_c = \frac{[C][D]}{[A][B]} = \frac{2 \times 2}{2 \times 2} = 1
$$

Example 9 :

x dissociated at equilibrium. Find the equilibrium constant. At a certain temperature, $2HI \rightleftharpoons H_2 + I_2$ only 50% HI is

$$
s \times x < t \times 0 \text{ the first order, and the probability of the number of moles at equilibrium 1-u to the number of moles at equilibrium 1-u to the number of moles at equilibrium constant. The number of moles at equilibrium constant, the number of moles at equilibrium constant. The number of moles at equilibrium constant, the number of moles at equilibrium constant. The number of one is the maximum increase, the number of one is the maximum increase. The number of one is the maximum increase. The number of one is the number of the number
$$

One mole of ethyl alcohol was treated with one mole of ${}^{\circ}$ C. 2/3 of the acid changes into ester at equilibrium. Find the equilibrium constant for the reaction of hydrolysis of ester

I'mass action

In the Moles at equilibrium $1 - \alpha$ and
 $\frac{200 \text{ C}_2 \text{H}_5 \parallel \text{H}_2\text{O}}{[H_2^2]} = \frac{x^2}{\left[\frac{ax}{y}\right] \times \left[\frac{bx}{y}\right]} = \frac{x^2}{\left[\frac{ax}{y}\right] \times \left[\frac{bx}{y}\right]}$
 $\frac{1}{\left[\frac{ax}{y}\right] \times \left[\frac{bx}{y}\right]} = \frac{x^2}{\left[\frac{ax}{y}\right] \times \left[\frac{bx}{y}\right]}$ $\lfloor v \rfloor \rfloor \lfloor v \rfloor$
 $\lfloor v \rfloor \rfloor \lfloor v \rfloor$
 $\lfloor (a-x)(b-x) \rfloor \rfloor$
 $\frac{x^2}{(a-x)(b-x)}$. If $a = b = 1$ then $K_c = \frac{x^2}{(1-x)^2}$

Example 10 :

Cone mole of ethyl alcohol was traces

does not involve volume term thus K_c is not affecte $\frac{x^2}{(1-x)}$. If $a = b = 1$ then $K_c = \frac{x^2}{(1-x)^2}$

Example 10:

See and throubus bottom term thus K_c is not affected

ge in volume $K_c \propto V^0$

sepending the term of the sequilibrium

Sol. C₂H₃OH(9)+CH₃COOH(9)=C

co **Sol.** $C_2H_5OH(\ell) + CH_3COOH(\ell) \rightleftharpoons CH_3COOC_2H_5(\ell) + H_2O(\ell)$ Initially $1 \qquad 1 \qquad 0 \qquad 0$ Moles at equilibrium $1-\alpha$ 1 – α 0 α 1 – α 1 Given $\alpha = 2/3$ Therefore, $C_2H_5OH = 1 - 2/3 = 1/3$ 10:

and the mole of ethyl alcohol was treated with one mole of

unitary ${}^{\circ}C$. 2/3 of the acid changes into ester at

uilibrium. Find the equilibrium constant for the reaction

hydrolysis of ester

H₂OH(ℓ)+CH₃C of ethyl alcohol was treated with one mole of

... °C. 2/3 of the acid changes into ester at

m. Find the equilibrium constant for the reaction

sis of ester
 $2+CH_3COOH$ (ℓ) $\rightleftharpoons CH_3COOC_2H_5(\ell) + H_2O(\ell)$

1 0 0

quilibr of HI is dissociated therefore α = .5

HI] = 1 - 0.5 = 0.5 ; [H₂] = 0.5/2 ; [L₂] = 0.5/2
 $\frac{[H_2][I_2]}{[HI]^2} = \frac{0.5/2 \times 0.5/2}{[0.5]^2} = \frac{1}{4} = 0.25$
 $\frac{[H_2][I_2]}{[HI]^2} = \frac{0.5/2 \times 0.5/2}{[0.5]^2} = \frac{1}{4} = 0.25$
 $\$ HI] = 1 - 0.5 = 0.5 ; [H₂] = 0.5/2 ; [L₂] = 0.5/2
 $\frac{E}{E}$ _c = $\frac{[H_2][I_2]}{[HI]^2} = \frac{0.5/2 \times 0.5/2}{[0.5]^2} = \frac{1}{4} = 0.25$

1:

10.5² = $\frac{1}{4} = 0.25$

1:

10.5² = $\frac{1}{2} = 0.25$

10.5² = $\frac{1}{2} = 0.25$
 $\alpha/2$ $\alpha/2$

ore $\alpha = .5$

0.5/2; [I₂] = 0.5/2
 $\frac{1}{2} = \frac{1}{4} = 0.25$

treated with one mole of

acid changes into ester at

m constant for the reaction

CH₃COOC₂H₅(ℓ) + H₂O(ℓ)

0
 α α

= 1/3

3
 ^oC. 2/3 of the acid changes into ester at

brium. Find the equilibrium constant for the reaction

drolysis of ester

drolysis of estre
 $-\alpha$ at equilibrium
 $\alpha = 2/3$

fore, C₂H₃OH = 1 - 2/3 = 1/3

CH₃COOH = 1 - 2 brum. Find the equilibrium constant for the reaction
 Orblysis of estr
 $\text{DH}(\ell) + \text{CH}_3\text{COOH}(\ell) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(\ell) + \text{H}_2\text{O}(\ell)$
 $\text{or} \quad 1 - \alpha$ a
 $\alpha = 2/3$
 $\alpha = 2/3$
 $\text{CH}_3\text{COOH} = 1 - 2/3 = 1/3$
 $\text{CH}_3\text{CO$

CH₃COOH = 1-2/3 = 1/3
CH₃COOC₂H₅ = 2/3, H₂O = 2/3
K_c =
$$
\frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]} = \frac{2/3 \times 2/3}{1/3 \times 1/3} = 4
$$

Example 11 :

For the reaction, $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$, the partial pressures of CO_2 and CO are 2.0 and 4.0 atm respectively at equilibrium. Find the K_p for the reaction

Sol.
$$
K_p = \frac{[p_{co}]^2}{[p_{co2}]} = \frac{[4]^2}{[2]}
$$
 \therefore $K_p = 8$

REACTION QUOTIENT AND EQUILIBRIUM CONSTANT

Consider the following reversible reaction $A + B \rightleftharpoons C + D$

 θ is the ratio of the product of θ . active masses of the products and product of active masses of the reactants.

 $Q_c = \frac{[C][D]}{[A][B]}$

The concentration are not necessarily equilibrium concentration.

[At equilibrium $Q_c = K_c$]

Case I : If Q^c < K^c then : [Reactants] > [Products] then the system is not at equilibrium.

The value of $\frac{[Product]}{[Reader]}$ is small.

 \therefore For establishment of equilibrium the reaction will go in forward direction. [Reactants \rightarrow Products]

Case II : **If** $Q_c = K_c$ then :

The system is at equilibrium and the concentration of the $Q.7$ species C, D, B, A are at equilibrium.

Case III : If $Q_c > K_c$ **then :** [Products] > [Reactants] The system is not at equilibrium.

The value of $\frac{[Product]}{[Reader]}$ is large

 \therefore For establishment of equilibrium the reaction will go in $Q.8$ backward direction [Products \rightarrow Reactants]

Note : Suppose there are three reaction having equilibrium constant K_1 , K_2 and K_3 and $K_1 > K_2 > K_3$ Hence the order of the tendency of the reaction to proceed $Q.9$ in forward direction (i.e. formation of product) Reaction 1 > Reaction 2 > Reaction 3 In short, $K \uparrow$, forward reaction \uparrow Ishment of equilibrium the reaction will go in

section [Products \rightarrow Reactants]

section [Products \rightarrow Reactants]

section and K₁ > K₂ > K₃

der of the tendency of the reaction to proceed
 K_2 and K₃ and K₁

Example 12 :

For the reaction NOBr (g) \rightleftharpoons NO (g) + $\frac{1}{2}$ Br₂ (g) $K_p = 0.15$ atm at 90°C. If NOBr, NO and Br₂ are mixed at this temperature having partial pressures 0.5 atm, 0.4 atm and 0.2atm respectively, will Br_2 be consumed or formed ?

Sol.
$$
Q_p = \frac{[P_{Br_2}]^{1/2} [P_{NO}]}{[P_{NOBr}]} = \frac{[0.2]^{1/2} [0.4]}{[0.50]} = 0.36
$$

\n $K_p = 0.15$ \therefore $Q_p > K_p$
\nHence, reaction will shift in backward direction.
\n \therefore Br₂ will be consumed

TRY IT YOURSELF-1

- **Q.1** At 500°C, the equilibrium constant for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 6.02×10^{-2} mol⁻² L. What is the value of K_p at the same temperature?
- **Q.2** For the reaction $\overline{A} + 3\overline{B} \rightleftharpoons 2C + D$, initial mole of A is twice that of B. If at equilibrium moles of B and C are

- **Q.3** $N_2 + O_2 \rightleftharpoons 2NO$ has equilibrium constant 2.5×10^{-5} , calculate equilibrium constant for the reaction $NO \rightleftharpoons \frac{1}{2}N_2 + \frac{1}{2}O_2.$
- **Q.4** For the reaction $2SO_2 + O_2 \rightleftharpoons 2SO_3$, equilibrium molar concentrations of SO_2 and SO_3 are equal and equilibrium constant is equal to 10. If reaction takes place in a 5 litre vessel then calculate equilibrium mole of O_2 .
- **Q.5** For the reaction PCl₅ (g) \rightleftharpoons PCl₃ (g) + Cl₂ (g). The moles of each component PCl_5 , PCl_3 and Cl_2 at equilibrium were found to be 2. If the total pressure is 3atm. The K_p will be – (A) 1 atm (B) 2 atm
	- $(C) 3$ atm $(D) 1.5$ atm
- **Q.6** In a reaction vessel of 2 litre capacity 3 moles of N_2 reacts with 2 moles of O_2 to produce 1 mole of NO. What is the molar concentration of N_2 at equilibrium? (A) 1.25 (B) 1.50 $(C) 0.75$ (D) 2.0
- **Q.7** At 87°C, the following equilibrium is established $H_2(g) + S(s) \rightleftharpoons H_2S(g)$; $K_p = 7 \times 10^{-2}$

If 0.50 mole of hydrogen and 1.0 mole of sulphur are heated to 87°C in 1.0 L vessel, what will be the partial pressure of H_2S at equilibrium?

(A) 0.966 atm (B) 1.38 atm (C) 0.0327 atm (D) 9.66 atm

- system is not at equilibrium.

leated to 87°C in 1.0 L vessel, what will be the partial

value of [Reactent] is large

(A) 0.966 smm

or establishment of equilibrium the reaction will go in (A) 0.966 smm

or establishment : IT Q₂ > K, then in [Products] > [Reactants] IT 0.50 mole of hydrogen and 1.0 mole of spin-

en is not a equilibrium,

en is not a equilibrium the reaction will go in

Reactant [Reactants]

c(A) 09664mm

(A) 1.38 smm
 Q.8 At a certain temperature the equilibrium constant K_c is 0.25 for the reaction $A_2(g) + B_2(g) \rightleftharpoons C_2(g) + D_2(g)$ If we take 1 mole of each of the four gases in a 10 litre container, what would be equilibrium concentration of $A_2(g)$ 032/atm

certain temperature the equilibrium constant K_c is

cor the reaction A₂(g) + B₂(g) $\rightleftharpoons C_2$ (g) + D₂(g)

take 1 mole of each of the four gases in a 10 litre

iner, what would be equilibrium concentrati 2/am

ain temperature the equilibrium constant K_c is

ain temperature the equilibrium constant K_c is

(a) and the reaction A₂(g) + B₂(g) $\Rightarrow C_2(g) + D_2(g)$

ie 1 mole of each of the four gases in a 10 litre

r, what
	- **Q.9** Consider following reactions in equilibrium with equilibrium concentration 0.01 M of every species
		- (I) $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$. (II) $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)$.

(III) N₂(g) + 3H₂(g)
$$
\rightleftharpoons
$$
 2NH₃(g).

Extent of the reactions taking place is –

- (A) $I > II > III$ (B) $I < II < III$
- (C) $II < II I < I$ (D) $III < I < II$
- **Q.10** 1 mole N_2 and 3 mol H_2 are placed in a closed container at a pressure of 4 atm. The pressure falls to 3 atm at the same temperature when the following equilibrium is attained. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. The equilibrium constant K_p for dissociation of NH₃ is – take 1 mole of each of the four gases in a 10 litre

iner, what would be equilibrium concentration of

(a)

inter what would be equilibrium concentration of

(b)

inter what would be equilibrium concentration of
 $2CI_3(g) \$ $\frac{3 \times 3}{2 \times 1}$

Sales in a 10 litre

m concentration of

equilibrium with

every species
 $\frac{3-1}{2}$
 $\frac{1}{2}$
 $\frac{3 \times 3}{2 \times 1}$
 $\frac{3 \times 3}{2 \times 1}$

(A)
$$
\frac{1}{0.5} \times (1.5)^3 \text{ atm}^{-2}
$$
 (B) $0.5 \times (1.5)^3 \text{ atm}^2$
(C) $\frac{0.5 \times (1.5)^3}{3 \times 3} \text{ atm}^2$ (D) $\frac{3 \times 3}{0.5 \times (1.5)^3} \text{ atm}^{-2}$

extraction of

the equilibrium concentration of

the following reactions in equilibrium with

ium concentration 0.01 M of every species
 $\frac{1}{5}(g) \rightleftharpoons PCI_3(g) + CI_2(g).$
 $(g) \rightleftharpoons H_2(g) + I_2(g).$
 $(g) \rightleftharpoons H_2(g) + I_2(g).$
 $(g) \right$ Frame concentration of

in equilibrium with

of every species

..

..

..

..

..

..

..

..

...
 $\text{or} \pm \text{or} \pm \$ **Q.11** For the reaction, $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$. $K_c = 1.8 \times 10^{-6}$ at 184 \degree C and R = 0.083 JK⁻¹ mol⁻¹. When K_p and K_c are compared at 184°C, it is found that – (A) $K_p > K_c$ (B) $K_p < K_c$ (C) $K_p = K_c$ (D) $K_p \geq K_c$ depends upon pressure of gases

Q.12 A 20.0 litre vessel initially contains 0.50 mole each of H_2 (D) and I_2 gases. These substances react with finally reach \blacksquare an equilibrium condition. Calculate the equilibrium concentration of HI if K_{eq} = 49 for the reaction

ANSWERS

(1) 1.5×10^{-5} atm⁻². **(2) (D) (3)** 2×10^{2} **(4)** 0.5 mole **(5)** (A) **(6)** (A) **(7)** (A) **(8)** 0.13 **(9)** (B)

(10)(B) **(11)** (A) **(12)** (B)

LECHATELIER'S PRINCIPLE

According to this principle, "If a system at equilibrium is 2. subjected to a change of concentration, pressure or temperature, the equilibrium shift in a direction that tends to undo the effect of the change".

(A) Effect of change in concentration : In a chemical equilibrium, increasing the concentration of the reactants result in shifting the equilibrium in favour of products (i.e. equilibrium shift in forward direction). Increasing concentration of the products results in shifting the equilibrium in favour of reactants (i.e in the backward direction).

 Concentration Equilibrium shift

Concentration of reactant \uparrow Forward direction

Concentration of product \uparrow Backward direction

- **(B) Effect of change in pressure :** If the pressure is increased reaction will take place in that direction, which will bring lowering of pressure or lowering in number of mole.
	- \overrightarrow{a} Increase of pressure shifts the equilibrium in that \overrightarrow{a} . direction where number of moles decreases. $(\Delta n \le 0)$ eg. $N_2 + 3H_2 \rightleftharpoons 2NH_3$; $\Delta n = 2 - 4 = -2 < 0$
	- i.e. on \uparrow pressure equilibrium shift in forward direction. **(ii)** Decrease of pressure shifts the equilibrium in that direction where number of mole increase. $(\Delta n > 0)$

eg.
$$
PCl_5 \rightleftharpoons PCl_3 + Cl_2
$$
 $\Delta n = 2 - 1 = +ve$

i.e. \downarrow in pressure equilibrium shift in forward direction (iii) Change of pressure has no effect if $\Delta n = 0$

eg. $H_2 + I_2 \rightleftharpoons 2HI$ \therefore $\Delta n = 2 - 2 = 0$

Note : Addition of an inert gas at constant pressure has the same effect as produced by the increased volume of the container. Thus, for equilibrium $\Delta n = 0$, when an inert gas is added at constant pressure, the equilibrium remains unaffected (since V does not appear in the expression of K_p) while for equilibrium having $\Delta n \neq 0$, the addition of an inert gas at constant pressure causes reaction to move in the direction of more moles of the gases.

(C) Change in temperature :

- **(i) For endothermic reaction :** If we increase the temperature equilibrium shift in forward direction which proceed with absorption of heat.
- **(ii) For exothermic reaction :** If we increased the temperature the equilibrium shifts in backward direction which proceed with evolution of the heat.

(D) Role of Catalyst : Positive catalyst increase the rate of both the forward and backward reaction equally so the equilibrium will be attained in less time. Negative catalyst decrease the rate of both forward and backward reaction so equilibrium will be attained in more time. **Example 18 CONDUCE AT ANTIFICATE CONTROVANCED LEARNING**
 EXECUTE AT AND CONDUCT AT AND CONDUCT AT AND CONDUCT AT AND INCREDIBLY AND THE AT AND INTERTATION CASE ()
 EXECUTE AT AND INTERTATION CONDUCT AT AND AT ANTIFICA EXERCISE AND MANUST CONSUMING THE CONSUMER SUBMINER CONSUMER SURFAINING
 EXERCISE ADMOND CONSUMED AND SURFAINS (CONSUMER)

Elibrium will be attained in less time. Negative catalyst

rease the rate of both forward and

Application of Le Chatelier principle on physical equilibria: 1. Melting of Ice :

$$
H_2O(s) \underset{\text{Ice}}{\xrightarrow{\bullet} H_2O(\ell)}
$$

Here

(More volume) (Less volume)

With increase in pressure melting point of ice is decreased and equilibrium will be shifted in a direction of less volume.

2. Vapourisation of liquid :

(less volume) (more volume)

Vaporisation of liquid is endothermic process in the nature i.e. the evaporation of liquid in to its vapour is completed by absorption of heat so rise of temperature will favour vaporisation and equilibrium is shifted in forward direction. On the other hand on increasing the pressure the equilibrium will be shifted in direction of less volume means water cannot be converted in to vapour and boiling point increases.

3. Melting of Sulphur :

Sulphur (s) \rightleftharpoons sulphur (ℓ)

Less volume More volume On increasing pressure the equilibrium will shift towards less volume, means solid will be not converted in to liquid, thus melting point of sulphur increases. will be shifted in direction of less volume means

ot be converted in to vapour and boiling point

rature Vaporisation ↑ forward shift

wire B.P ↑ Backward shift

Sulphur :
 \Rightarrow sulphur (ℓ)

here volume

ing pressure not be converted in to vapour and boiling point

berature Vaporisation \uparrow forward shift

sure B.P \uparrow Backward shift

f**Sulphur :**
 \downarrow \uparrow \uparrow

Example 13 :

Describe the effect of (a) Addition of H_2 (b) Addition of $CH₃OH$ (c) Removal of CO (d) Removal of $CH₃OH$ on the equilibrium of the reaction :

 $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

Sol. The equilibrium constant for the given reaction is

$$
K_c = \frac{[CH_3OH]}{[H_2]^2 [CO]}
$$

- (a) Addition of H_2 will push the reaction in the forward direction.
- (b) Addition of $CH₃OH$ will push the reaction in the backward direction.
- (c) When CO is removed, the reaction will move in the backward direction.
- (d) When $CH₃OH$ is removed, the reaction will move in the forward direction.

x

Example 14 :

The equilibrium constant of the reaction at 25°C $CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4.3H_2O(s) + 2H_2O(g)$ is 1.084×10^{-4} atm². Find out under what conditions of relative humidity, $CuSO₄$.5H₂O will start loosing its water of crystallization according to above reaction. (Vapour pressure of water at 25°C is 24mm of Hg). Relative humidity **STUD**

STUD

Del 14:

requilibrium constant of the reaction at 25°C

CuSO₄₋SH₂O(g) C=CuSO₄-SH₂O(g) F(10CO₄-SH₂O(g) F(10CO₄-SH₂O(g) F(10CO₄-SH₂O owill start looking its water

variable the metal of the

 $=\frac{\text{ratural mass}}{2}$ \therefore 9 2° at that temperature

Sol. $K_p = (P_{H2O})^2$

 $P_{H_2O} = \sqrt{1.084 \times 10^{-4}} = 1.041 \times 10^{-2}$ atm ≈ 8 mm of Hg

In a room, pressure of water is greater than 8mm of Hg then $CuSO₄$. $3H₂O$ will absorb water from air and will form $CuSO₄.5H₂O$ and will keep absorbing until partial pressure of H_2O becomes 8mm of Hg. relative humidity, CuSO₄.5H₂O will start loosing its water
of crystallization according to above reaction.
(Vapour pressure of water at 25°C is 24mm of Hg).
Relative humidity
 $= \frac{\text{Partial pressure of H}_2\text{O vapours}}{\text{Vapour pressure of H}_2\text{O at that temperature}}$

If P_{H_2O} < 8mm of Hg then CuSO₄.5H₂O will loose water of crystallization and reaction will move in forward direction.

i.e., if relative humidity
$$
\frac{8}{24}
$$
 < 33.33%

then $CuSO_4.5H_2O$ will loose water of crystallization.

FREE ENERGY CHANGE AND EQUILIBRIUM CONSTANT

Let ΔG° be the difference in free energy of the reaction when all the reactants and products are in standard state (1 atm pressure and 298 K) and K_c and K_p be the thermodynamic equilibrium constant of the reaction. Both are related to each other at temperature T by following relation.

 $\Delta G^{\circ} = -2.303 \text{ RT} \log K_C \& \Delta G^{\circ} = -2.303 \text{ RT} \log K_p$
Equation relates equilibrium constant of a reaction to a thermochemical property.

where $\Delta G^{\circ} = \Delta H^{\circ} - T \Delta S^{\circ}$

here ΔH° = Standard enthalpy change

 ΔS° = Standard entropy change

Three cases may arise : (i) $\Delta G^{\circ} = 0$ then $K_c = 1$ (ii) $\Delta G^{\circ} > 0$ i.e. + ve then K_c < 1

In this case reverse reaction is feasible i.e. less concentration of products at equilibrium state.

(iii) $\Delta G^{\circ} < 0$ i.e. –ve then $K_c > 1$

In this case forward reaction is feasible i.e. large concentration of products at equilibrium state.

In short, Free energy change Equilibrium constant Concentration of reactants and Products

$$
\Delta G^{\circ} = 0 \text{ K}_{c} = 1
$$
\n
$$
\Delta G^{\circ} > 0 \text{ K}_{c} < 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$
\n
$$
\Delta G^{\circ} < 0 \text{ K}_{c} > 1
$$

Example 15 :

NO and Br_2 at initial partial pressures of 98.4 and 41.3 torr, respectively, were allowed to react at 300 K. At equilibrium the total pressure was 110.5 torr. Calculate the value of the

equilibrium constant and the standard free energy change at 300 K for the reaction $2NO(g) + Br_2(g) \rightleftharpoons 2NOBr(g)$

At equilibrium 98.4-x 41.3 - $\frac{x}{2}$ x and the state of the stat 2 \overline{a}

Total pressure at equilibrium is 110.5 torr

WANIEBIALS	STUDY MATERIAL: CHEMSTRY
\n DE 14: \n $P_1 = \frac{1}{2} \times \frac{$	

$$
33.33\% = -2.92 \text{ k cal} = 12.2 \text{ kJ}
$$

[If R is used as 1.99 cal/mol K, then ΔG° will be in cal. If R is used as 8.314 J/mol K, then ΔG° will be in joules. But K_p must be in $(\text{atm})^{\Delta n}$. .

DEGREE OF DISSOCIATION BY DENSITY MEASUREMENT

Those type of reactions in which there is change in number of moles after dissociation the extent of dissociation can be determined by vapour density measurement. Let us consider the following reaction.

 $\text{PCl}_5 \quad \rightleftharpoons \quad \text{PCl}_3 + \text{Cl}_2$ Initially 1 0 0 Moles at equilibrium $1 - \alpha$ α α $(\alpha \rightarrow$ degree of dissociation)

Total no. of moles at equilibrium = $(1 - \alpha) + \alpha + \alpha = (1 + \alpha)$ Now, 'V' is the volume occupied by 1 mole of PCI_5 which have vapour density 'D' before dissociation and after dissociation is 'd'. So under same conditions, the volume occupied by ISSOCIATION BY DENSITY MEASUREMENT

DISSOCIATION BY DENSITY MEASUREMENT

De of reactions in which there is change in number

after dissociation the extent of dissociation can

mined by vapour density measurement.

mostl **SYDENSITYMEASUREMENT**
which there is change in number
n the extent of dissociation can
density measurement.
 \therefore \Rightarrow PCl₃ + Cl₂
0 0
 α α
 \rightarrow degree of dissociation)
brium = $(1 - \alpha) + \alpha + \alpha = (1 + \alpha)$
cupied by 1 m be accrimined by vapour density measurement.

Let us consider the following reaction.
 PCl_5 \rightleftharpoons $PCl_3 + Cl_2$

Initially $1 \qquad 0 \qquad 0$

Moles at equilibrium $1 - \alpha \qquad \alpha \qquad \alpha$

Total no. of moles at equilibrium $=(1 - \alpha) +$ d by vapour destact or dissociation can
detect or dissociation can
del by vapour density measurement.

 $PCI_5 \implies PCI_3 + CI_2$
in 0 0
($\alpha \rightarrow$ degree of dissociation)

noles at equilibrium = $(1 - \alpha) + \alpha + \alpha = (1 + \alpha)$
where we vol

 $(1 + \alpha)$ moles at equilibrium will be $(1 + \alpha)$ V litre.

Since, Density
$$
\propto \frac{1}{\text{Volume}}
$$
. Therefore, $D \propto \frac{1}{V}$

$$
d \propto \frac{1}{(1+\alpha)V}
$$
 or $\frac{D}{d} = 1 + \alpha$ or $\alpha = \frac{D-d}{d}$

Note : When one mole of reactant on dissociation gives 'n' moles of gaseous products the above equation takes in

the form of
$$
\alpha = \frac{D-d}{(n-1)d}
$$

Example 16 :

At 250 \degree C the vapour density of PCl₅ is 100 calculate the degree of dissociation at this temperature.

CHAPTERICAL EQUILBRICM	
Example 16:	Which of the following is correct, if the total pre- degree of dissociation at this temperature.
SoI. D for PCI ₅ = $\frac{\text{Molecular mass}}{2} = \frac{208.5}{2} = 104.25$	(b) K will remain same to be very high of the following
SoI. D for PCI ₅ = $\frac{\text{Molecular mass}}{2} = \frac{208.5}{2} = 104.25$	6) K will increase (C) K will increase initially and decrease when is very high of the following reactions, the equi- is very high
Q.1 The favourable conditions for melting of ice is – (A) Low pressure (C) Low temperature (D) Absence of catalyst (E) High pressure (E) High pressure (E) High pressure (E) High pressure (E) High pressure (E) High pressure (E) A is	6) PCI ₅ (g) \rightleftharpoons PCI ₃ (g) + CI ₂ (g) c) $\frac{\text{Molow magnetic}}{2} = \frac{208.5}{2} = 104.25$
Q.2 AG ^o for the dissociation of the following reactions, the equi- of classicality of the following reactions, the equi- of the following reaction of small amount of small amount of the following	
Q.3 aG ^o for the dissociation of the following reaction solution at 27°C is 6 909 kcal/mol If 8 moles of A is	6) The equilibrium will remain unaffected in all

TRY IT YOURSELF-2

- **Q.1** The favourable conditions for melting of ice is (A) Low pressure (B) High pressure (C) Low temperature (D) Absence of catalyst
- **Q.2** ΔG° for the dissociation of the dimer $(A \rightleftharpoons 2A)$ in benzene solution at 27^oC is 6.909 kcal/mol. If 8 moles of A is dissolved in 10dm³ of benzene at 27°C. What is the ratio of equilibrium concentration of monomer to dimer $([A]/A_2]$? Given : $R = 2$ cal/mol. K

(A)
$$
1:200
$$
 (B) $1:100$

(C)
$$
200:1
$$
 (D) $800:1$

- **Q.3** The exothermic formation of CIF_3 is represented by the equation: $Cl_2(g) + 3F_2(g) \rightleftharpoons 2ClF_3(g)$; $\Delta H = -329 \text{ kJ}$ Which of the following will increase the quantity of $CIF₃$ in an equilibrium mixture of Cl_2 , F_2 and ClF_3 .
	- (A) Increasing the temperature
	- (B) Removing $Cl₂$
	- (C) Increasing the volume of container
	- (D) Adding F_2
- **Q.4** Densities of diamond and graphite are 3.5 and 2.3 gm/mL. C (diamond) \rightleftharpoons C (graphite) ; Δ_r H = –1.9 kJ/mole favourable conditions for formation of diamond are –
	- (A) high pressure and low temperature
	- (B) low pressure and high temperature
	- (C) high pressure and high temperature
	- (D) low pressure and low temperature
- **Q.5** Following two equilibrium is simultaneously established in a container

 $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g).$

$$
CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)
$$

If some Ni (s) is introduced in the container forming $Ni(CO)₄$ (g) then at new equilibrium

- (A) PCl_3 concentration will increase
- (B) PCl_3 concentration will decrease
- (C) Cl_2 concentration will remain same
- (D) CO concentration will remain same
- **Q.6** The vapour density of N_2O_4 at a certain temperature is $*$ 30. What is the % dissociation of N_2O_4 at this temperature?

- **Q.7** For the equilibrium, $PC_{\frac{1}{2}}(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ at 298 K, $K = 1.8 \times 10^{-7}$. What is ΔG° for the reaction?
- **Q.8** On increasing the pressure, in which direction will the gas phase reaction proceed to re-establish equilibrium, is predicted by applying the Le Chatelier's principle. Consider the reaction. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Which of the following is correct, if the total pressure at which the equilibrium is established, is increased without changing the temperature?

- (A) K will remain same
- (B) K will decrease
- (C) K will increase

(D) K will increase initially and decrease when pressure is very high

- **Q.9** In which of the following reactions, the equilibrium remains unaffected on addition of small amount of argon at constant volume?
	- $(A) H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
	- $(B) PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
	- $(C) N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
	- (D) The equilibrium will remain unaffected in all the three cases

ANSWERS

IMPORTANT POINTS

 K_P is related to K_C as $K_P = K_C (RT)^{\Delta n}$

Standard free energy & Equilibrium constant

 $\Delta G = \Delta G^0 + RT \ln Q$; the condition for a system to be at equilibrium is that $\Delta G = 0$

Thus at equilibrium,

$$
\Delta G^0 = -RT \ln K \, , \, \ln K = \frac{-\Delta G^0}{RT}
$$

 ΔG^0 = –ve or K > 1: forward reaction is feasible, ΔG^0 = +ve or $K < 1$: reverse reaction is feasible,

 ΔG^0 = 0 or K = 1: reaction is at equilibrium (very rare).

ADDITIONAL EXAMPLES

Example 1 :

6 gram hydrogen reacts with 9.023×10^{23} chlorine molecules to form HCl. If total pressure of the reaction is 800 mm then find the partial pressure of HCl.

Sol.
$$
H_2 + Cl_2 \rightleftharpoons
$$
 2HCl

3 mole $\frac{9.023 \times 10^{23}}{6.023 \times 10^{23}}$ 23 0 Initial mole

 $3 - 1.5$ mole $1.5 - 1.5$ mole 3.0 Mole at equilibrium Total mole at equilibrium = $1.5 + 3.0 = 4.5$ mole

Partial pressure of HCl = $\frac{3.0}{4.5} \times 800 = \frac{1600}{3} = 533$ mm

Example 2 :

In which case does the reaction go farthest to completion : K=1; K = 10^{10} ; K = 10^{-10} and why?

Sol. The ratio $\frac{[Product]}{[Reader]}$ is maximum when K = 10¹⁰ and thus $\frac{Product}{Readnt}$ is maximum when K = 10¹⁰ and thus

reaction goes farthest to completion when $K = 10^{10}$

Example 3 :

The equilibrium constant K_c for $A_{(g)} \rightleftharpoons B_{(g)}$ is 1.1. Which gas has a molar concentration greater than 1 ?

reaction goes farthest to completion when K = 10¹⁰
\n**Example 3 :** The equilibrium constant K_c for A_(g) ⇒ B_(g) is 1.1. Which
\ngas has a molar concentration greater than 1 ?
\n**Sol.** For A_(g) ⇒ B_(g) ; K_c =
$$
\frac{[B]}{[A]} = 1.1
$$

\n⇒ [B] > [A]. If [B] = 1; [A] = 0.91
\nCase I 0.91 < \leq [A] ≤ 1
\nand [A] > 1
\n**Example 4 :**
\n0.1 mole of N₂O_{4(g)} was sealed in a tube under atmospheric
\nprobabilityum
\n $\log A = 1.3 \times 10^{-2} = 1$
\n \Rightarrow [B] > [A]. If [B] = 1; [A] = 0.91
\nCase II [A] > 1
\n**Example 5 :** In the reaction, C(s) + CO₂(g) ⇒ 2CO(g)
\n\n**Example 4 :**
\n0.1 mole of N₂O_{4(g)} was sealed in a tube under atmospheric
\n $\log A = 2000$
\n $\log A = 2000$
\n $\log A = 2000$
\n $\log B = 2000$
\n $\log B = 2000$
\n $\log B = 2000$
\n $\log C = 200$
\n $\log C = 200$
\n $\log C = 200$
\n $\log A = 2000$
\n $\log B = 1.5$
\n $\log C = 1.5 \times 12 = 8$
\nor K_p = $\frac{(2\alpha)^2}{(0.1 - \alpha)} \times [\frac{P}{0.1 + \alpha}]$
\n $\therefore \alpha = 0.017$ \therefore NO₂ = 0.017 × 2 = 0.034 mole
\n \there

Example 4 :

0.1 mole of $N_2O_{4(g)}$ was sealed in a tube under atmospheric **Sol.** conditions at 25°C. Calculate the number of mole of NO_{2(g)} present, if the equilibrium

 $N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}(K_p=0.14)$ is reached after some time. **Sol.** $N_2O_4 \stackrel{\sim}{\rightleftharpoons} 2NO_2$

$$
\begin{array}{cc} 0.1 & 0 \\ (0.1 - \alpha) & 2\alpha \end{array}
$$

$$
K_{p} = \frac{(2\alpha)^{2}}{(0.1 - \alpha)} \times \left[\frac{P}{0.1 + \alpha}\right]^{1}
$$

\nTotal pressure
\nor $K_{p} = \frac{40\alpha^{2}}{(0.1 - \alpha)}$ or $\frac{40\alpha^{2}}{(0.1 - \alpha)} = 0.14$
\nPartial pressu
\nPartial pressu

$$
\therefore \ \alpha = 0.017 \ \therefore \ NO_2 = 0.017 \times 2 = 0.034 \ \text{mole}
$$

Example 5 :

What concentration of CO_2 be in equilibrium with 2.5×10^{-2} mol litre⁻¹ of CO at 100° C for the reaction $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$; K_c=5.0 **DYMATERIAL: CHEMISTRY**

2 be in equilibrium with

3 at 100° C for the reaction
 \cdot CO₂(g) ; K_c = 5.0
 $\left[\text{CO}_2\right]$

2.5 × 10⁻²

12.5 × 10⁻² mol litre⁻¹

n of H₂S if 1 mole of H₂S is

e vessel at 1000 K ?

STUDYMATERIAL: CHEMISTRY
\n**STUDYMATERIAL: CHEMISTRY**
\n
$$
\begin{array}{ll}\n & \text{STUDYMATERIAL: CHEMSTRY} \\
\hline\n\end{array}
$$
\n
$$
\begin{array}{ll}\n & \text{Example 5:} \\
 \text{What concentration of CO}_2 \text{ be in equilibrium with} \\
 & 2.5 \times 10^{-2} \text{ mol litre}^{-1} \text{ of CO at } 100^{\circ} \text{ C for the reaction} \\
\text{for form HCl. If total pressure of the reaction is} \\
\text{in find the partial pressure of HCl.} \\
\text{Cl}_2 \rightleftharpoons 2HCl \qquad \text{Sol.} \\
\begin{array}{ll}\n & K_c = \frac{[CO_2]}{[CO]} \\
 & \text{Sol.} \\
\text{Staplet 6:} \\
\text{the equilibrium} \\
\text{at equilibrium} = 1.5 - 1.5 \text{ mole} \\
\text{the equilibrium} = 1.5 + 3.0 = 4.5 \text{ mole} \\
\text{since of HCl:} \\
\text{HCl:} \\
\text{HCl:
$$

Example 6 :

What is the % dissociation of H_2S if 1 mole of H_2S is introduced into a 1.10 litre vessel at 1000 K ? K_c for the reaction : $2H_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$ is 1×10^{-6}

EXAMPLES

Example 5:

What concentration of CO₂ be in equilibrium with

sixth 9.023 × 10²³ chlorine

FeO(s) + CO(g) = Pe(s) + CO₂ (g) = Re(s) + Co₂ (g) **Sol.** 2H₂S_(g) \rightleftharpoons 2H_{2(g)} + S_{2(g)} Mole before dissociation 1 0 0 Mole after dissociation $(1-\alpha)$ α $\alpha/2$ Where α is degree of dissociation of H₂S Volume of container = 1.10 litre equilibrium
 $\%$ dissociation of H₂S if 1 mole of H₂S is

to a 1.10 litre vessel at 1000 K ? K_c for the
 $I_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$ is 1×10^{-6}
 $2H_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$ is 1×10^{-6}
 $2H_2S_{(g)} \rightleftharpoons$ O_2 at equilibrium

= 2.5 × 10⁻² × 5 = 12.5 × 10⁻² mol litre⁻¹

is the % dissociation of H₂S if 1 mole of H₂S is

aced into a 1.10 litre vessel at 1000 K ? K_c for the

n: $2H_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$ is 1 H S 1.1 2 $\begin{aligned}\n\frac{1}{\left|\frac{\text{S}}{\text{S}}\right|} &\therefore 5 = \frac{\left[\text{CO}_2\right]}{2.5 \times 10^{-2}} \\
\text{equilibrium} \\
\text{equilibrium} \\
\frac{1}{\left|\frac{\text{S}}{\text{S}}\right|} &\text{Equation of } \frac{1}{\left|\frac{\text{S}}{\text{S}}\right|} &\text{Equation of } \frac{1}{\left|\frac{\text{S}}{\text{S}}\right|} \\
\text{Equation of } \frac{1}{\left|\frac{\text{S}}{\text{S}}\right|} &\text{Equation of } \frac{1}{\left|\frac{\text{S}}{\text{S}}\$ CO₂ be in equilibrium with

fCO at 100° C for the reaction

(s) + CO₂(g); K_e = 5.0
 $=$ $\frac{[CO_2]}{2.5 \times 10^{-2}}$

m
 $5=12.5 \times 10^{-2}$ mol litre⁻¹

ation of H₂S if 1 mole of H₂S is

litre vessel at 1000 K ? K_e CO₂ be in equilibrium with

fCO at 100° C for the reaction

(s) + CO₂(g) ; K_c = 5.0
 $=$ $\frac{[CO_2]}{2.5 \times 10^{-2}}$

m
 $5 = 12.5 \times 10^{-2}$ mol litre⁻¹

ation of H₂S if 1 mole of H₂S is

litre vessel at 1000 K ? K_c ²² mol litre⁻¹ of CO at 100° C for the reaction
 \cdot CO(g) \rightleftharpoons Fe(s) + CO₂(g) ; K_c = 5.0

CO₂)

∴ 5 = $\frac{[CO_2]}{2.5 \times 10^{-2}}$
 $\frac{1}{2}$ at equilibrium

= 2.5 × 10⁻² × 5 = 12.5 × 10⁻² mol litre⁻¹
 3 at 100 C for the reaction

+ CO₂(g); K_c = 5.0

<u>[CO₂]</u>

2.5 × 10⁻² mol litre⁻¹

nn of H₂S if 1 mole of H₂S is

e vessel at 1000 K ? K_c for the

2(g) + S_{2(g)} is 1 × 10⁻⁶

1−α) α α/2

 + CO₂(g) ; K_c = 5.0

[CO₂]

12.5 × 10⁻² mol litre⁻¹

on of H₂S if 1 mole of H₂S is

e vessel at 1000 K ? K_c for the
 $2(g)$ + S_{2(g)} is 1 × 10⁻⁶
 $1_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$
 $1-\alpha$ α α/2

ociation of 2 1.1 is the % dissociation of H₂S if 1 mole of H₂S is

duced into a 1.10 litre vessel at 1000 K ? K_c for the

ion : 2H₂S_(g) ⇒ 2H_{2(g)} + S_{2(g)} is 1 × 10⁻⁶

2H₂S_(g) ⇒ 2H_{2(g)} + S_{2(g)} is 0

before di

$$
\therefore K_c = \frac{\left[H_2\right]^2 \left[S_2\right]}{\left[H_2 S\right]^2} = \frac{\left[\frac{\alpha}{1.1}\right]^2 \left[\frac{\alpha}{2 \times 1.1}\right]}{\left[\frac{1-\alpha}{1.1}\right]^2} = 1 \times 10^{-6}
$$

Assuming $1 - \alpha = 1$ since α is small because $K_c = 10^{-6}$

$$
\therefore \quad \frac{\alpha^3}{2(1.1)} = 10^{-6} \quad ; \ \alpha = 1.3 \times 10^{-2} = 1.3\%
$$

Example 7 :

In the reaction, $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$, the equilibrium pressure is 12 atm. If 50% of CO_2 reacts, calculate K_p .

Sol. $\text{C(s)} + \text{CO}_2(g) \rightleftharpoons 2\text{CO}(g)$ Gaseous mole 1 0 before dissociation –

and why ?
\nand why ?
\nand why ?
\nWhen K = 10¹⁰ and thus
\nplot after dissociation (1-α)
\n
$$
K_c = \frac{[H_2]^2 [S_2]}{[H_2S]^2} = \frac{\left[\frac{\alpha}{1.1}\right]^2 \left[\frac{\alpha}{2 \times 1.1}\right]}{\left[\frac{1-\alpha}{1.1}\right]} = 1 \times 10^{-6}
$$
\n
$$
= 1.1
$$
\n
$$
= 0.91
$$
\n
$$
= 0
$$

 \therefore Total mole = 1.5 Total pressure given at equilibrium = 12 atm

Partial pressure of CO =
$$
\frac{1}{1.5} \times 12 = 8
$$

Partial pressure of CO₂ =
$$
\frac{0.5}{1.5} \times 12 = 4
$$

$$
K_p = \frac{P_{CO}^2}{P_{CO_2}} = \frac{(8)^2}{4} = 16 \text{ atm}
$$

Example 8 :

When α -d glucose is dissolved in water, it undergoes a partial conversion to β –d glucose to exhibit mutarotation. This conversion stops when 63.6% of glucose is in β from. Assuming that equilibrium has been attained, calculate K_C for mutarotation.

Sol. α –d glucose $\rightleftharpoons \beta$ –d glucose At equilibrium 36.4 63.6

$$
\therefore K_c = \frac{63.6}{36.4} = 1.747
$$

Example 9 :

What should be the respective active masses of methyl alcohol and carbon tetrachloride, if their densities are 0.5 and 1.2g/ml ?

Sol. (1). Active mass = number of g molecules per litre

We know that
$$
D = \frac{M}{V}
$$
 and $M = D \times V$ (a--

(1) The amount of CH₃OH in one litre = $0.5 \times 1000 = 500$ g

- (2) The amount of CCl_4 in one litre = $1.2 \times 1000 = 1200$ g
- (a) Active mass of CH₃OH = $\frac{\text{Weight in g}}{\text{Molecular weight } \times \text{Volume}}$ $\frac{2x}{a-x} = \sqrt{9 \times 10^{-4}} =$

[Molecular weight of CH₃OH = $12 + 3 + 16 + 1 = 32$]

$$
= \frac{500}{32 \times 1} = 15.62 \text{g} \text{ mole/litre}
$$
 x = 0.0

(b) Active mass of CCl₄ = $\frac{1200}{154 \times 1}$ = 7.79g mole/litre At 87°C, the foll
H₂(g) + S (s

[Molecular weight of CCl₄ = $12 + 35.5 \times 4$] = $12 + 142 = 154$

Example 10 :

For the reaction $A + 2B \rightleftharpoons 2C + D$, initial concentration of A is a and that of B is 1.5 times that of A. Concentration of A and D are same at equilibrium. What should be the concentration of B at equilibrium ?

Sol. (B).

 $A(g) + 2B(g) \rightleftharpoons 2C(g) + D(g)$

a 1.5 a 0 0 No. of initial mole $(a-x)$ $(1.5 a-2x)$ 2x x No. of moles at equilibrium Concentration of A and D are same at equilibrium

 $a - x = x$; $2x = a$; $x = a/2$

Therefore, concentration of B = 1.5 (a) – 2 (a/2) $[B] = 0.5a$; $[B] = a/2$

Example 11 :

Value of K_c at 300°C for N₂ + O₂ \rightleftharpoons 2NO is 9 × 10⁻⁴ and equimolecular amounts of N_2 and O_2 are used. The concentration of NO at equilibrium will be – (A) 0.0148a (B) 0.296a

(C) 0.148a (D) 0.0296a

Sol. (B).

^M ^D 32 1 154 1 N² (g) + O² (g) 2NO (g) K^c = 9 × 10–4 a a 0 Initial mole (a – x) (a – x) 2x Mole at equilibrium 2 2 c 2 2 2 [NO] 4x ^K [N] [O] (a x) = 9 × 10–4 2x 4 2 9 10 3 10 a x

$$
2x = 0.03 (a - x) ; 2x = 0.03a - 0.03x
$$

\n
$$
2x + 0.03x = 0.03a ; 2.03x = 0.03a
$$

\n
$$
x = 0.0148a ; 2x = 0.0296a
$$

Example 12 :

At 87°C, the following equilibrium is established : $H_2(g) + S(s) \rightleftharpoons H_2S(g), K_c = 8 \times 10^{-2}$ If 0.3 mole hydrogen and 2 mole sulphur are heated to 87^oC in a 2L vessel, what will be the partial pressure of $H₂S$ approximately at equilibrium. [Use $R = 0.08$ atm. L/mol. K) (A) 0.32 atm (B) 0.43 atm (C) 0.62 atm (D) 4.0 atm $K_c = \frac{[NO]^2}{[N_2][O_2]} = \frac{4x^2}{(a-x)^2}$ = 9 × 10⁻⁴
 $\frac{2x}{a-x} = \sqrt{9 \times 10^{-4}}$ = 3 × 10⁻²
 $2x = 0.03 (a-x)$; 2 x = 0.03a – 0.03x
 $2x + 0.03x = 0.03a$; 2 0.03x = 0.03a
 $x = 0.0148a$; 2 x = 0.0296a

2:

2:

2:

2:

2:

7°C, th $\frac{[NO]^2}{[N_2][O_2]} = \frac{4x^2}{(a-x)^2} = 9 \times 10^{-4}$
= $\sqrt{9 \times 10^{-4}} = 3 \times 10^{-2}$
03 (a-x) ; $2x = 0.03a - 0.03x$
03x= 0.03a ; $2.03x = 0.03a$
148a ; $2x = 0.0296a$
e following equilibrium is established :
+ S (s) $\rightleftharpoons H_2S$ (g), (a-x) 2x Mole at equilibrium

= $\frac{[NO]^2}{[N_2][O_2]} = \frac{4x^2}{(a-x)^2} = 9 \times 10^{-4}$
 $\frac{1}{2} = \sqrt{9 \times 10^{-4}} = 3 \times 10^{-2}$

0.03 (a-x) ; 2x = 0.03a – 0.03x

0.03x = 0.03a ; 2.03x = 0.03a

0.0148a ; 2x = 0.0296a

the following equ C, the following equilibrium is established :
 $2(9) + S (s) \rightleftharpoons H_2S (g), K_c = 8 \times 10^{-2}$

mole hydrogen and 2 mole sulphur are heated to

in a 2L vessel, what will be the partial pressure of

pproximately at equilibrium.
 $=$ $\frac{2x}{a-x} = \sqrt{9 \times 10^{-4}} = 3 \times 10^{-2}$
 $2x = 0.03 (a-x)$; 2x = 0.03a - 0.03x
 $2x + 0.03x = 0.03a$; 2.03x = 0.03a
 $x = 0.0148a$; 2x = 0.0296a

2:

2:

27°C, the following equilibrium is established :

27°C, the following equili 21 $(a-x)^2$
 $0^{-4} = 3 \times 10^{-2}$

22 $x = 0.03a - 0.03x$

22 $(2.03x - 0.03a)$

22 $(x - 0.0296a)$

22 $x = 0.02$

23 $x = 0.02$

24 $x = 0.043$

22 $x = 0.022$

22 $x = 0.022$

22 $=\sqrt{9 \times 10^{-4}} = 3 \times 10^{-2}$

.03 (a - x) ; 2x = 0.03a - 0.03x

03x = 0.03a ; 2.03x = 0.03a

1148a ; 2x = 0.0296a

he following equilibrium is established :

+ S (s) \rightleftharpoons H₂S (g), K_c = 8 × 10⁻²

e hydrogen and 2 m

Sol. (A).
$$
K_c = \frac{[H_2 S (g)]}{[H_2 (g)]} \Rightarrow 8 \times 10^{-2} = \frac{x}{0.3 - x}
$$

0.024 - 0.08 = x ; 0.024 = 1.08x ; x=0.022
 $P_{H_2S} = \frac{0.022 \times 0.08 \times 360}{0.22 \times 0.08 \times 360} \approx 0.32 \text{ atm}$

IONIC EQUILIBRIUM

INTRODUCTION

Those compounds which can conduct electricity either in fused state or in solution state in any solvent are called (e) Electrolytes.

Compounds which contain ions in pure solid state and can conduct electricity in the pure liquid state are called True electrolytes e.g. NaCl, KCl etc.

Compounds which do not conduct electricity in fused state but their solution in a solvent can conduct electricity are (f) called potential electrolytes e.g. HCl, $CH₃COOH$, NaOH, $NH₄OH$ etc.

There are two types of electrolytes –

- **(a) Strong Electrolytes :** Compounds which completely ionised in aqueous solution and are good conductor of electricity e.g. HCl, H_2SO_4 , HNO₃, NaOH, KOH, CH₃COONa, NH₄Cl etc. (α = 100%)
- **(b) Weak Electrolytes :** Compounds which ionise partially in aqueous solution and are poor conductor of electricity e.g. CH₃COOH, HCN, NH₄OH, Al(OH)₃ etc. (α < 100%) F Those compounds which do not conduct electricity either 1. in fused state or their solution in any solvent are called Non Electrolytes.

Example 1 :

Identify the non-electrolyte, strong electrolyte and weak electrolytes in aq. medium of the following substance. NaOH, HCl, NaCl, $NH_2\text{CONH}_2$ (urea), $H_2\text{SO}_4$, $H_3\text{PO}_4$, HCN, NH₄OH, KOH, C₂H₅COOH, NH₄Cl, HNO₃, $C_{12}H_{12}O_{11}$ (sugar), $C_6H_{12}O_6$ (glucose)

Sol. Non-electrolyte : NH_2 CON H_2 , C₁₂H₁₂O₁₁, C₆H₁₂O₆ Weak electrolyte : H_3PO_4 , HCN, NH₄OH, C₂H₅COOH 4. N Strong electrolytes : NaOH, HCl, NaCl, H₂SO₄, KOH, $NH₄Cl, HNO₃$. (sugar), $C_6H_{12}O_6$ (glucose)

yet: NH_2CONH_2 , $C_{12}H_{12}O_{11}$, $C_6H_{12}O_6$

yet: H_3PO_4 , HCN , NH_4OH , C_2H_5COOH

dualized in than 'a' increase.

yet: H_3PO_4 , HCN , NH_4OH , C_2H_5COOH

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ARRHENIUS THEORY

- **(a)** When an electrolyte is dissolved in water it breaks into oppositely charged i.e. cations and anions this process is called ionisation.
- **(b)** The solution is electrically neutral, so the number of positive charge is equal to the number of negative charge.
- **(c)** Process of ionisation is reversible. An equilibrium is established between ionised and unionised electrolyte. Hence law of mass action can be applied to this equilibrium $AB \rightleftharpoons A^+ + B^-$

$$
K_1 = \frac{[A^+][B^-]}{[AB]}
$$

 K_1 is called ionisation constant or dissociation constant.

- **(d)** Electricity does not help in ionisation of electrolyte. It only causes migration of cation of cathode & anion to anode.
- The fraction of total amount of an electrolyte which gets ionised is called its degree of ionisation or degree of dissociation **STUDYMATERIAL: CHEMISTRY**

v does not help in ionisation of electrolyte. It only

gration of cation of cathode & anion to anode.

on of total amount of an electrolyte which gets

called its degree of ionisation or degree **STUDYMATERIAL: CHEMISTRY**

y does not help in ionisation of electrolyte. It only

igration of cation of cathode & anion to anode.

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 M

tricity does not help in ionisation of electrolyte. It only

ses migration of cation of cathode & anion to anode.

fraction of total amount of an electrolyte which gets

fraction of total

- The equivalent conductance of a solution depends upon the number of ions and their charge in solution.
- **(g)** The properties of solution governed by the ions present in the solution i.e. colour of solution.
- The colligative properties are effected by ionisation and depend upon number of cations and anions in the solution. **Note:** (i) α is always less than 1

(ii) Degree of dissociation is assumed 100% at infinite dilution.

Factors Responsible for degree of ionisation :

- **Dilution :** When dilution is increased in a weak electrolyte, the degree of ionisation increases and at infinite dilution any substance gets almost completely ionised.
- **2. Temperature :**Increase of temperature increases the degree of ionisation, and decrease of temperature decreases the degree of ionisation.

Kinetic energy of ions increases on increasing temperature and attractive forces between oppositely charged ions become less.

- **3. Nature of solute :** Ionic character of electrolyte increase than α' increase. e.g. strong acid and base is totally ionised.
- Hytes in aq. medium of the following substance.

HCl, NaCl, NH₂CONH₂ (times), H₂SO₄, H₂P_{O4}, and attractive forces between op

O₁₁ (sugar), C₂H₃COOH, NH₄Cl, HNO₃, and attractive forces between op

O₁ **4. Nature of Solvent :** Degree of ionisation depends on the nature of solvent. Solvent molecules weakening the forces of attraction is called dielectric constant. Higher the polarity of the solvent molecule, more is its dielectric constant. Thus degree of ionisation of an electrolyte is high in a solvent of high dielectric constant.

Degree of ionisation of a compound in different solvents is different, because their dielectric constant are different.

5. Present of common ion : The extent or degree of dissociation of an electrolyte is lowered in the presence of a strong electrolyte having a common ion.

Example 2 :

Which is greater α_1 or α_2 for the following equation : (i) $NH_4OH + H_2O \rightarrow \alpha_1$ (ii) $NH_4OH + D_2O \rightarrow \alpha_2$

Sol. Dielectric constant of D₂O is less than the H₂O so $\alpha_1 > \alpha_2$

IONIC EQUILIBRIUM

OSTWALD DILUTION LAW

Strong electrolytes are almost completely ionised, but the weak electrolytes are partially ionised. Therefore, Ostwald put forward the law for weak electrolytes.

Degree of ionisation can be determined by applying the law of mass action on dilute solutions of weak electrolytes. On the basis of that, equilibrium constant and ionisation constant can be determined.

Take an example of dissociation of a weak acid HA, which takes place as follows.

> $HA \rightleftharpoons H^+ + A^-$ 1 0 0 Initial mole $1-\alpha$ α α Mole at equilibrium

If volume is v, then

1 v v V α (*x* α *x* α Mole at equilibrium

Therefore, $K_a = \frac{L}{|HA|} = \frac{N}{v} \times \frac{N}{v} \times \frac{1}{1-\alpha}$ HCl+H₂O $\times \frac{a}{v} \times \frac{a}{1-\alpha}$ HCl+H₂¹

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

 K_a is called dissociation constant of the weak acid. Value of x is very low for weak electrolytes, thus $1 - x = 1$. Then **(B)**

$$
K_a = \frac{x^2}{v} \quad ; \quad x^2 = K_a \times v
$$

$$
x \propto \sqrt{v} \quad \text{or} \quad x \propto \sqrt{1/C}
$$

s place a following to the sealing of $X_a = \frac{\alpha^2}{(1-\alpha)y}$, $X_a = \sqrt{v}$ or $x = \sqrt{1/C}$
 $X_a = \frac{2}{v} \times \frac{2}{v} = K_a \times v$
 $X_b = \frac{2}{v} \times 2 = K_a \times v$
 $X_b = \frac{2}{v} \times 2 = K_a \times v$
 $X_b = \frac{2}{v} \times 2 = K_a \times v$
 $X_b = \frac{2}{v} \times 2 = K_a \times v$
 $X_b = \frac{2}{v} \$ Therefore degree of ionisation of any substance is proportional to the square root of dilution. Thus degree of ionisation increases when dilution is increased. A substance tends to get completely ionised on infinite dilution.

Degree of ionisation increases on decreasing concentration.

- Application of Dilution Law : Dissociation constant 'K' and molar concentration 'c' of the electrolyte being known, ' α ' can be calculated by using the relation.
- **Limitations of Ostwald's dilution law :** It holds good any for weak electrolytes and falls completely in the case of strong electrolytes.

Example 3 :

What concentration of acetic acid is needed to give a hydrogen ion concentration of 3.5×10^{-4} M ?

$$
\begin{array}{c}\n\text{(K}_{\text{a}} = 1.8 \times 10^{-5})\\ \n\text{(A) } 3.5 \times 10^{-4} \text{M}\n\end{array}
$$

$$
3.5 \times 10^{-4} \,\mathrm{M} \tag{B} \,6.80 \times 10^{-3} \,\mathrm{M}
$$

(C)
$$
4.2 \times 10^{-4}
$$
 M (D) 7.2×10^{-4} M

Sol. (B). Let the concentration of acetic acid be C.

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

C-C α C α C α

$$
K_a = 1.8 \times 10^{-5} = \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} = C\alpha^2
$$

EBRICM
\nLITION LAW
\nctrolytes are almost completely ionised, but the
\nctrolytes are partially ionised. Therefore, Ostwald
\n'd the law for weak electrolytes.
\nand the derivation can be determined by applying the
\nsraction on dilute solutions of weak electrolytes.
\nand the determination constant and ionisation
\nand be determined.
\n
$$
\sqrt{C} = \frac{3.5 \times 10^{-4}}{\sqrt{1.8 \times 10^{-5}}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
$$
\n
$$
\sqrt{C} = \frac{3.5 \times 10^{-4}}{\sqrt{1.8 \times 10^{-5}}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
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\n
$$
\sqrt{C} = \frac{3.5 \times 10^{-4}}{\sqrt{1.8 \times 10^{-5}}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
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\n
$$
\sqrt{C} = \frac{3.5 \times 10^{-4}}{1.8 \times 10^{-5}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
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\n
$$
\sqrt{C} = \frac{3.5 \times 10^{-4}}{1.8 \times 10^{-5}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
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\n
$$
\sqrt{C} = \frac{6.80 \times 10^{-3} \text{ mol L}^{-1}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
$$
\n
$$
\sqrt{C} = \frac{1.25 \times 10^{-4}}{1.8 \times 10^{-5}} = \frac{1.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
$$
\n
$$
\sqrt{C} = \frac{6.80 \times 10^{-3} \text{ mol L}^{-1}}
$$

$$
C = \frac{(3.5 \times 10^{-4})^2}{1.8 \times 10^{-5}} = \frac{12.25 \times 10^{-8}}{1.8 \times 10^{-5}} = 6.80 \times 10^{-3} \text{ mol L}^{-1}
$$

CONCEPT OFACIDSAND BASES

1. ARRHENIUS CONCEPT

According to Arrhenius concept

 α α **V** given H in aqueo α **v** given H⁺ in aqueous solution. **(A) Acid :** The acid is defined as those compounds which

 $\frac{HCl + H_2O}{\text{acid}} \rightleftharpoons H^+ + Cl^-$

- **(i) Strong acids :** Almost completely ionized in water. Hence give more H^+ ions. e.g. : HCl, HNO_3 , H_2SO_4 , etc.
- **(ii) Weak acids :** Feebly ionized in water. Hence give less H^+ ions. e.g. :HCN, CH₃COOH, H₂CO₃, H₂S etc.
- **Base :** The base defined are those compounds which give OH– in aqueous solution.

 $NaOH + H₂O \rightleftharpoons Na⁺ + OH⁻$
Base

- **(i) Strong bases :** Almost completely ionized in water. Hence give more OH⁻ ions. e.g. : NaOH, KOH, RbOH etc.
- **(ii) Weak bases :** Feebly ionized in water. Hence give less OH⁻ ions. e.g. : NH₄OH, C₂H₅NH₃⁺, OH⁻, Ba(OH)₂, $\mathrm{Al(OH)}_{3}$, Ca $\mathrm{(OH)}_{2}$, etc.

Limitations of Arrhenius Theory :

- **(i)** It fails to explain the behaviour of acids and bases in non–aqueous solutions.
- **(ii)** It fails to explain the neutralization giving salt in the absence of a solvent e.g.,

 $CO_2 + CaO \rightarrow CaCO_3$; NH₃ + HCl \longrightarrow NH₄Cl

- **(iii)** It fails to explain the acidic character of certain salts like $AICl_3$, BF_3 and basic character of NH_3 , PH_3 etc.
- (iv) It fails to explain as to how H^+ exists in water.

Strength of acid or base:

o get completely ionised on infinite dilution.

a C or ionisation increases on decreasing

a A(OH)₃, Ca(OH)₂, Ca(OH)₂, etc.

ation of Dilution Law : Dissociation constant K' (b) It fails to explain the ehactwise t ion increases on decreasing

in the interdions of Arrhenius Theory:

on Law: Dissociation constant K'

(i) It fails to explain the behaviour of acids

(ii) It fails to explain the behaviour of acids

(ii) It fails to expl Proto foliation. It
as degree of the foliocol and the H₂ on Sec and Cali (ii) Weak bases: Feebly ionized in water. H

onised on infinite dilution.

In increases on decreasing
 $\mathbf{A}(\mathbf{H})$, Ca(OH)₂, Ca(OH)₂, Ca(OH in is increased. A substance

on infinite dilution.

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cases on decreasing

Limitations of Arrientials Theory:

Dissociation constant K'

(i) It fails to explain the behaviour o under is increased. Associated on the case of $\frac{1}{2}$ and $\frac{1}{2$ **(i)** Strength of acid or base depends on the extent of its ionisation. Hence equilibrium constant K_a or K_b respectively of the following equilibria gives a quantitative measurement of the strength of the acid or base. in water. Hence give less
 $H_5NH_3^+$, OH⁻, Ba(OH)₂,
 $V:$

wur of acids and bases in

ization giving salt in the
 $+ HCl \longrightarrow NH_4Cl$

character of certain salts

racter of NH₃, PH₃ etc.
 $+$ exists in water.

ends on th orientally 1, 0.01 and the SI

oriental and bases in

on giving salt in the

CI \longrightarrow NH₄Cl

acter of certain salts

er of NH₃, PH₃ etc.

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constant K_a or K_b

equilibria giv of acids and bases in

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Cl \longrightarrow NH₄Cl

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ter of NH₃, PH₃ etc.

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s on the extent of its

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strength of the acid
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 \rightarrow NH₄Cl

er of certain salts

of NH₃, PH₃ etc.

i in water.

the extent of its

nstant K_a or K_b

uilibria gives a

ength of the acid
 $\frac{H_3O^+][A^-]}{[HA]}$
 $=\frac{[BH^+][OH^-]}{[B]}$

(ii) HA+H₂O
$$
\rightleftharpoons
$$
 H₃O⁺+A⁻; K_a = $\frac{[H_3O^+][A^-]}{[HA]}$

(iii) Similarly,B+H₂O
$$
\rightleftharpoons
$$
 BH⁺+OH⁻; K_b = $\frac{[BH^+][OH^-]}{[B]}$

Here H_2O is solvent.

2. BRONSTED-LOWRY CONCEPT

According to this concept, the acid is defined as a substance which give up proton $\&$ the base is defined as a substance which accepts proton.

The species B accepts a proton $&$ thus behave as a base, while the species BH^+ gives up a proton & thus acts as an acids. $B + H^+ \rightleftharpoons B^+H$

(Base) (Proton) (acid)

Such related pair of an acid $\&$ a base is said to be conjugate (i) to are another so B is a conjugate base of the acid $B^+H \&$ B ⁺H is the conjugate acid of the base B. **eg. :**

$$
H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-
$$
\nAcid Base Conjugate acid Conjugate base

\nAcid Base Conjuate

\n
$$
H_2O \rightleftharpoons H_3O^+ + CO
$$
\n
$$
H_2O^+ + SO_4^{2-}
$$
\n
$$
H_2O^+ + SO_4^{2-}
$$
\n
$$
H_2O^+ + [Fe(H_2O)_5(OH)]^{2+}
$$

- Conjugate acid-base pair differ by only one proton.
- Strong acid will have weak conjugate base and vise versa. Reaction will always proceed from strong acid to weak acid or from strong base to weak base. According to bronsted – Lowry concept, neutralization is a process in which a proton is trasnfered from an acid to a base.

Conclusion of this theory are –

- **(i)** A substance acts as an acid only when another substance capable of accepting a proton is present.
- (ii) In aqueous solution H^+ ions exists an hydrated ion H^+H_2O or H_3O^+ & not as free H⁺ ion.
- **(iii)** Not only molecules but even ion may act as acids or bases.
- **(iv)** Water, the most common solvent can act both as an acid as well as a base because it can give off a proton as well as can receive. It its dual behaviour may be represented by the following equation.

Example 4 :

In which of the following reactions does $NH₃$ act as an acid?

(A) NH_3 (aq) + H⁺ (aq) \rightarrow NH₄⁺

(B) NH₃(
$$
\ell
$$
) + Na(s) \rightarrow NaNH₂(aq) + $\frac{1}{2}$ H₂(g)
(C) NH₄(α) + HCl(α) , NH₄Cl(α)

(C) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

(D) None of these

Sol. (B). In the following reaction, NH_3 changes to $NaNH_2$ which contains NH_2^- ion. This means that NH_3 has donated a proton to Na and hence acts as an acid.

Example 5 :

- Which is the strongest Bronsted base in the following anion- (A) ClO⁻ (B) ClO₂⁻ $-$ – $-$ – (C) $ClO₃⁻$ (D) $ClO₄$ $-$ – $-$ – $-$
- **Sol.** (A). HClO is weakest acid among HClO, $HCIO_2$, HClO₃ and $HClO₄$.

Behaviour of acids and bases in aqueous solution :

(i) Acids in water : Water functions as a weak base and accepts a proton from the acid, thus

 $HA + H_2O \rightleftharpoons H_3O^+ + A^-$
A[–] can be considered as the conjugate base of anion of salt of the acid HA. So an aqueous solution of an acid contains hydronium (H_3O^+) ions (and not H⁺ ions)

(ii) Strong and Weak acids : (Relationship between a conjugate pair). An acid is considered to be a strong acid, it can give up its proton readily. In an aqueous solution of HCl, there

is equilibrium HCl + H₂O \rightleftharpoons H₃O⁺ + Cl⁻
Since HCl can give up its proton readily, this equilibrium will lie far to the right. Consequently the base Cl⁻ cannot take up the proton from H_3O^+ readily, i.e., Cl^- is a weak base. In general, the conjugate base of a strong acid will be a week base. In the case of a weak acid like phenol, s proton readily. In an aqueous solution of HCl, there
ulibrium HCl + H₂O \rightleftharpoons H₃O⁺ + Cl⁺
e HCl can give up its proton readily, this equilibrium
e HCl can give up its proton readily, this equilibrium
up the pr considered to be a strong acid, it can give
dily. In an aqueous solution of HCl, there
ICI + H₂O \rightleftharpoons H₃O⁺ + Cl⁻
iCI + H₂O \rightleftharpoons H₂O⁺ + Cl⁻
ive up its proton readily, this equilibrium
or from H₃O

 $C_6H_5OH + H_2O \rightleftharpoons H_3O^+ + C_6H_5O^-$
The equilibrium will lie far to the left. So the conjugate base $(C_6H_5O^-)$ of a weak acid (C_6H_5OH) will be moderately strong. Again there is little ionization in the aqueous solution of a weak acid so the acid is weak electrolyte and the solution has a low conductance. m H₃O' readily, i.e., CI⁻ is a weak
onjugate base of a strong acid will
case of a weak acid like phenol,
 \Rightarrow H₃O⁺ + C₆H₅O⁻
ie far to the left. So the conjugate
k acid (C₆H₅OH) will be moderately
is littl n general, the conjugate base of a strong acid will
eek base. In the case of a weak acid like phenol,
 $H_2OH + H_2O \rightleftharpoons H_3O^+ + C_6H_5O^-$
quilibrium will lie far to the left. So the conjugate
 $V_6H_5O^-$) of a weak acid (C_6

Based on studies of acids it is established that,

 $HClO_4$ > HCl > HNO_3

(iii) Relationship between the ionization constants of acids and their conjugate bases :

The equilibrium reaction of $CH₃COOH$ in water is : CH₃COOH (aq) + H₂O(ℓ) \rightleftharpoons CH₃COO⁻ (aq) + H₃O⁺ (aq)

$$
K_{eq}[H_2O] = \frac{[CH_3COO^-][H_2O^+]}{[CH_3COOH]} = K_a
$$
(1)

The conjugate base of acetic ion $(CH_3CO_2^-)$ acts as base in its reaction with water as

 $CH_3COO^- + H_2O(\ell) \rightleftharpoons CH_3COOH (aq) + OH^-(aq)$

base. In general, the conjugate base of a strong actual with
be a week base. In the case of a week acid like phenol,

$$
C_6H_5OH + H_2O \rightleftharpoons H_3O^+ + C_6H_5O^-
$$

The equilibrium will lie far to the left. So the conjugate
base $(C_6H_5O^-)$ of a weak acid (C_6H_5OH) will be moderately
strong. Again there is little ionization in the aqueous
solution of a weak acid so the acid is weak electrolyte and
the solution has a low conductance.
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 $HClO_4 > HCl > HNO_3$
Relationship between the ionization constants of acids
and their conjugate bases:
 CH_3COOH (aq) + H₂O(ℓ) $\rightleftharpoons CH_3COO^-$ (aq) + H₃O⁺ (aq)
 $K_{eq}[H_2O] = \frac{[CH_3COO^-][H_2O^+]}{[CH_3COOH]}$ = K_a (1)
The conjugate base of acetic ion $(CH_3CO_2^-)$ acts as base
in its reaction with water as
 $CH_3COO^- + H_2O(\ell) \rightleftharpoons CH_3COOH$ (aq) + OH⁻ (aq)
 $K'_{eq}[H_2O] = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$ = K_b
 $[OH^-] = \frac{K_w}{[H^+]}$ (2)
 $\therefore K_b = \frac{[CH_3COOH]K_w}{[CH_3COO^-][H^+]}$; $K_b = \frac{K_w}{K_a}$
Thus, stronger the acid (the larger K_a), weaker will be its
conjugate base (the smaller K_b) and vice versa.
Therefore, for an acid-conjugate base pair.
 $K_w = K_a \times K_b$

$$
\therefore \quad K_b = \frac{\text{[CH}_3\text{COOH]}K_w}{\text{[CH}_3\text{COO}^-]\text{[H}^+]}; \quad K_b = \frac{K_w}{K_a}
$$

 2^{12} Thus, stronger the acid (the larger K_a), weaker will be its conjugate base (the smaller K_b) and vice versa. Therefore, for an acid-conjugate base pair.

$$
K_w = K_a \times K_b
$$

IONIC EQUILIBRIUM

3. LEWIS CONCEPT

(A) Lewis acid : Compounds having nature of accepting electron pair are called acids. e.g. : AlCl₃, ZnCl₂, SnCl₂, SnCl₄, BF₃, Ag⁺, H⁺, Cu²⁺

Classification of Lewis acids :

- (a) Molecule with incomplete octet of its central atom e.g. : AlCl₃, BF₃, CaCl₂, FeCl₃, ZnCl₂, BeCl₂
- (b) Molecule with central atom containing empty d–orbital. e.g. : SiF_4 , SnCl_2 , SnCl_4 , PF_3 , PF_5 , SF_4 , SeF_4
- (c) Positively charged ions : e.g.: Al^{+3} , Cu^{+2} , Ag^+ , NO_2^+ , Cl^+ , Br^+ , RNH_3^- , ROH_2^+ etc.
- (d) Molecules having multiple bonds between atoms differing in electronegativities. e.g. : CO_2 , SO_2 , SO_3 etc.
- (e) Atom with electron sextet : e.g. : O, S
- **(B) Lewis base :** Compounds having nature of donating electron pair are called bases.

e.g. : OH⁻, Cl⁻, CN⁻, NH₂⁻, NH₃, RÖH, R - Ö - R

Classification of Lewis bases :

- (a) Neutral molecules : Molecules containing atom carrying lone pair of electrons. e.g. : $\ddot{\text{NH}}_3$, $\ddot{\text{RNH}}_2$, $\text{R}_2\ddot{\text{NH}}$, $\text{R}_3\ddot{\text{N}}$ \ddot{N}
- (b) Negatively charged ions : e.g. : Cl⁻, Br⁻, I⁻, C₂H₅⁻, C₂H₅O⁻, C₂H₅S⁻, etc.
- (c) Multi bonded compound which is formed coordinate compound with transition element. e.g. : CO, No, Ethylene, Acetylene etc.

CLASSIFICATION OF SOLVENTS

- **(I) Protophilic :** The solvent which have a greater tendency to accept proton e.g H_2O , alcohol, liquid ammonia. etc. e.g. : RNH_2 , Liquid NH₃, Pyridine, NH₂–NH₂ etc.
- **(II) Protogenic :** The solvent which have a greater tendency to donate proton. e.g. : $\mathrm{H}_2\mathrm{SO}_4$, $\mathrm{H}_2\mathrm{F}_2$, HCN, $\mathrm{HNO}_3\mathrm{H}_2\mathrm{O}$ etc.
- **(III) Amphoteric :** The solvents which can both accept & donate a proton. e.g. : H_2O , alcohol etc. H is always present
	- $H_2O, NH_3, HSO_4^-, OH^-, NH_2-NH_2$
- **(IV) Aprotic :** The solvent which can neither accept nor donate a proton. e.g. :Benzene, CS_2 , CCl_4 , SO_2 , COCl_2 , BrF_3 , N_2O_4 .
- **(V) Levelling solvent :** All the strong acids react always completely with water. So all strong acids in aqueous solution appear equally strong. Water so thus called as a leveling solvent because it levels all the acids to the same strength leveling effect.

Note :

- If instead of water solvent, we take mineral acids in pure acetic acid solvent (which is poor proton acceptor as compared to water) it is found acids become weak and can be differentiated.
- Ex. HCl + CH₃COOH[–] \rightleftharpoons Cl[–] + C
Acid Base Base Cl^- + CH₃COOH₂⁺
Base Acid Acid Base Base Acid * In above example acetic acid and Cl^- ions both complete for protons and the former being a poor proton acceptor does it much less effectively than water. Thus HCl in acetic acid solvent appears to be a $*$ much weaker acid than that in water.

Mineral acids in acetic acid solvent follow the following order of their strengths. $\rm HNO_3^{\phantom i}$ < $\rm HCl$ < $\rm H_2SO_4^{\phantom i}$ < $\rm HBr$ < $\rm HClO_4^{\phantom i}$ Mineral acids in acetic acid solvent follow the

following order of their strengths.

HNO₃ < HCl < H₂SO₄ < HBr < HClO₄
 ODUCT OF WATER

or is a weak electrolyte, which gets ionised to H⁺ and
 Γ to a very lo **SPONDADVANCED LEARNING**

neral acids in acetic acid solvent follow the

lowing order of their strengths.
 10_3 < HCl < H₂SO₄ < HBr < HClO₄

DUCT OF WATER

s a weak electrolyte, which gets ionised to H⁺ and
 σ

IONIC PRODUCT OFWATER

Water is a weak electrolyte, which gets ionised to H^+ and OH— to a very low extent.

 $H_2O \rightleftharpoons H^+ + OH^-$
According to law of mass action

$$
=\frac{\text{[H+][OH-}{\text{[H2O]}}
$$

 $2^{\mathbf{U}}$ Here K is called ionisation constant of water. Therefore, $K[H_2O] = [H^+][OH^-]$

 $K_W = [H^+] \tilde{[O}H^-]$ Here $K_W = K[H_2O]$

 K_W is called ionic product of water.

In pure water,
$$
[H^+] = [OH^-]
$$

 $10^{-7} \quad 10^{-7}$

Material state of accepting $\mathbb{C} \cup \mathbb{C} \cup$ pounds having nature of accepting

led acids.

SnCl₂, SnCl₄, BF₃, Ag⁺, H⁺, Cu²⁺

emis acids:

incomplete octet of its central atom

incomplete octet of its central atom

leading to the transmition the same del , R N³ Value of ionic product of water K_W is 1.0×10^{-14} at 25° C Value of K_{W} depends on temperature. Value of K_{w} increases with increase in temperature and decreases with decrease in temperature and decreases with decrease in temperature. **Note :**

Effect of Temperature

H₂O
$$
\rightleftharpoons
$$
 H⁺ + OH⁻
\nAccording to law of mass action
\n
$$
K = \frac{[H^+][OH^-]}{[H_2O]}
$$
\nHere K is called ionisation constant of water.
\nTherefore, K[H₂O] = [H⁺][OH⁻]
\n
$$
K_W = [H^+][OH^-]
$$
 Here $K_W = K[H_2O]$
\nK_W is called ionic product of water.
\nIn pure water, [H⁺] = [OH⁻]
\n10⁻⁷ 10⁻⁷
\nValue of 6 K_W depends on temperature. Value of K_w increases
\nwith increase in temperature and decreases with decrease
\nin temperature and decreases in temperature.
\nNote:
\nEffect of Temperature
\n
$$
T\rightarrow K_w\uparrow
$$

\n
$$
T\uparrow \rightarrow K_w\uparrow
$$

\n
$$
T\uparrow \rightarrow K_w\uparrow
$$

\n
$$
T\uparrow \rightarrow K_w\uparrow
$$

\n
$$
- \log [H^+] = P^H
$$

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

\n
$$
- \log [OH^-] = P^{OH}
$$

\n
$$
- \log K_w = - \log [H^+] - \log [OH^-]
$$

\n
$$
pK_w = p^H + p^{OH}
$$

\nEffect of temperature on K_w
\nTemp.
\n
$$
K_w
$$

\n
$$
0.3 \times 10^{-14}
$$

Effect of temperature on K

pHAND HYDROGEN ION CONCENTRATION

The term pH was given by Sorenson. pH is defined as the negative logarithm to the base 10 of the hydrogen ion concentration i.e. $pH = -log[H^+]$

pH lies from 0 to 14. For pure water $pH = 7$. For acidic solution pH is less than 7 and for basic solution pH is more than 7.

- Solution with $pH = 0$ is acidic
- pH value decreases with heating the solution
- $-ve$ pH value is not possible experimentally (but theoretically it indicate acidic solution)
- Value of pH of acid, base $&$ salts having same concentration is different.
- pH value of solution is determined by universal indicator, EMF method or use of pH meters.

.

- pH scale extends from 0 to 14
- * Mid point of scale is pH = 7, represent neutrality.
- pH values greater than 14 are possible for concentrated strong base and negative pH values are possible for concentrated strong acid.

Example 6 :

What should be the number of OH^- ion one litre water, if concentration of OH⁻ ion is 10^{-7} gram ions/litre -

Sol. Number of OH⁻ = Concentration of $[OH^-] \times$ Avogadro no. $= 10^{-7} \times 6.023 \times 10^{23} = 6.023 \times 10^{16}$

Example 7 :

What should be the ionic product of water, if the value of ionisation constant of water is 1.8×10^{-16} ?

Sol. $K_w = K \times [H_2O]$ $= 1.8 \times 10^{-16} \times 55.4 = 1 \times 10^{-14}$

Example 8 :

At 25 $\rm ^{\circ}C$ the [H⁺] of an acidic solution is 10⁻⁵ then find out its pH.

Sol. $pH = -\log[H^+]$ $=-\log 10^{-5} = +5 \log 10 = 5$

Example 9 :

At 25 \degree C the [OH⁻] is 10⁻⁹ then find out the value of pH of the solution.

Sol. pH + pOH = 14
\n[H⁺] [OH⁻] = 10⁻¹⁴
\n[H⁺] =
$$
\frac{10^{-14}}{10^{-9}} = 10^{-5}
$$

\nph = $-\log$ [H⁺] = $-\log$ [10⁻⁵] = + 5 log 10 = 5

STRENGTH OFACID

(i) Oxi acid (Non metal + O-H)
\n**Case I :** When non metals are different
\nAcidic character
$$
\propto
$$
 Eelectro negatively of non-metal
\ne.g. : $\begin{bmatrix} H_3BO_3 & H_2CO_3 & HNO_3 \end{bmatrix}$

 H_2SiO_3 HPO₃ H_2SO_4 HClO₄ $HASO₃$ $H₂SeO₄$ $HBrO₄$
 $H₂TeO₄$ $HIO₄$ EN 1 Acidic character J

 \longrightarrow EN \uparrow , Acidic character \uparrow **Case–II :** When non metal are similar Acidic character ∞ Oxidation Number e.g.: $HNO₂< HNO₃$ $(+3)$ $(+5)$ $\rm H_2SO_3 \leq H_2SO_4$ $\rm HClO$ < $\rm HClO_2$ < $\rm HClO_3$ < $\rm HClO_4$ $(+1)$ $(+3)$ $(+5)$ $(+7)$ **Exception :** H_3PO_2 > H_3PO_3 > H_3PO_4 $(+1)$ $(+3)$ $(+5)$ (ii) In hydride : $B_2H_6CH_4$
 $Sine_1$ $B_2H_6CH_4$ $B_2H_6CH_4$ NH₃ H₂O HF
SiH₄ PH₃ H₂S HCl SiH_4 PH₃ H₂S HCl GeH₄ AsH₃ H₂Se HBr
SnH₄ SbH₃ H₂Te HI Size 1 Acidic character \uparrow Acidic character |

m metal are similar
 NO_3
 H_2SO_4
 $\text{CO}_2 \leq \text{HClO}_3 \leq \text{HClO}_4$
 $(+5)$ (+7)
 (25) $(+7)$
 (-5) $(+7)$
 (-5)
 (-7)
 (-8) $(+6)$
 (-8)
 (-1)
 I group

 EN increase, Acidic character increase **(iii)** In carbonic acid : $R - \overline{C} - O - H$ O

Acidic character $\infty - I$ group ∞ $\frac{1}{\pm I}$ 1 e.g. : (a) H -COOH > CH₃COOH > CH₃CH₂COOH

 \longrightarrow $+I$ group \uparrow , Acidic character \downarrow (b) CH₃COOH <ICH₂COOH <BrCH₂COOH <FCH₂COOH

 \longrightarrow $-I$ group \uparrow , Acidic character \uparrow **(iv)** Explanation of mathematical acids : Acidic character \propto [H⁺] $\propto \sqrt{K_a}$ Acidic character $\uparrow = [H^+] \uparrow$, pH \downarrow , [OH⁻] \downarrow , pOH \uparrow , K_a \uparrow , $pK_a\downarrow$ **(v)** Oxides (Acidic anhydride) : H-COOH > CH₃COOH > CH₃COOH

H-COOH > CH₃COOH > CH₃CH₂COOH

+I group 1, Acidic character ↓

CH₃COOH < ICH₂COOH < BrCH₂COOH < FCH₂COOH

-

-

-I group 1, Acidic character 1

lanation of mathematical acids $H_2SO_4 \longrightarrow SO_3$ H-COOH > CH₃COOH > CH₃COOH

+I group \uparrow , Acidic character \downarrow

CH₃COOH < ICH₂COOH < BrCH₂COOH < FCH₂COOH

-I group \uparrow , Acidic character \uparrow

lanation of mathematical acids :

lic character \propto [H⁺ $H_2CO_3 \longrightarrow CO_2$ +I group 1, Acidic character \downarrow
CH₃COOH < ICH₂COOH < BrCH₂COOH < FCH₂COOH
-I group 1, Acidic character 1
lanation of mathematical acids :
lic character \propto [H⁺] $\propto \sqrt{K_a}$
lic character 1 = [H⁺] 1, pH↓, $H_2SO_3 \longrightarrow SO_2$ -1 group ↑, Acidic character ↑

lanation of mathematical acids :

lic character \propto [H⁺] $\propto \sqrt{K_a}$

lic character \uparrow = [H⁺] ↑, pH↓, [OH⁻]↓, pOH↑, K_a↑,

↓

8s (Acidic anhydride) :

H₂SO₄ ---H_{2O}→ SO₃ $2 \times HNO₃ \longrightarrow N₂O₅$

STRENGTH OF BASE

(i) Hydroxide (metal + O–H)

EN \uparrow , Basic character \downarrow

(ii) Basic strength of amine :

Basic strength $\propto +I$ group $\propto \frac{1}{-I}$ group

e.g. :
$$
(CH_3)_2NH > CH_3NH_2 > (CH_3)_3N > NH_3
$$
\n $(C_2H_5)_2NH > C_2H_5NH_2 > NH_3 > (C_2H_5)_3N$ \n(iii) Explanation of mathematical bases :

Basic character ∞ [OH⁻] ∞ $\sqrt{K_b}$ Basic character $\uparrow = [OH^-] \uparrow$, pOH \downarrow , pH \uparrow , $[H^+]\downarrow$, $K_b \uparrow$, pK_b \downarrow **(iv)** Oxides(Basic anhydride) :

Hydroxide Alkaline anhydride \longrightarrow BaO **BaO Applications of Common ion effect :** 2 × [Al(OH)³

Relative strength of acids :

$$
\frac{\text{Strength of acid 1}}{\text{Strength of acid 2}} = \sqrt{\frac{\text{K}_{a_1}}{\text{K}_{a_2}}}
$$

 $pK_a = -\log K_a$; $pK_b = -\log K_b$ Higher pK_a means lower acid strength.

Dissociation constant of poly basic acids :

$$
H_3PO_4 \rightarrow H^+ + H_2PO_4^{-1}
$$
\n
$$
(K_{a_1})
$$
\n
$$
H_2PO_4^{-1} \rightarrow H^+ + H PO_4^{-2}
$$
\n
$$
(K_{a_2})
$$
\n
$$
H PO_4^{-2} \rightarrow H^+ + PO_4^{-3}
$$
\n
$$
(K_{a_3})
$$
\n
$$
H_3 PO_4 \rightarrow 3H^+ + PO_4^{-3}
$$
\n
$$
(K_a)
$$
\n
$$
K_a = K_{a_1} \times K_{a_2} \times K_{a_3}
$$
\n
$$
K_{a_1} > K_{a_2} > K_a
$$

But pka_1 < pka_2 < pka_3 Simillarly in poly acidic base.

> $K_b = K_{b_1} \times K_{b_2} \times K_{b_3}$ the ionisation $K_{b_1} > K_{b_2} > K_{b_3}$

But $P^{K_{b_1}} < P^{K_{b_2}} < P^{K_{b_3}}$

COMMON ION EFFECT

1 CH₃COOH, it ionises as : CH₃COOH \Leftrightarrow CH₃COO[–]+H⁺ If we consider ionisation of a weak electrolyte say,

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

1 ions will continue to change in molecule and molecule will COMMONIONEFFECT

If we consider ionisation of a weak electrolyte

(EH₃COOH, it ionises as: CH₃COOH \Leftrightarrow CH₃COOH

(EH₃COOH, it ionises as: CH₃COOH \Leftrightarrow CH₃COOH

(EH₃COOH, it ionises as: CH₃COOH \geq CH H Ba(OH)₂ – T(OH)₃ Dasse character (-1 Co-1 Ca(OH)₂ – T(OH)³ Dasse character (-1 Ca(OH)₂ – CH₃OU

EN T, Basic character \downarrow

En T, Basic character (-1 constant at same temperature constant at same temperatur \overrightarrow{CaO} CaO of its common ion is added ionisation of weak electrolyte EN T, Basic character $\sqrt{k_1}$

c strength of amine :

c extength α + I group α - I group

Basic strength α + I group α - I group
 $\left(\frac{C_1H_3}{2}NH > CH_3NH_2 > CH_3N + C_2H_3N$
 $\left(\frac{C_2H_5}{2}MH > CH_3NH_2 > CH_3} \right)\right)$

s Example of a material stopped only constant it will continue to change

2 × [Al(OH)₃] = $\frac{1}{-H_2O}$ + $\frac{1}{H_2O_3}$ = $\frac{1}{(C_2H_3)N + C_2H_3N + C_3 - H_1C_3N}$

2 × [Al(OH)₃] = $\frac{1}{-H_2O}$ + $\frac{1}{H_2O_3}$ = $\frac{1}{H_$ An $\alpha + 1$ group α
 α and α and α and α and α and α and α by α and α con Assic strength α + I group α

(CH₃)₂NH > CH₃H₃N > NH₃ of CH₃H₃N > NH₃ of C₁H₃N₂N is and the CH₃CO₁ (C₃H₃)³N = C₂H₅NH₂ > NH₃ > (C₂H₃)³N = solution to increase CH₃CO₀ (CH₃)NH > CH₃NH₂>(H₃)NH > CH₃NH₂>(H₃)NH ₂⁻ (CH₃)NH > CH₃NH₂>(H₃)NH₂⁻ (H₃)NH₃⁻ (CH₃(C₂H₃)H₂⁻ 2 (H₃C₂H₃)H₂⁻ 2 (H₃C₂H₃)H₂⁻ 2 (H₃C₂ (H₃C₂ (H₃C₂ (H EFFECT

SUPERRECT

SUPERRECT

SUPERRECT

SUPERRECT

SUPERRECT

SUPERRECT

SUPERRECT

SUPERRECT

CH₃COO⁻ $\left[$ H⁺

CH₃COO⁻ $\left[$ H⁺

CH₃COO⁻ $\left[$ H⁺

CH₃COOH¹

1 come such that concentration of CH₃ **EXECT**
 CONVENOVANCED LEARNING
 EFECT
 EFECT EFFECT

SOMMOVANGED LEARNING

H, it ionises as: CH₃COOH \Leftrightarrow CH₃COO⁻ + H⁺
 $\begin{bmatrix} CH_3COO^- \end{bmatrix}$ $\begin{bmatrix} H^+ \end{bmatrix}$
 $\begin{bmatrix} CH_3COOH \end{bmatrix}$

ill come such that concentration of CH₃COOH,

and H⁺ will not get chang A stage will come such that concentration of $CH₃COOH$, $CH₃COO⁻$ and H⁺ will not get changed further, their concentrations on a definite mathematical calculation will give a constant called equilibrium constant which will remain constant at same temperature. It does not mean that reaction has stopped only concentration will remain same, continue to get ionised this is said to be Dynamic Equilibrium. Now suppose some how $CH₃COO⁻$ ions are added to the solution to increase CH_3COO^- ions concentration since mathematical calculation will tend to change itself to change that eq. constant, electrolyte (CH_3COOH) will do same thing in order to keep K_a constant because it has to be a constant it will reduce its ionisation thus $\left[CH_3COOH\right]$ will increase, $[H^+]$ will decrease and K_a will remain constant. This is called Common ion Effect that if in a solution of weak electrolyte a solution of strong electrolyte with one is suppressed.

- \longrightarrow Na₂O' (a) H₂S is used in presence of HCl as a group reagent for the entry N₂O' \longrightarrow Al₂O₃ it oxidised H₂S to sulphur. analysis of second group. $HNO₃$ cannot be used because
	- **(b)** $NH₄OH$ and $NH₄Cl$ are used as group reagent for the analysis of third group.
	- $\overline{K_{a_1}}$ saturated solution of NaCl. **(c)** Purification of common salt by passing HCl gas in a
	- **Example 10 : (d)** Soap is precipitated from its solution by adding some NaCl.

Determine the effect on the concentrations of NH_3 , NH_4^+ and OH– when small amounts of each of the following is added to a solution of $NH₃$ in water.

(a) HCl (b) NH_3 (c) NaOH (d) NH_4Cl (e) KNO₃ **Sol.** The equilibria of NH_3 in water will be

 $-(K_{a_1})$ $\qquad \qquad \text{and} \qquad \text{NH}_3 + H_2O \rightleftharpoons NH_4^+(aq.) + OH^-(aq.)$

- (K_{a_2}) (K_{a_2}) (a) If HCI is added it will full fish H and CI folds as it is a strong electrolyte. Now H⁺ will form H₂O with OH⁻ (K_{a_3}) ions (neutralisation) so $[NH_3]$ will decrease, $[NH_4^+]$ & (a) If HCl is added it will furnish H and Cl^- ions as it is a [OH–] increase since equilibria will shift in forward direction.
- $= K_{a_1} \times K_{a_2} \times K_{a_3}$. $K_{a_1} > K_{a_2} > K_{a_3}$ (b) If more NH₃ is added, then [NH₃], [NH₄⁺] & [Cl⁻] will increase.
	- (c) If NaOH is added it will furnish $[Na^+]$ and $[OH^-]$ ions since it is a strong electrolyte and [OH–] will suppress the ionisation of weak base NH_3 so $[NH_3]$ increases, [NH₄⁺] decreases and [OH⁻] increases.
	- (d) NH₄Cl will also apply common ion effect on NH₄OH but now [NH₃] of increases, [NH₄⁺] increases but [OH⁻] decreases.
	- (e) No effect expected.

pH CALCULATION OF DIFFERENTTYPES OF SOLUTIONS

(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 but greater than 10^{-8} M. In this case H⁺ ions coming from water cannot be neglect. So, $[H^+]$ = normality of strong acid $+ H^{+}$ ions coming from water in presence of this strong acid.
- **(iii)** If concentration is less than 10^{-8} M, in this case H⁺ ions coming from water is considered only.

Example 11 :

Calculate pH of 10^{-8} M HCl solution.

Sol. $H_2O \rightleftharpoons H^+ + OH^-$
 $10^{-8} + x + OH^ 10^{-8} + x$ $k_{\rm w} = [H^+] [OH^-]$ $10^{-14} = x (x + 10^{-8})$ $x^2 + x \times 10^{-8} - 10^{-14} = 0$ **EXERCISE SOLUTIONS** (d) **pH of mixture of two strong bases:**
 Example 16 16 16 16 16 16 16 16 16 16 16 16 17 17
 Example 17 in somoning from water can be neglected.
 EXAMPLE SOLUTIONS (d) **PH of mixture of values** $x = \frac{10 - 940 + 1000}{2} = 0.95 \times 10^{-7}$ **EXERIME TRION OF DIFFERENT TYPES OF SOLUTIONS** (d) **pH of mixture of two strong bases:**

SILUDY MALERIAL: C

care H¹ in scoming from water can be neglected.

If $T =$ mormality of strong acid solution.

If $T =$ mormalit $=\frac{-10 \pm \sqrt{10} + 4 \times 10}{10} = 0.95 \times 10^{-7}$ $[H^+] = 10.5 \times 10^{-8} = 1.05 \times 10^{-7}$ $pH = -\log[H^+]$; $pH = 7 - \log 1.05 \approx 7$

(b) Strong base solution :

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

Example 12 :

Calculate pH of 10^{-7} M 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^+ $(x+10^{-7})$ x $K_{\text{av}} = [H_{7}^{+}] [OH^{-}] = 10^{-14} = x (x + 10^{-7})$ $x^{2} + 10^{-7} \times -10^{-14} = 0$ x = $\frac{-10^{-8} \pm \sqrt{10^{-16} + 4 \times 10^{-14}}}{2}$ = 0.95 × 10⁻⁷

H⁺] = 10.5 × 10⁻⁸ = 1.05 × 10⁻⁷

H⁺] = 10.5 × 10⁻⁸ = 1.05 × 10⁻⁷

HH = -log [H⁺]; pH = 7 -log 1.05 × 7

and the (H+T) which will be equal to normali +x × 10⁻⁸ - 10⁻¹ + 0.61

= $\frac{-10^{-3.8} \pm \sqrt{10^{-16} + 4 \times 10^{-14}}}{1}$ = 0.95 × 10⁻⁷

+ $\frac{1}{2}$ = 0.5 × 10⁻³ = 1.05 × 10⁻⁷

+ $\frac{1}{2}$ = 0.5 × 10⁻⁷

+ $\frac{1}{2}$ = 0.5 × 10⁻⁷

= $\frac{1}{2}$ = 0.818 × 10⁻⁷

= 10-7)
 10^{-7}
 $10^{-14} = 0$
 $-10^{-14} =$

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

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

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

\n
$$
x = \frac{\sqrt{5} - 1}{2} \times 10^{-7} = 0.618 \times 10^{-7} \quad (\sqrt{5} = 2.2336)
$$

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

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

\n
$$
[OH^{-}] = N = \frac{N_{1}V_{1}}{V_{1}}
$$

\n
$$
[OH^{-}] = N = \frac{N_{2}V_{2}}{V_{1}}
$$

\n
$$
[H N_{2}V_{2} > N_{1}V_{1}; [OH^{-}] = N = \frac{N_{2}V_{2}}{V_{2}}
$$

\n
$$
[H N_{2}V_{2} > N_{1}V_{1}; [OH^{-}] = N = \frac{N_{2}V_{2}}{V_{2}}
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\n
$$
[H N_{2}V_{2} > N_{1}V_{1}; [OH^{-}] = N = \frac{N_{2}V_{2}}{V_{2}}
$$

\n
$$
[H N_{2}V_{2} > N_{1}V_{1}; [OH^{-}] = N = \frac{N_{2}V_{2}}{V_{2}}
$$

\n
$$
[H N_{2}V_{2} > N_{1}V_{1}; [OH^{-}] = N = \frac{N_{2}V_{2}}{V
$$

(c) \mathbf{p} **H** 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_rV_1 Number of H⁺ ions from I-solution =N_IV₁ 1 2

Number of H⁺ ions from II-solution = N_2V_2

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

$$
NV = N_1V_1 + N_2V_2
$$

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

$$
[H^+] = N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}
$$

(d) pH of mixture of two strong bases: Similar to above calculation

$$
[OH^{-}] = N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2} ; [H^{+}] = \frac{10^{-14}}{[OH^{-}]}
$$

Example 13 :

STUDYMATERIAL: CHEMISTRY
 ostrong bases:

ulation
 $\frac{1V_1 + N_2V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$

IOH is mixed with 500 mL of

OH)₂. To the resulting solution 99 L **TUDY MATERIAL: CHEMISTRY**
 rong bases:

tation
 $\frac{1 + N_2 V_2}{1 + V_2}$; [H⁺] = $\frac{10^{-14}}{[OH^-]}$

H is mixed with 500 mL of

(l)₂. To the resulting solution 99 L

te pH.
 $\frac{1}{1 + N_2 V_2}$ **STUDY MATERIAL: CHEMISTRY**

of mixture of two strong bases:

lar to above calculation
 $[OH^-] = N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$

3:

mL of 10^{-5} M NaOH is mixed with 500 mL of
 $\times 10^{-5}$ M of Ba(OH)₂. STUDYMATERIAL: CHEMISTRY
strong bases:
lation
 $\frac{V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$
DH is mixed with 500 mL of
 H)₂. To the resulting solution 99 L
ate pH.
 $_{2}$ =2 × 2.5 × 10⁻⁵ **STUDYMATERIAL: CHEMISTRY**

re of two strong bases:

sove calculation

= $N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$
 $0^{-5} M NaOH$ is mixed with 500 mL of

1 of Ba(OH)₂. To the resulting solution 99 L RIAL: CHEMISTRY
 $[H^+] = \frac{10^{-14}}{[OH^-]}$

the 500 mL of

sulting solution 99 L
 -5 **HEMISTRY**
 $\frac{10^{-14}}{[OH^-]}$
L of
olution 99 L 500 mL of 10^{-5} M NaOH is mixed with 500 mL of 2.5×10^{-5} M of Ba(OH)₂. To the resulting solution 99 L water is added. Calculate pH. **STUDYMATERIAL: CHEMISTRY**
 pH of mixture of two strong bases:

Similar to above calculation
 $[OH^-] = N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$
 ple 13 :

500 mL of 10⁻⁵ M NaOH is mixed with 500 mL of
 $2.5 \times$ **STUDYMATERIAL: CHEMISTRY**

wo strong bases:

alculation
 $\frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$

NaOH is mixed with 500 mL of

a(OH)₂. To the resulting solution 99 L

lculate pH.

DH)₂ = 2 × 2.5 × 10⁻⁵
 STUDYMATERIAL: CHEMISTRY
 f mixture of two strong bases:

lar to above calculation
 $[OH^-] = N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$; $[H^+] = \frac{10^{-14}}{[OH^-]}$

3:

mL of $10^{-5} M$ NaOH is mixed with 500 mL of
 $10^{-5} M$ of Ba(OH)₂. To

Sol. Normality of Ba $(OH)_2 = 2 \times 2.5 \times 10^{-5}$

$$
[OH^-] = \frac{500 \times 10^{-5} + 500 \times 2 \times 2.5 \times 10^{-5}}{1000} = 3 \times 10^{-5} \,\mathrm{M}
$$

 $M_1 = 3 \times 10^{-5} M$, $V_1 = 1L$, $V_2 = 100 L$

No. of moles of $[OH^-]$ initially = no. of moles of $[OH^-]$, in final solution.

$$
3 \times 10^{-5} = M_2 \times 100
$$

\n∴ M₂ = 3 × 10⁻⁷ < 10⁻⁶
\nH₂O ⇒ H⁺ + OH⁻
\nx (x + 3 × 10⁻⁷)
\nK_w = x (x + 3 × 10⁻⁷) = 10⁻¹⁴

Similar to above calculation
\n[OH⁻] = N =
$$
\frac{N_1V_1 + N_2V_2}{V_1 + V_2}
$$
; [H⁺] = $\frac{10^{-14}}{[OH^-]}$
\nrule 13:
\n500 mL of 10⁻⁵ M NaOH is mixed with 500 mL of
\n2.5 × 10⁻⁵M of Ba(OH)₂. To the resulting solution 99 L
\nwater is added. Calculate pH.
\nNormally of Ba(OH)₂ = 2 × 2.5 × 10⁻⁵
\n[OH⁻] = $\frac{500 \times 10^{-5} + 500 \times 2 \times 2.5 \times 10^{-5}}{1000} = 3 \times 10^{-5}$ M
\nM₁ = 3 × 10⁻⁵ M, V₁ = 1L, V₂ = 100 L
\nNo. of moles of [OH⁻] initially = no. of moles of [OH⁻], in
\nfinal solution.
\n3 × 10⁻⁵ = M₂ × 100
\n \therefore M₂ = 3 × 10⁻⁷ < 10⁻⁶
\nH₂O \rightleftharpoons H⁺ + OH⁻
\nx (x + 3 × 10⁻⁷)
\nK_w = x (x + 3 × 10⁻⁷) = 10⁻¹⁴
\n \therefore x = $\left(\frac{\sqrt{13}-3}{2}\right) \times 10^{-7} = 0.302 \times 10^{-7}$
\n[OH⁻]_{net} = $\left[3 + \frac{\sqrt{13}-3}{2}\right] \times 10^{-7} = \left[\frac{3 + \sqrt{13}}{2}\right] \times 10^{-7}$
\n= 3.302 × 10⁻⁷
\nph of mixture of a strong acid and a strong base:
\nAcid base neutralisation reaction will take place.
\nThe solution will be acidic or basic depending on which

(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 is taken in excess. If V_1 volume of a strong acid solution of normality N₁ is mixed with V₂ volume of a strong base solution normality N_2 , then Number of H⁺ ions from I-solution = N_1V_1 Number of OH⁻ ions from II-solution = N_2V_2 x $(x+3 \times 10^{-7})$
 $K_w = x (x+3 \times 10^{-7}) = 10^{-14}$
 \therefore $x = \left(\frac{\sqrt{13}-3}{2}\right) \times 10^{-7} = 0.302 \times 10^{-7}$
 $[OH^-]_{net} = \left[3 + \frac{\sqrt{13}-3}{2}\right] \times 10^{-7} = \left[\frac{3 + \sqrt{13}}{2}\right] \times 10^{-7}$
 $= 3.302 \times 10^{-7}$
 $= 3.302 \times 10^{-7}$
 PH of mixture of a s $\begin{aligned}\n& 10^{-7} \\
&= \left[\frac{3 + \sqrt{13}}{2} \right] \times 10^{-7}\n\end{aligned}$ I a strong base:

iill take place.

c depending on which
 V_1 volume of a strong

ed with V_2 volume of a

, then

= N₁V₁

on = N₂V₂
 $\frac{1V_1 - N_2V_2}{V_1$ $\left[\frac{3+\sqrt{13}}{2}\right] \times 10^{-7}$
strong base:
take place.
lepending on which
volume of a strong
with V₂ volume of a
nen
 N_1V_1
= N_2V_2
 $\frac{1-N_2V_2}{N_1+V_2}$
 $\frac{V_2-N_1V_1}{V_1+V_2}$ On
 $(x + 3 \times 10^{-7})$
 $(x + 3 \times 10^{-7}) = 10^{-14}$
 $\times 10^{-7} = 0.302 \times 10^{-7}$
 $\frac{3}{2} \times 10^{-7} = \left[\frac{3 + \sqrt{13}}{2} \right] \times 10^{-7}$
 10^{-7}
 12×10^{-7}
 $12 \times 10^{-$ 10⁻⁷
 $\left[\frac{3+\sqrt{13}}{2}\right] \times 10^{-7}$
 a strong base:

Il take place.

depending on which

1 volume of a strong

d with V₂ volume of a

then
 N_1V_1
 $n = N_2V_2$
 $V_1 - N_2V_2$
 $V_1 + V_2$
 $2V_2 - N_1V_1$
 $V_1 + V_2$ $\times 10^{-7}$
 $= 10^{-14}$
 $= 0.302 \times 10^{-7}$
 $\times 10^{-7} = \left[\frac{3 + \sqrt{13}}{2} \right] \times 10^{-7}$

acid and a strong base:

eaction will take place.

c or basic depending on which

cess. If V₁ volume of a strong

N₁ is mixed with V [OH⁻]_{net} = $\left[3 + \frac{\sqrt{13}-3}{2}\right] \times 10^{-7} = \left[\frac{3+\sqrt{13}}{2}\right] \times 10^{-7}$

= 3.302 × 10⁻⁷

= 3.302 × 10⁻⁷

Acid base neutralisation reaction will take place.

The solution will be acidic or basic depending on which

comp 1
 1 strong base:

Il take place.

depending on which

1 volume of a strong

d with V₂ volume of a

then
 ${}^{1}N_{1}V_{1}$
 $n = N_{2}V_{2}$
 $V_{1} - N_{2}V_{2}$
 $V_{1} + V_{2}$
 $2V_{2} - N_{1}V_{1}$
 $V_{1} + V_{2}$ trong base:

ake place.

epending on which

volume of a strong

with V_2 volume of a

en
 $\begin{bmatrix} V_1 \\ V_1 \\ V_2 \end{bmatrix}$
 $-\frac{N_2V_2}{V_2}$
 $+\frac{V_2}{V_2}$
 $\begin{bmatrix} V_2 - N_1V_1 \\ V_1 + V_2 \end{bmatrix}$
 $\begin{bmatrix} M & Ba(OH)_2 + N_1 \end{bmatrix}$ ×10 ' = 0.302 × 10⁻⁷
 $\frac{3}{2}$ > 10⁻⁷ = $\left[\frac{3+\sqrt{13}}{2}\right]$ × 10⁻⁷

02 × 10⁻⁷
 trong acid and a strong base:

tation reaction will take place.

e acidic or basic depending on which

in excess. If V₁ volume of $\left[\frac{3+\sqrt{13}}{2}\right] \times 10^{-7}$
strong base:
take place.
depending on which
volume of a strong
with V₂ volume of a
nen
 N_1V_1
= N_2V_2
 $\frac{V_1 - N_2V_2}{V_1 + V_2}$
 $\frac{V_2 - N_1V_1}{V_1 + V_2}$ = 0.302 × 10⁻⁷

× 10⁻⁷ = $\left[\frac{3+\sqrt{13}}{2}\right]$ × 10⁻⁷

-7

cid and a strong base:

action will take place.

or basic depending on which

ess. If V₁ volume of a strong

^{[1}] is mixed with V₂ volume of a

lility N ation win the actual of the state plenting on window

ent is taken in excess. If V_1 volume of a strong

then of hormality N_1 is mixed with V_2 volume of a

ase solution normality N_2 , then

of H⁺ ions from I-s Since the water in text, in the view of the state of solution of normality N_1 is mixed with V_2 volume of a
strong base solution normality N_1 is mixed with V_2 volume of a
strong base solution normality N_2 , t to assecutations and calculate of the scale.

a solution will be acide or basic depending on which

mponent is taken in excess. If V_1 volume of a strong

a solution of normality N_1 is mixed with V_2 volume of a

m

If N₁V₁ > N₂V₂;
$$
[H^+] = N = \frac{N_1V_1 - N_2V_2}{V_1 + V_2}
$$

Solution will be acidic in nature.

If N₂V₂ > N₁V₁; [OH⁻] = N =
$$
\frac{N_2V_2 - N_1V_1}{V_1 + V_2}
$$

Solution will be basic in nature.

, **Example 14 :**

Calculate pH of mixture of (400 mL, $\frac{1}{200}$ M Ba(OH)₂ + $(400 \text{ mL}, \frac{1}{50} \text{M} \text{ HCl}) + (200 \text{ mL of water})$

Sol.
$$
[H^+] = \frac{400 \times \frac{1}{50} - 400 \times \frac{1}{200} \times 2}{1000} = 4 \times 10^{-3}
$$

So, pH = 3 – 2 log 2 = 2.4

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IONIC EQUILIBRIUM

(f) pH of a weak acid (monoprtoic) solution :

Weak acid does not dissociated 100% therefore we have to calculate the percentage dissociation using (K_a) dissociation constant of the acid. **M**

(**monoprtoic) solution :**

ot dissociated 100% therefore we have

percentage dissociation using (K_a)
 $\begin{array}{ccc}\n\text{The weak acid will dis}\n\end{array}$
 $\begin{array}{ccc}\n\text{The weak acid will dis}\n\end{array}$
 $\begin{array}{ccc}\n\text{H A} & \rightleftharpoons \\
\text{P} & \leftarrow \\
\text{L1} & \leftarrow \\
\text{C}_2(1-\alpha) & \left$

Using Ostwald's dilution law

$$
HA \rightleftharpoons H+ + A-
$$

t=0 C 0 0
At eq. C(1- α) C α C α

$$
K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}
$$

If
$$
\alpha \ll 1 \Rightarrow (1 - \alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}
$$

(is valid if $\alpha \le 0.5$)

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

On increasing the dilution $\Rightarrow C \downarrow = a \uparrow$ and $[H^+] \downarrow \Rightarrow pH \uparrow$

Example 15 :

Calculate pH of (a)10⁻¹ M CH₃COOH (b)10⁻³ M CH₃COOH t $K_a = 2 \times 10^{-5}$

At eq.
$$
C(1-\alpha)
$$
 C α
\n
$$
K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}
$$
\n
$$
K_a = (C_2\alpha + C_1)\alpha
$$
\n
$$
K_a = 2 \times 10^{-5}
$$
\n
$$
K_a = 2 \times 10^{-5}
$$
\n
$$
K_a = 2 \times 10^{-5}
$$
\n
$$
K_a = \frac{C\alpha^2}{\alpha} \Rightarrow \alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2 \times 10^{-5}}{10^{-1}}} = \sqrt{2 \times 10^{-4}}
$$
\n
$$
K_a = \frac{C\alpha^2}{\alpha} \Rightarrow \alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2 \times 10^{-5}}{10^{-1}}} = \sqrt{2 \times 10^{-4}}
$$
\n
$$
K_a = (C_2(1 - \alpha) - C_1\alpha)
$$
\nSo, $[H^+] = 10^{-1} \times \sqrt{2} \times 10^{-2} \Rightarrow pH = 3 - \frac{1}{2} \log 2 = 2.85$ \n
$$
K_a = \frac{(C_2\alpha_2 + C_1C_1\alpha)}{C_2(1 - \alpha_2)} = \frac{C_2\alpha_2 + C_1C_1\alpha}{C_2(1 - \alpha_2)} = \frac{C_2\alpha_2 + C_1C_
$$

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

Since α , α both

(b)
$$
\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2 \times 10^{-5}}{10^{-3}}} = \sqrt{2 \times 10^{-2}}
$$
 $(\alpha > 0.1)$ $K_{a_1} = K_{a_2}$

Since α > 0.1, exact calculation is required.

$$
K_a = \frac{C\alpha^2}{1-\alpha} \Rightarrow 2 \times 10^{-5} = \frac{10^{-3} \times \alpha^2}{1-\alpha} \Rightarrow \alpha = 13.14\%
$$

[H⁺] = 10⁻³ × 0.1314 = 1.314 × 10⁻⁴
 \Rightarrow pH = 4 - log (1.314) ≈ 3.8

(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 $(SA = Strong acid)$; (WA = Weak acid)

The weak acid will dissociate as follows,

ILIBRIUM	Self				
REAR acid (monoprotoi) solution :	The weak acid will dissociate as follows, late the percentage dissociation using (K_a)	The weak acid will dissociate as follows, late the percentage dissociation using (K_a)	C_2 0 0 0 C $(1-\alpha)$ 0 $C_2\alpha + C_1$ 0 0 C $(1-\alpha)$ 0 $C_2\alpha$ 0 0		
$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}$	$K_a = \frac{(C_2\alpha + C_1)C_2\alpha}{C_2(1-\alpha)}$ 0 $(\alpha \lllt; \lt; 1)$				
$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}$	The weak acids dissociation will be further suppressed because of presence of strong acid, common ion effect) because of presence of strong acid, common ion effect) $K_a = (C_2\alpha + C_1)\alpha$				
$\Rightarrow (1-\alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$	$10 + i\text{ non concentration} = C_1 + C_2\alpha$ $\Rightarrow (1-\alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$	$10 + i\text{ non concentration} = C_1 + C_2\alpha$ $\Rightarrow (1-\alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$	$10 + i\text{ on concentration}$ 0 $-\text{C}_1 + \text{C}_2\alpha$ $\Rightarrow (1-\alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$	$10 + 10 + i\text{ on concentration}$ 0 $-\text{C}_1$ $\Rightarrow (1-\alpha) \approx 1 \Rightarrow K_a = C\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$	10

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

 $K_a = (C_2 \alpha + C_1) \alpha$ To H⁺⁺ ion concentration = C₁ + C₂ α

 $2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$ then contribution from the water can be neglected, if comparable then we take [H⁺] from water also. **Note :** If the total $[H^+]$ from the acid is more than $10^{-6}M$, comparable then we take $[H^+]$ from water also.

(h) pH of a mixture of two weak acid (both monoprotic) solution:

Both acids will dissociate partially.

* Let the acid are $HA_2 \& HA_2$ and their final concentrations are C_1 & C_2 respectively, then

2 5 C 2 10 ^K K 2 10 1 C ¹⁰ ^K 2 10 2 10 (> 0.1) 2 3 2 1 1 HA¹ H+ + A¹ – t = 0 C¹ 0 0 At eq. C¹ (1 – ¹) C1¹ + C2² C1¹ 1 1 1 1 2 2 a1 1 1 C (C C) ^K C (1) HA² H+ + A² – t = 0 C¹ 0 0 At eq. C² (1 – ²) C2² + C1¹ C2² 2 2 1 1 2 2 2 1 (C C) C ^K C (1) a 1 1 2 2 1 a 1 1 2 2 2 1 2 K (C C) : K (C C) a1 1 a 2 ²

$$
K_{a_2} = \frac{(C_2 \alpha_2 + C_1 \alpha_1) C_2 \alpha_2}{C_2 (1 - \alpha_1)}
$$

 2^{18} 2.02
(Since α_1 , α_2 both are small in comparision to unity)

$$
K_{a_1} = (C_1\alpha_1 + C_2\alpha_2) \alpha_1 : K_{a_2} = (C_1\alpha_1 + C_2\alpha_2) \alpha_2
$$

\n increasing the dilution
$$
\Rightarrow C \psi = a \mid \text{and } [H^+]\psi \Rightarrow pH \quad \text{Let } \alpha \text{ and } \alpha \text{ and } \alpha \text{ and } \alpha \text{ are } C_1 \& C_2 \text{ respectively, then}
$$

\n\n 5: \n $\text{ar } C_1 \& C_2 \text{ respectively, then}$ \n

\n\n 5: \n $\text{ar } C_1 \& C_2 \text{ respectively, then}$ \n

\n\n 6: \n $\text{ar } C_1 \& C_2 \text{ respectively, then}$ \n

\n\n 7: \n $\text{ar } C_1 \& C_2 \text{ is } C_2 \text{ is }$

Note : If the dissociation constant of one of the acid is very much greater than that of the second acid then contribution from the second acid can be neglected. So, $[H^+] = C_1 \alpha_1 + C_2 \alpha_2 = C_1 \alpha_1$

Calculate pH of solution obtained by mixing vol. of 0.02 M HOCl $& 0.2$ M CH₃COOH solution given that

$$
K_{a_1}
$$
[HOCI] = 2×10⁻⁴; K_{a_2}[CH₃COOH] = 2×10⁻⁵

also calculate OH⁻, OCl⁻, CH₃COO⁻.

Sol. Final solution volume become double $C_1 = 0.01, C_2 = 0.1$

$$
[H^+] = \sqrt{K_{a_1}C_1 + K_{a_2}C_2}
$$

= $\sqrt{2 \times 10^{-4} \times 0.01 + 2 \times 10^{-5} \times 0.1}$
= $\sqrt{2 \times 10^{-6} + 2 \times 10^{-6}} = 2 \times 10^{-3}$
pH = 3 - log 2 = 3 - 0.3010 = 2.69

$$
\alpha_1 = \frac{2 \times 10^{-4}}{2 \times 10^{-3}} = 10^{-1} \quad ; \quad \alpha_2 = \frac{2 \times 10^{-5}}{2 \times 10^{-3}} = 10^{-2}
$$

HOC1 \rightleftharpoons H⁺ + OCI⁻
C₁ (1- α_1) C₁ α_1 + C₂ α_2 C₁ α_1
[OCI⁻] = C₁ α_1 = 0.01 × 10⁻¹ = 1 × 10⁻¹

CH₃COOH
$$
\rightleftharpoons
$$
 H⁺ + CH₃COO⁻ Label to
\n $C_2(1-\alpha_2)$ C₁ α_1 + C₂ α_2 C₂ α_2 Q₃ CH₃NH
\n[CH₃COO⁻] = C₂ α_2 =0.01 × 10⁻² = 1 × 10⁻¹ 10⁻¹ HCl and
\nconcent
\n[OH⁻¹]= $\frac{K_w}{\alpha_2}$ = $\frac{10^{-14}}{2}$ = 0.5 × 10⁻¹¹ = 5 × 10⁻¹² M (A) 1.6

$$
[OH^-] = \frac{100}{[H^+]^2} = \frac{100}{2 \times 10^{-3}} = 0.5 \times 10^{-11} = 5 \times 10^{-12} M
$$
\n
$$
[HOC] = 10^{-2} (1 - 0.01) = 9 \times 10^{-3} M
$$
\n
$$
[CH_3COOH] = 10^{-1} (1 - 0.01) \approx 10^{-1}
$$
\n
$$
= 10^{-1} (1 - 0.01) \approx 10^{-1}
$$

ISOHYDRIC SOLUTIONS

- **(i)** Solution of electrolytes are said to be isohydric if the concentration of the common ion present in them is the same and on mixing such solution, there occurs no change in the degree of dissociation of either of the electrolyte.
- (ii) Let the isohydric solution is made by HA_1 and HA_2 acids, then $[H^+]$ of both should be equal i.e.

$$
\sqrt{K_{a_1}C_1} = \sqrt{K_{a_2}C_2}
$$
 or $\frac{K_{a_1}}{K_{a_2}} = \frac{C_2}{C_1}$

TRY IT YOURSELF-1

Q.1 What molar concentration of ammonia will provide a hydroxyl ion concentration of 1.5×10^{-3} ?

 $(K_b = 1.8 \times 10^{-5})$ (A) 0.125 M (B) 0.12 M (C) $0.13 M$ (D) $0.14 M$

Q.2 The hydrolysis constant for $ZnCl_2$ will be –

(A)
$$
K_h = \frac{K_w}{K_b}
$$
 (B) $K_h = \frac{K_w^2}{K_b}$

(C)
$$
K_h = \frac{K_w^2}{K_{b_c}^2}
$$
 (D) $K_h = \frac{K_b}{K_w^2}$ (8) (B) (9) (A) (10)

 K_w^2 where, K_b is effective dissociation constant of base Zn^{++} .

- **STUDYMA**

DRE 16:

Calculate pH of solution obtained by mixing vol. of

DRE 16:

Calculate pH of solution obtained by mixing vol. of
 $K_{a_1}[\text{HOCI}] = 2 \times 10^{-4}$; $K_{a_2}[\text{CH}_3\text{COOH}] = 2 \times 10^{-5}$
 $K_{a_3}[\text{HOCI}] = 2 \times 10^{-4}$ 5 (C) 4 (D) 1 **STUDYMATERIAL: CHEMISTRY**

obtained by mixing vol. of
 R_{3} COOH solution given that
 R_{42} [CH₃COOH] = 2 × 10⁻⁵
 R_{42} [CH₃COOH] = 2 × 10⁻⁵
 R_{42} [CH₃COOH] = 2 × 10⁻⁵
 R_{42} (D₁) = 2 × 10⁻⁵
 R_{42 **STUDYMATER**

De 16:

Calculate pH of solution obtained by mixing vol. of

Calculate pH of solution obtained by mixing vol. of
 K_{a_1} [HOCI] = 2 × 10⁻⁴; K_{a_2} [CH₃COOH] = 2 × 10⁻⁵

Mass calculate OH⁻, OCI⁻, **EXERCISE THE MILTERTY CHEM**

STUDYMATERIAL: CHEM

6:

6.

MATORIAL: CHEM

6.

MATORIAL: CHEM

HOCULE 0.2 2 NOT-3; NASILONID given that

HOCULE 0.2 2 NOT-3; NASILONID given that
 $\begin{array}{ll}\n\text{HIOCI} & \text{D2} \times 10^{-4} \times \text{R}_{2} \text$ **EXERUIDENTIFIERIAL:** CHE

SIMENDIVITE CHE

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SIMIRATE CHE

MHOCI & 0.2 M CH₃COOH solution optimal by mixing vol. of

MHOCI & 0.2 M CH₃COOH solution given that

HOCI = 2 × 10⁻⁴; K_{a₂CH₃COOH = 2} **EXECUTE:**

EXECUTE:

Del de 0.2 M CH₃COOH solution given that
 $2 \times 10^{-4} \times 0.01 + 2 \times 10^{-5} \times 0.1$
 $\times 10^{-6} + 2 \times 10^{-6} = 2 \times 10^{-3}$
 \Rightarrow H of the resulting solution
 $\times 10^{-6} + 2 \times 10^{-6} = 2 \times 10^{-3}$
 $\times 10^{-4} \times 0.01$ **EXECUTE:**

Del of solution obtained by mixing vol. of

DCI & 0.2 M CH₃COOH solution given that
 $2I = 2 \times 10^{-4}$, $2CI$, $2CI$, $2CI$

and $2I = 10^{-1}$; $\alpha_2 = \frac{2 \times 10^{-5}}{2 \times 10^{-3}}$
 $\alpha_3 = 10^{-1}$; $\alpha_4 = \frac{2 \times 10^{-4} \times$ **STUDYMAT**

STUDYMAT

STUDYMAT

STUDYMAT

Correlation obtained by mixing vol. of

MHOCl & 0.2 M CH₃COOH solution given that

(A) The first and second dissociation constant of

(C) 4 (B)

calculate OH , OCl⁻, K_{a₂CH} $\frac{-5}{2}$ **Q.6** Acidity of BF₃ can be explained on the basis of which of $2 = \frac{10}{2 \times 10^{-3}} = 10^{-7}$ the following concepts? **STUDYMATERIAL: CHEMIS**

by mixing vol. of
 $(0.11 \times 10^{-5} \text{ m}) = 10^{-2}$
 2 10⁻⁵
 2 10⁻⁵
 2 10⁻⁵
 10 10 the resulting solution will be -
 $(0.1 \times 10^{-10} \text{ m}) = 10^{-2} \text{ m}$
 10 10 the first and second dissoci **STUDYMATERIAL: CHEMIS**

by mixing vol. of
 $(0.911) = 2 \times 10^{-5}$
 2.3 lec of 0.1 NHCl is added to 99 cc solution of NaCl.

solution given that
 (1.91 C)
 $(1.9$ **STUDYMATERIAL: CHEMISTRY**

indely mixing vol. of

OOH solution given that

(OA 1 and the resulting solution will be-

(CH₃COOH] = 2 x 10⁻⁵

(C) 4 (B)

(C) 4 (B)
 $\times 10^{-3}$ (A) Arrhenius concepts. 1= $\sqrt{k_A_1C_1 + k_B_2C_2}$

(C) 5.0 × 10⁻¹⁵

(C) 5.0 × 10⁻¹⁵

(C) 5.0 × 10⁻¹⁵

(C) 5.0 × 10⁻¹⁵

(C) 5.0 × 10⁻⁵

(C) 5.0 × C₁ + K_a - C₂ + C₂ $rac{2 \times 10^{-4} \times 0.01 + 2 \times 10^{-5} \times 0.1}{2 \times 10^{-3}} = 10^{-1}$; $\alpha_2 = \frac{2 \times 10^{-5}}{2 \times 10^{-3}} = 10^{-2}$
 $\alpha = 0.01 \times 10^{-1} = 1 \times 10^{-1}$
 $\alpha = 0.01 \times 10^{-1} = 1 \times 10^{-1}$
 $\alpha = 0.01 \times 10^{-1} = 1 \times 10^{-1}$
 $\alpha = 0.01 \times 10^{-1} = 1 \times 10^{-1}$
 Noting become double

N₁C₂ - 0.1 (A) 5.0 × 10⁻¹⁵

(A) 5.0 × 10⁻¹⁵

(D) 5.0 × 10⁻¹⁵

(D) 5.0 × 10⁻¹⁵

10⁻⁴ × 0.01+ 2 × 10⁻⁵

10⁻⁴ × 0.01+ 2 × 10⁻⁵

- log 2 = 3 - 0.3010 = 2.69

10⁻⁴ × 10⁻¹ is a co **Q.3** 1cc of 0.1 N HCl is added to 99 cc solution of NaCl. The pH of the resulting solution will be – (A) 7 (B) 3 **Q.4** The first and second dissociation constants of an acid H₂A are 1.0×10^{-5} and 5.0×10^{-10} respectively. The overall dissociation constant of the acid will be – (A) 5.0×10^{-5} (B) 5.0×10^{15} (C) 5.0×10^{-15} (D) 5.0×10^{5} **Q.5** The pH of neutral water at 25°C is 7.0. As the temperature increases, ionisation of water increases, however, the concentration of H+ ions and OH– ions are equal. What will be the pH of pure water at 60°C? (A) Equal to 7.0 (B) Greater than 7.0 (C) Less than 7.0 (D) Equal to zero (B) Bronsted Lowry concept (C) Lewis concept (D) Bronsted Lowry as well as Lewis concept. **Q.7** The ionisation of hydrochloric in water is given below: pure water at 60°C?

(B) Greater than 7.0

(D) Equal to zero

nn be explained on the basis of which of

necepts?

myr concept

oncept

wry as well as Lewis concept.

they docuble of the basis of which of
 $\frac{1}{2}$

wry a (C) Less than 7.0

(C) Less than 7.0

(D) Equal to zero

Acidity of BF₃ can be explained on the basis of which of

the following concept?

(B) Bronsted Lowry as well as Lewis concept.

(D) Bronsted Lowry as well as Lewi
	- $\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{Cl}^-(aq)$ Label two conjugate acid-base pairs in this ionisation.
	- $^{-12}$ M (A) 1.6×10^{-7}
(C) 3×10^{-5} **Q.8** CH₃NH₂ (0.1 mol, $K_b = 5 \times 10^{-4}$) is mixed with 0.08 mol of HCl and the solution is diluted to 1 litre. The H^+ ion concentration of the resulting solution will be – (A) 1.6×10^{-11} (B) 8×10^{-11} (D) 1.25×10^{-4}
		- **Q.9** If CO_2 be passed into water, what will remain present in solution –

$$
(A) CO_2, H_2CO_3, CO_3^{2-}, HCO_3^- (B) H_2CO_3, CO_2
$$

$$
1CO_3^-, CO_3^{2-}
$$
 (D) CO_2 , HCO_3^-

- CI $= C_1\alpha_1 = 0.01 \times 10^{-1} = 1 \times 10^{-3}$
 $C_2 (1 \alpha_2)$ $C_1\alpha_4 + C_2\alpha_2$ $C_2\alpha_2$
 $C_3(1 \alpha_2)$ $C_1\alpha_4 + C_2\alpha_2$ $C_2\alpha_2$
 $C_3(1 \alpha_2)$ $C_1\alpha_1 + C_2\alpha_2$ $C_2\alpha_2$
 $C_3(1 \alpha_3)$ $C_1\alpha_1 + C_2\alpha_2$ $C_2\alpha_2$
 $C_4(1 \alpha$ 0.5 × 10⁻¹¹ = 5 × 10⁻¹² M

(A) 1.6 × 10⁻¹¹ (B) 8 × 1

(C) 3 × 10⁻¹ (B) 8 × 1

(C) 1.25

(C) 3 × 10⁻⁵ (D) 1.25

solution –

(A) CO₂, H₂CO₃, CO₃⁻, HCO₃ (B) H

1 to be isohydric if the

(C) HCO₃, CO + CH₃COO⁻

+ CH₃COO⁻
 $C_2\alpha_2$ C₂ α_2
 $\approx 10^{-2} = 1 \times 10^{-1}$

(38 CH₃NH₂OU₁ mix in this ionisation
 $\approx 10^{-2} = 1 \times 10^{-1}$

(38 CH₃NH₂OU₁ mix ionisation is diluted to 1 litre. The H⁺

concentrati **Q.10** Given: $Ag^+ + NH_3 \rightleftharpoons Ag(NH_3)^+$; K₁ = 3.5 × 10⁻³ ; $\text{Ag(NH}_3)^+ + \text{NH}_3 \rightleftharpoons \text{[Ag(NH}_3)_2]^+$; K₂ = 1.7 × 10⁻³ The equilibrium constant for the reaction Ag^+ + 2NH₃ \rightleftharpoons $\text{[Ag(NH}_3)_2]$ ⁺ is $(A) 6.08 \times 10^{-6}$ (B) 6.08×10^{6} $(C) 6.08 \times 10^{-9}$ (D) 6.08×10^{9}
	- $\frac{K_{a_1}}{K} = \frac{C_2}{C}$ (C) 6.08 × 10⁻⁹ (D) 6.08 × 10⁻⁹ (D) 6.08 × 10⁻⁹ weak acid (HA, 1M) is $1/100$ th of that of a strong acid (HX, 1M), at 25°C. The K_a of HA is –

$$
\begin{array}{llll}\n\text{M,} & \text{M,} & \text{M,} \\
\text{M,} & \text{M,} & \text
$$

IONIC EQUILIBRIUM

CLASSIFICATION OF SALTS

Salt : A substance which ionises in water to produce ions other than H+ and OH– is called a salt. Salts are produced by the neutralisation reaction.

(a) Normal salt : In case the acid and base neutralise completely the salt formed is a normal salt. In case both acid and base are strong or weak, the salt is neutral, when the constituent acid is strong, the salt is acidic and when the constituent base is strong, the salt is alkaline.

 $\textbf{Example : NaCl}, \text{Na}_2\text{HPO}_3, \text{Na}_3\text{PO}_4, \text{NaH}_2\text{PO}_2$ and KNO_3

(b) Acidic salt : If a polybasic acid (e.g., H_2SO_4 , H_3PO_4 , H_2SO_3 etc.) is neutralised partly by a base, the salt formed is acidic.

Examples : NaHSO₃, NaHCO₃, NaH₂PO₄, Ca(H₂PO₄)₂ etc.

(c) **Basic salt :** If a polyacidic base $[e.g Ba(OH)₂, Bi(OH)₃,$ $Fe(OH)$ ₃ etc.] is neutralised partly by an acid, the salt formed is basic salt. $\pmb{\text{Example :} }\text{Mg(OH)Cl}, \text{Zn(OH)Cl}, \text{Al(OH)Cl}_2, \text{Pb(OH)NO}_3,$

 $Ba(OH)Cl, Bi(OH), Cl, Ca(OH)Cl, etc.$ **(d) Mixed salt :** Salts formed by the simultaneous

neutralisation of one acid by two bases or one base by two acids are called mixed salts.

 $\textbf{Example : } \text{NaKSO}_4$, ZnBrCl, Al(Cl₂)Br, Al(NO₃)Cl.Br, Ca(OCl)Cl etc.

(e) Double salt : A compound of two salts whose solution gives tests for all the constituent ions is called a double salt or lattice complex .

 $\textbf{Example}: \text{FeSO}_4.(\text{NH}_4)_2\text{SO}_4.6\text{H}_2\text{O},$ $[K_2SO_4.AI_2(SO_4)_3.24H_2O]$

(f) Complex salt : A compound whose solution does not give tests for the constituent ions is called a complex salt,

HYDROLYSIS OF SALT

1. **Salt of Weak acid and Strong base** [CH₃COONa, HCOONa] Let AB be salt i.e. HA is weak acid and BOH is strong base. $A^- + B^+ + H_2O \Longleftrightarrow HA + BOH$

After hydrolysis BOH will again get ionised to give B^+ and OH– \angle ions $BOH \Leftrightarrow B^+ + OH^-$

Thus only A⁻ ions are hydrolysed to give weak acid HA $\&$ OH— ions. This is known as Anionic Hydrolysis

 $A^{-} + H_2O \Leftrightarrow HA + OH^{-}$

$$
K_h = \frac{[AH][OH^-]}{[A^-]}
$$
(1)

⁼ [AH][OH] [A] Weak acid HA further ionised partially and set a reversible equation formally to attain equilibrium

 $HA \Leftrightarrow H^+ + A^-$

A⁻ + H₂O
$$
\Leftrightarrow
$$
 HA + OH⁻
\n $K_h = \frac{[AH][OH^-]}{[A^-]}$ (1)
\nWeak acid HA further ionised partially and set a reversible
\nequation formally to attain equilibrium
\n $HA \Leftrightarrow H^+ + A^-$
\n $K_a = \frac{[H^+][A^-]}{[HA]}$ (2)
\n $K_w = [H^+][OH^-]$ (3)
\nEq. (1), (2) and (3), $K_h = \frac{K_w}{K_a}$

As earlier, let C was the concentration (mole lit^{-1}) taken initially and 'h' be the degree of hydrolysis (mole ionised per mole of salt) then,

^A– + H2^O HA + OH– C 0 0 (initially) C(1 – h) Ch Ch (at equilibrium) K^h ⁼ ² [AH] [OH] Ch.Ch Ch [A] C (1 h) 1 h if h <<< 1 so 1 – h 1 ^K^h = Ch² h = K^h C h = h W ^a ^K ^K C K .C [OH–] = Ch = C W W a a K K .C K .C K pOH = ^a 1 1 7 pK log C 7 2 2 pH = 14 – pOH pH = ^a 1 1 7 pK log C 7 2 2 **Result : (a)** Solution will be basic ^a 1 1 1 1 pH 7 pK log C 7 7 log10 2 2 2 2 7 3 14 7 3 21 3 7 log10 9 2 2 2 2

(c) pH of solution is > 7 .

(d) Red litmus change in to blue litmus.

Example 17 :

Find out pH, h and [OH–] of milli molar solution of KCN if the dissociation constant of HCN is 10^{-7} .

Sol. (i) pH =
$$
7 + \frac{1}{2}pK_a + \frac{1}{2}\log C = 7 + \frac{1}{2} \times 7 + \frac{1}{2}\log 10^{-3}
$$

= $7 + \frac{7}{2} - \frac{3}{2}\log 10 = \frac{14 + 7 - 3}{2} = \frac{21 - 3}{2} = 9$

(ii)
$$
h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}} = 10^{-5}
$$

EXAMPLE 13.24 STUDY MATERIAL: CHEMSTRY
\n**(ii)**
$$
h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}} = 10^{-5}
$$

\n**(iii)** $[OH^-] = \sqrt{\frac{K_w \times C}{K_a}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}} = 10^{-5}$
\n**13.24 14.2 15.2 16.2 17.2 17.2 18.2 19.2 19.2 10.2 10.2 11.2 11.2 12.2 13.2 14.2 15.2 16.2 17.2 18.2 19.2 19.2 10.2 10.2 11.2 11.2 12.2 13.2 14.2 15.2 16.2 17.2 18.2 19.2 10.2 11.2 11.2 12.2 13.2 14.2 15.2 16.2 17.2 17.2 18.2 19.2 19.2 10.2 11.2 11.2 11.2 12.2 13.2 14.2 15.2 16.2 17.2 17.2 18.2 19.2 1**

2. Salt of a strong acid and a weak base :

 $\mathrm{NH}_4\mathrm{Cl}, \mathrm{NH}_4\mathrm{NO}_3, \mathrm{(NH_4)_2SO_4)}$

Let AB be salt, HA is strong acid and BOH is weak base. $A^- + B^+ + H_2O \Longleftrightarrow HA + BOH$

Only B^+ ion is hydrolysed so it is known as cationic hydrolysis.

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

$$
K_{h} = \frac{[BOH][H^{+}]}{[B^{+}]}
$$
(1)

Weak base BOH further ionised partially and set a reversible equation formally to attain equilibrium

 $BOH \Leftrightarrow B^+ + OH^-$

$$
K_h = \frac{[BOH][H^+]}{[B^+]}
$$

Weak base BOH further ionised partially and set a reversible equation formally to attain equilibrium [BOH]

 $BOH \Leftrightarrow B^+ + OH^-$

$$
K_b = \frac{[B^+][OH^-]}{[BOH]}
$$
 (2)

$$
K_w = [H^+][OH^-]
$$
 (3)

By eq. (1), (2) and (3), $K_h = \frac{K_w}{K}$ $K_e = \frac{|H|}{K}$ b_a

As earlier, let C was the concentration (mole lit^{-1}) taken initially and 'h' be the degree of hydrolysis (mole ionised per mole of salt) then,

$$
B^{+} + H_{2}O \Leftrightarrow BOH + H^{+}
$$

\nC 0 0 (initially)
\nC(1-h) Ch Ch (at equilibrium)

$$
K_{h} = \frac{[BOH][H^{+}]}{[B^{+}]} = \frac{Ch Ch}{C (1-h)} \quad ; \quad K_{h} = \frac{Ch^{2}}{1-h}
$$
 Let degree of hydroly

If $h \ll 1$ so $1 - h \approx 1$

B⁺ + H₂O
$$
\Leftrightarrow
$$
 BOH + H⁺
\nC 0 0 (initially)
\nC(1-h) Ch Ch (at equilibrium)
\nK_h = $\frac{[BOH][H^+]}{[B^+]} = \frac{ChCh}{C(1-h)}$; K_h = $\frac{Ch^2}{1-h}$
\nIf h << 1 so $1-h \approx 1$
\n $\Rightarrow K_h = Ch^2 \Rightarrow h = \sqrt{\frac{K_h}{C}}$
\n $h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_w}{K_b.C}}$
\n $[H^+] = Ch = C \sqrt{\frac{K_w}{K_b \times C}} = \sqrt{\frac{K_w \times C}{K_b}}$
\n $h = \sqrt{\frac{K_w}{K_b \times C}} = \sqrt{\frac{K_w \times C}{K_b}}$
\n $h = \sqrt{\frac{K_w}{K_a.K_b}}$
\n $h = \sqrt{\frac{K_w}{K_a.K_b}}$

STUDYMATERIAL: CHEMISTRY
\n
$$
\frac{-14 \times 10^{-3}}{10^{-7}} = 10^{-5}
$$
\n
$$
\frac{\text{REsult :}}{\text{REsult :}}
$$
\n
$$
\frac{\text{Result :}}{10^{-7}} = 10^{-5}
$$
\n
$$
\frac{\text{Result :}}{\text{(a) Hydrolysis is cataionic (b) Nature of solution is acidic.}}{\text{(c) P}^H \text{ of solution is less than 7.}} \text{Result :}
$$
\n
$$
\frac{\text{(d) Blue litmus change into red litmus.}}{\text{(e) P}^H \text{ of solution is less than 7.}}
$$

Result :

- **(a)** Hydrolysis is cataionic
-
- **(c)** PH of solution is less than 7.
- **(d)** Blue litmus change into red litmus.

Example 18 :

How many grams of $NH₄Cl$ should be dissolved per litre of solution to have a pH of 5.13 ? K_b for NH₃ is 1.8×10^{-5} .

Sol. NH₄Cl is a salt of strong acid and weak base for solutions of such salts.

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

10.26 = 14. $\log C$, 4.74

 \Rightarrow 10.26 = 14 – log C – 4.74 $\Rightarrow \log C = 9.26 - 10.26 = -1.0$ $\therefore C = 10^{-1}$ M $NH_ACl = 10⁻¹ M$

$$
W_{\text{NH}_4\text{Cl}} = 10^{-1} \times 53.5 \text{g L}^{-1} = 5.35 \text{ g L}^{-1}
$$

3. Salt of a weak acid and a weak base

N C V 1 10⁻¹

H ⁻ $\sqrt{K_{av}} \times C$ and $\sqrt{K_{ax}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}} = 10^{-5}$

(a) Hydrolysis is extationic (e) p¹¹ of solution is less than 7.

a strong acid and a weak base:

NH₄NO₃, (NH₄₂)-SO₄)

be sell, $\sqrt{\frac{K_w \times C}{K_a}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}} = 10^{-5}$

(a) Hydrolysis is cataionic

(b) Nature of solution is acidic.

(c) PH of solution is less than 7.

(d) Blue litmus change into red lit

(c) PH of solution is less than 7 ataionic

s less than 7.

s less than 7.

f NH₄Cl should be dissolved per litre

ipH of 5.13 ? K_b for NH₃ is 1.8×10^{-5} .

trong acid and weak base for solutions

log C – pK_b]
 $\{C-4.74$
 $0.26 = -1.0$ \therefore C = $\text{OH} = \left[7 - \frac{1}{2} \text{pK}_b - \frac{1}{2} \log C\right] < 7$

Ht:

Ht:

Hydrolysis is cataionic

Harture of solution is acidic.

H of solution is less than 7.

Sue litmus change into red litmus.

8:

Sue many grams of NH₄Cl should be d Let AB be salt i.e. HA and BOH both are weak acid and weak base respectively. Here both A^- and B^+ ions will get hydrolysed because neither of HA and BOH are completely ionisable. any grams of NH₄Cl should be dissolved per litre

ion to have a pH of 5.13 ? K_b for NH₃ is 1.8×10^{-5} .

is a salt of strong acid and weak base for solutions

salts.
 $= \frac{1}{2} [pK_w - log C - pK_b]$
 $26 = 14 - log C - 4.74$
 $C =$ grams of NH₄Cl should be dissolved per litre

to have a pH of 5.13 ? K_b for NH₃ is 1.8×10^{-5} .

salt of strong acid and weak base for solutions

ts.
 $\begin{bmatrix} [PK_w - log C - pK_b] \end{bmatrix}$
 $\begin{bmatrix} 14 - log C - 4.74 \end{bmatrix}$
 $= 10^{-1$ s alts.
 $I = \frac{1}{2} [pK_w - log C - pK_b]$
 $26 = 14 - log C - 4.74$
 $2C = 9.26 - 10.26 = -1.0$ ∴ $C = 10^{-1} M$
 $A_4 C1 = 10^{-1} M$
 $N H_4 C1 = 10^{-1} \times 53.5 g L^{-1} = 5.35 g L^{-1}$

a weak acid and a weak base

to be salt i.e. HA and BOH both are weak ac 14-log C-4.74

9.26-10.26=-1.0 ∴ C=10⁻¹ M
 $= 10^{-1}$ M
 $= 10^{-1} \times 53.5g$ L⁻¹ = 5.35 gL⁻¹
 k acid and a weak base

alt i.e. HA and BOH both are weak acid and

spectively. Here both A⁻ and B⁺ ions will get

ec 26 = 14 - log C - 4.74

c = 9.26 - 10.26 = -1.0 ∴ C = 10⁻¹ M

4₄Cl = 10⁻¹ M

NH₄Cl = 10⁻¹ × 53.5g L⁻¹ = 5.35 gL⁻¹
 a weak acid and a weak base

be salt i.e. HA and BOH both are weak acid and

as expective 1 = 10⁻¹ × 53.5g L⁻¹ = 5.35 gL⁻¹
 k acid and a weak base

alt i.e. HA and BOH both are weak acid and

spectively. Here both A⁻ and B⁺ ions will get

ecause neither of HA and BOH are completely
 $2O \Leftrightarrow HA + BOH$
 eak acid and a weak base
salt i.e. HA and BOH both are weak acid and
the respectively. Here both A⁻ and B⁺ ions will get
because neither of HA and BOH are completely
 $H_2O \Leftrightarrow HA + BOH$
 $H_3[[BOH]$
 $[A^-][B^+]$ (1)
 $H^+ +$

⁼ [B] [OH] ⁼ [BOH][H] [B] C (1 h) 1 h A– + B+ + H2^O HA + BOH K^h HAH+ + A– a b

$$
K_{w} = \frac{[H^{+}][A^{-}]}{[HA]}
$$
 (2)

$$
BOH = B^+ + OH^-
$$

$$
K_b = \frac{[B^+][OH^-]}{[BOH]}
$$
(3)

From equation (1), (2) $\&$ (3) we have

$$
K_{h} = \frac{K_{W}}{K_{a}.K_{b}}
$$

Let degree of hydrolysis be h

$$
A^{-} + B^{+} + H_{2}O \Leftrightarrow AH + BOH
$$

\n
$$
C \t C \t C \t 0 \t 0
$$

\n
$$
C(1-h) C(1-h) \t Ch \t 0
$$

\n
$$
C h
$$

\n
$$
h^{2}
$$

H⁻
\n
$$
A^{-} + B^{-} + H_2O \Leftrightarrow HA + BOH
$$

\n $\begin{array}{ccc}\n-1 & \text{.... (1)} & K_h = \frac{[HA][BOH]}{[A^-][B^+]} & \text{....(1)} \\
1 & \text{.... (3)} & HA \Leftrightarrow H^+ + A^-\n\end{array}$
\n $\begin{array}{ccc}\n\text{A. } & K_h = \frac{K_w}{K_b} & K_a = \frac{[H^+][A^-]}{[HA]} & \text{....(2)} \\
\text{B. } & K_h = \frac{[B^+][OH^-]}{[HA]} & \text{....(2)} \\
\text{A. } & BDH + H^+ & K_a = \frac{[B^+][OH^-]}{[BOH]} & \text{....(3)} \\
\text{B. } & \text{B. } \\
\text{B. } & \text{B. } \\
\text{B. } & \text{B. } \\
\text{B. } & \text{B. } \\
\text{B. } & \text{B. } & \text{B. } & \text{B. } & \text{B. } \\
\text{B. } & \text{B. } & \text{B. } & \text{B. } & \text{B. } \\
\text{B. } & \text{B. } & \text{B. } & \text{B. } & \text{B. } \\
\text{C. } & \text{C. } & \text{D. } & \text{A. } & \text{B. } \\
\text{D. } & \text{D. } & \text{D. } & \text{D. } & \text{D. } \\
\text{D. } & \text{D. } & \text{D. } & \text{D. } & \text{D. } \\
\text{E. } & \text{E. } \\
\text{E. } & \text{E. } \\
\text{E. } & \text{E. } & \text{E. } & \text{E$

IONIC EQUILIBRIUM

It is very important to note that here degree of hydrolysis 5. is independent of concentration of salt taken initially. It just depends upon K_a and K_b . **EXECUTE:**

The transit of once that here degree of hydrolysis 5. Salt of a weak p

of concentration of salt taken initially. It

(i) Salt contain

pon K_a and K_b .

in stepwis

in stepwis
 $A^{-} + H^{+}$
 H^{+}
 H^{+}

h increases on increasing temperature.

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

⁼ [A] [H] [H+] = K^a . h [H+] = K^a [H+] = 1/2 w a ^b K .K K 1 – ¹ pK^b

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

Conclusions :

Example 19 :

Calculate pH and degree of hydrolysis of 10^{-2} M NH₄CN solution. Given that K_a of HCN = 5×10^{-10} and K_b of (aq. NH₃) = 2×10^{-5} .

$$
[H^+] = \left(\frac{K_w K_a}{K_b}\right)^{1/2}
$$
\n
$$
pH = 7 + \frac{1}{2}pK_a - \frac{1}{2}pK_b
$$
\n**Conclusions:**\nConclusions:

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\nSubstituting the values of

4. Salt of strong acid & strong base

- **(a)** Here HA and BOH both are strong acid and strong base respectively.
- **(b)** As A– will get hydrolysed to HA, HA will again get ionised in to A⁻ ions. The same thing happens with B^+ ions so hydrolysis of both A^- and B^+ ions is not possible.
- (c) Also after coming into solution A^- and B^+ ions do not have tendency to react with each other they want to remain free so only A^- and B^+ are present in solution.

Result :

- **(a)** The solution of such salts remain neutral, i.e. $pH = 7$
- **(b)** It is not a process of hydrolysis.
- **(c)** No effect on litmus.

5. Salt of a weak polyprotic acid and strong base

here degree of hydrolysis 5. Salt of a weak polyprotic acid and strong base

is of salt taken initially. It

(i) Salt containing multivalent cation or anion :

Hydrolysis of multivalent cation or anion take

in stepwise m Example the hydrolysis 5. Salt of a weak polyprotic acid and strong base

to taken initially. It

(i) Salt containing multivalent cation or anion thydrolysis of multivalent cation or anion that

in stepwise manner and mor Example the properties are the properties and the properties of hydrolysis 5. Salt of a weak polyprotic acid and strong base

the taken initially. It (i) Salt containing multivalent cation or anion :

Hydrolysis of multiv degree of hydrolysis 5. Salt of a weak polyprotic acid and strong base

and taken initially. It

(i) Salt containing multivalent cation or anion :

Hydrolysis of multivalent cation or anion takes place

in stepwise manner **(i) Salt containing multivalent cation or anion :** Hydrolysis of multivalent cation or anion takes place in stepwise manner and more than one hydrolytic products are formed. For example the hydrolysis of Fe²⁺ ions will occur in following steps :
Fe²⁺ + 2H₂O \rightleftharpoons Fe (OH)⁺ + H₃O⁺ **SPRAMAGES**
 PRAMAGES
 SOLVANDE DENSITY OF START START SET ASSESSMENT AND SET AND REAL SET AND REAL SET AND REAL SET ASSESSMENT AND REAL SET ASSESSMENT AND THAT SO EXECUTE AND SET ASSESS AN ADVANCED LEARNING
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multivalent cation or anion :

multivalent cation or anion takes place

nanner and more than one hydrolysis of

occur in following steps :
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inner and more than one hydrolysis of

cur in following steps :

Fe (OH)⁺ + H₃O⁺
 **EXECUTE ANTIFY EXECUTE AN INTEGRAL SET AN ARTIST CONDUCT AN ANDELES AN INTEGRAL SET AND MOVED IN SURFACES AN ART AND SURFACES AND SURFACES AND AN INCREDUCT IN THE UP (Fe(OH)⁺ H₃O⁺
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FOR THE SURFALL OF EXAMPLE THEOREM INTERED INTERED IN Example 12**
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is of multivalent cation or anion takes place

is e manner and more than one**

Fe²⁺ + 2H₂O \rightleftharpoons Fe (OH)⁺ + H₃O⁺
Fe(OH)⁺ + 2H₂O \rightleftharpoons Fe(OH)₂ + H₃O⁺

$$
K_{h_1} = \frac{[Fe(OH)^+] [H_3O^+]}{[Fe^{2+}]} = \frac{K_w}{K_{b_2}}
$$

$$
K_{h_2} = \frac{[Fe(OH)_2][H_3O^+]}{[Fe(OH)^+]} = \frac{K_w}{K_{b_1}}
$$

Fe(OH)₂. Since $K_{b_1} \gg >> K_{b_2}$ thus $K_{h_1} \gg >> K_{h_2}$ hence the hydrolysis of $Fe(OH)^+$ in comparision can be neglected and

$$
[H_3O^+] = \sqrt{K_{h_1}C} \text{ or } pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}\log C
$$

e hydrolysis constants K_{h_1} and K_{h_2} are
 $N_{h_1} = \frac{[Fe(OH)^+] [H_3O^+]}{[Fe^{2+}]} = \frac{K_w}{K_{b_2}}$
 $N_{h_2} = \frac{[Fe(OH)_2][H_3O^+]}{[Fe(OH)^+]} = \frac{K_w}{K_{b_1}}$

ere K_{b_1} and K_{b_2} are dissociation constants of
 $(OH)_2$. Since K If selected and Intertainm and interting

Frequence I and Intertainment and Intertainment (I-

Fe²⁺ + 2H₂O \rightleftharpoons Fe(OH)⁺ + H₃O⁺

Fe(OH)⁺ + 2H₂O \rightleftharpoons Fe(OH)₂ + H₃O⁺

The hydrolysis constants K_{h d more than one hydrolytic

or example the hydrolysis of

ollowing steps :
 $y^+ + H_3O^+$

ts K_{h_1} and K_{h_2} are

ts K_{h_1} and K_{h_2} are
 $\frac{y^+1}{K_{b_2}} = \frac{K_w}{K_{b_1}}$

e dissociation constants of
 $\gg K_{b_2}$ ation or anion takes place

ore than one hydrolytic

xample the hydrolysis of

wing steps:
 H_3O^+
 $L_1H_3O^+$
 K_{h_1} and K_{h_2} are
 $= \frac{K_w}{K_{b_2}}$
 $= \frac{K_w}{K_{b_1}}$

issociation constants of
 K_{b_2} thus $K_{h_$ **(ii) Salt containing weak conjugate cation and an** $\mathbf a$ **mphiprotic cation:** When a salt like NaHCO₃, NaHS, NaH_2PO_4 , Na_2HPO_4 etc. is dissolved in water the amphiprotic anion can either accept a proton from water (Hydrolysis) or can donate a proton to water. The pH of the solution can be calculated as and K_{b_2} are dissociation constants of

nce $K_{b_1} \gg\gg\gg K_{b_2}$ thus $K_{h_1} \gg\gg K_{h_2}$

drolysis of Fe(OH)⁺ in comparision can

and
 $K_{h_1}C$ or $pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}log C$

ning weak conjugate cation and an
 Fre \int $K_{h_2} = \frac{[Fe(OH)_2][H_3O^+]}{[Fe(OH)^+]} = \frac{K_w}{K_{h_1}}$

where K_{h_1} and K_{h_2} are dissociation constants of

Fe(OH)₂. Since $K_{h_1} \gg \gg \gg K_{h_2}$ thus $K_{h_1} \gg \gg K_{h_2}$

nence the hydrolysis of Fe(OH)⁺ in comp $=\frac{[Fe(OH)^3][H_3O^3]}{[Fe^2^+]} = \frac{K_w}{K_{b_2}}$
 $=\frac{[Fe(OH)_2][H_3O^4]}{[Fe(OH)^4]} = \frac{K_w}{K_{b_1}}$

re K_{b_1} and K_{b_2} are dissociation constants of

DH₂. Since $K_{b_1} \gg \gg \gg K_{b_2}$ thus $K_{h_1} \gg \gg K_{h_2}$

ee the hydrolysis of F $[H_3O^+] = \sqrt{K_{h_1}C}$ or $pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}\log C$
 Salt containing weak conjugate cation and an amphiprotic cation: When a salt like NaHCO₃, NaHS, NaH₂PO₄, Na₂HPO₄ etc. is dissolved in water the amphipro 3^O $J = \sqrt{K_{h_1}C}$ or $pH = \sqrt{2}pk_{b_2} - \frac{1}{2}logC$

It containing weak conjugate cation and an

phiprotic cation: When a salt like NaHCO₃, NaHS,
 H_2PO_4 , Na₂HPO₄ etc. is dissolved in water the

phiprotic anion can nce $K_{b_1} \gg \gg \gg K_{b_2}$ thus $K_{h_1} \gg \gg K_{h_2}$
drolysis of Fe(OH)⁺ in comparision can
and
and
 $\overline{K_{h_1}C}$ or $pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}log C$
ing weak conjugate cation and an
cation: When a salt like NaHCO₃, NaH re(Or1)₂. Since $\kappa_{b_1} \gg \kappa_{b_2}$ thus $\kappa_{h_1} \gg \kappa_{h_2}$
hence the hydrolysis of Fe(OH)⁺ in comparision can
be neglected and
 $[H_3O^+] = \sqrt{K_{h_1}C}$ or $pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}log C$
Salt containing weak conjugat B_1
 B_2 are anoted in the hydrolysis of Fe(OH)⁺ in comparision can

decided and
 $B_1 = \sqrt{K_{h_1}C}$ or $pH = 7 - \frac{1}{2}pk_{b_2} - \frac{1}{2}logC$

containing weak conjugate cation and an
 $B_1 = \sqrt{K_{h_1}C}$ or $pH = 7 - \frac{1}{2}pk_{b$

$$
pH = \frac{1}{2} (pK_1 + pK_{a_2})
$$

 amphiprotic cation : (iii) Salt containing strong conjugate cation and an

Examples of such salts are NH_4HCO_3 , NH_4HS etc.

$$
[\mathbf{H}_3 \mathbf{O}^+] = \sqrt{\frac{\mathbf{K}_{a_1} \mathbf{K}_{w}}{\mathbf{K}_{b}}} - \mathbf{K}_{a_1} \mathbf{K}_{a_2}
$$

(iv) Hydrolysis of amphiprotic anion :

Let us consider hydrolysis of amphiprotic anion only, i.e., when counter cation is not hydrolysed, example of some salts of this category are NaHCO₃, NaHS, Na_2HPO_4 , NaH_2PO_4 . r can donate a proton to water. The pH

can be calculated as
 $\cdot pK_{a_2}$)
 ig strong conjugate cation and an
 ition:

uch salts are NH₄HCO₃, NH₄HS etc.

the solution can be calculated as
 $\frac{a_1 K_w}{K_b} - K_{a_1} K_{$ + pK_{a2})

ing strong conjugate cation and an

cation:

such salts are NH₄HCO₃, NH₄HS etc.

of the solution can be calculated as
 $\frac{K_{a_1}K_w}{K_b} - K_{a_1}K_{a_2}$

imphiprotic anion :

er hydrolysis of amphiprotic ani

Here, $H_2PO_4^-$ and HPO_4^2 are amphiprotic anions. pH after their hydrolysis can be calculated as,

pH of H₂PO₄⁻ in aqueous medium =
$$
\frac{pK_{a_1} + pK_{a_2}}{2}
$$

pH of H₂PO₂⁻⁴ in aqueous medium = $\frac{2}{2}$ $+ pK_{32}$

Here, $H_2PO_4^2$ is conjugate base of $H_2PO_4^-$ and H_3PO_4 is conjugate acid of $H_2PO_4^-$.

Similarly, PO^{3-4} isconjugate base of HPO^{-2} ₄ and HPO^{-}_{4} is conjugate acid of PO^{3-}_{4} . When these salts are dissolved in water, $[H_3O^+]$ concentration can be determined as,

$$
[H_3O^+] = \sqrt{K_{a_1}(K_w / k_b + K_{a_2})}
$$

$$
pH = -\log \sqrt{K_{a_1}(K_w / k_b + K_{a_2})}
$$

BUFFER SOLUTION

A solution that resists change in pH upon addition of small amount of strong acid and strong base is called Buffer Solution. Capacity of a solution to resist alteration in its pH value is known as buffer capacity.

Salient features of buffer solutions

- **(i)** Its pH does not change on standing for long.
- **(ii)** Its pH is slightly changed by the addition of small amount of acid & base.
- **(iii)** It has definite pH.
- **(iv)** It has reverse acidity or alkalinity.

Types of buffer solution

(i) Simple buffers :

- **(a)** A salt of weak acid & weak base in water example CH₃COONH₄, NH₄CN
- **(b)** Proteins & Amino acids.
- **(c)** A mixture of an acid salt & normal salt of a poly basic acid example $\text{Na}_2\text{HPO}_4 + \text{Na}_3\text{PO}_4$

(ii) Mixed buffers : These are of two types

(a) Acidic buffer mixtures : Solution of a weak acid and its salt with strong base is acid buffer. e.g. : NaHCO₃ + H₂CO₃, CH₃COOH + CH₃COONa, Boric acid (H_3BO_3) + Borax (Na₂B₄O₇) **Acidic buffer action :** It is shown by following equilibrium.

 $\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+ \text{(partly ionised)}$ (2) $CH_3^5COONa \longrightarrow CH_3COO^- + Na^+$ (fully ionised) $(3) H₂O \rightleftharpoons H⁺ + OH⁻$

pH of an acidic buffer :

Let in acidic buffer CH₃COOH + CH₃COONa
(1)
$$
CH_{2}COOH \rightarrow CH_{2}COCH + H^{+}
$$

$$
(1) CH3COOH \rightleftharpoons CH3COO- + H+
$$

$$
Ka = \frac{[CH3COO-]}{HaCOO-}
$$

$$
K_a = \frac{\text{[CH}_3\text{COO}^\dagger\text{[H}^+]}{\text{[CH}_3\text{COOH}]}
$$

- **(2)** In the above equation we will have to assume two (vi) things things
- Since $CH₃COOH$ is a weak acid it's ionisation may be neglected so at equilibrium, it's concentration may be assumed the same as it was before ionisation
- * CH3COO– ions concentration may be assumed totally due to $CH₃COONa$ because it is highly ionised as compared to $CH₃COOH$ So, $[CH_3COO^-] = [CH_3COONa] = [Salt]$

ED LEALARNING	STUDY MATERIAL: CHEMISTRY
Similarly, PO ³⁻ ₄ is conjugate base of HPO ⁻² ₄ and HPO ⁻¹ is conjugate acid of PO ³⁻ ₄ .	$When these salts are dissolved in water, [H3O+] = K_a = \frac{[Salt] [H^{\dagger}]}{[Acid]}; [H^{\dagger}] = \frac{K_a[Acid]}{[Salt]}\n$
concentration can be determined as, [H ₃ O ⁺] = $\sqrt{K_{a_1}(K_w / k_b + K_{a_2})}$ pH = $-\log \sqrt{K_{a_1}(K_w / k_b + K_{a_2})}$ pH = $-\log \sqrt{K_{a_1}(K_w / k_b + K_{a_2})}$ pH = $-\log \sqrt{K_{a_1}(K_w / k_b + K_{a_2})}$ 	

These equation are called **Handerson equations.**

(b) Basic buffer mixture : A weak base with it's salt with strong acid.

e.g. : $NH₄OH + NH₄Cl$, glycine + Glycine Hydrochloride

Basic buffer action :

The best known example is a mixture of ammonium hydroxide and ammonium chloride.

 $NH_4OH \rightleftharpoons NH_4^+ + OH^-$ (weakly ionised)

 $NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$ (Highly ionised)

The NH $_4^+$ ions from completely ionised NH₄Cl suppress the ionisation of $NH₄OH$ which is already a weak ionisable substance. Thus the concentration of OH– ions remains constant.

When a few drops of a base (say NaOH) are added, the OH⁻ ions (from NaOH) combine with NH_4^+ ions to form nearly unionised NH₄OH $\&$ thus the concentration of OH– ions and hence PH value remains constant.

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

(from buffer) (from added base)

On the other hand, when a few drops of acid (say HCI) are added, the H^+ ions (from acid) combine with the excess of NH₄OH to form H_2O ammonium ions.

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

 (from buffer) (from added acid) Thus the addition acid does not increase H^+ ion concentration and hence again pH remains unchanged. pOH and pH values of basic buffers can be determined from the following Henderson's

equation.
$$
pOH = pK_b + log \frac{[Salt]}{[Base]}
$$

Application of buffer solutions

- **(i)** Qualitative analysis of mixture
- **(ii)** Quantitative analysis or estimations
- **(iii)** Industrial process such as manufacture of paper, dyes, inks, paint, drugs etc.
- **(iv)** Digestion of food
- Preservation of foods & Fruits
- Agriculture & Dairy products preservation.

Buffer capacity : May be defined as change in concentration of buffer acid required for changing its p^H value by one unit.

 $B.C. = \frac{\text{of acid or base added to litre of Buffer}}{\text{of buffer}}$ Change in concentration or number of moles

Change in pH

Example 20 :

Calculate the pH of a buffer prepared by mixing 300 cc of 0.3M NH₃ & 500 cc of 0.5 NH₄Cl K_b for NH₃ = 1.8×10^{-5} $(A) 8.11$ (B) 9.81 $(C) 8.82$ (D) None of these

- **Sol. (C).** Total volume of the buffer solution
-

 $= 300 \text{ cc} + 500 \text{ cc} = 800 \text{ cc}$ Q.6 Number of milli moles of NH₃ = $300 \times 0.3 = 90$ Molarity of NH₃ (Base) in the buffer = $90/800$ M Number of milli moles of $NH_4Cl = 500 \times 0.5 = 250.0$ Molarity of [salt] in the buffer = 250/800 M Henderson's equation for basic buffer is:

$$
pOH = -\log K_b + \log \frac{[Salt]}{[Base]}
$$

= -\log 1.8 \times 10^{-5} + \log \frac{250/800}{90/800}
= 4.74 + \log (250/90) = 4.74 + 0.44 = 5.18
pH = 14 - pOH = 14 - 5.18 = 8.82

Example 21 :

In which case pH will not change on dilution (A) 0.01 M CH₃COONa + 0.01 M CH₃COOH buffer $(B) 0.01$ M CH₃COONH₄ (C) 0.01 M NaH_2PO_4 (D) in all cases

Sol. (D). Mixture of Sodium acetate and acetic acid is a buffer of pH value equal to pKa so its buffer capacity is maximum and hence its pH will not change significantly while CH_3COONH_4 is a salt of weak acid CH_3COOH Solu and weak base NH_4OH whose magnitude of K_a and K_b are equal. So its pH does not depend upon concentration. Further more, NaH_2PO_4 is, in fact, a single solute buffer.

TRY IT YOURSELF-2

Q.1 The pH of a solution prepared by mixing 50.0 mL of $0.02M$ NaOH and 50mL of 0.04 M CH₃COOH is

 $(K_a = 1.8 \times 10^{-5})$

- **Q.2** The pK_b of NH₄OH is 4.75. An aqueous solution of $NH₄OH$ is titrated with HCl. The pH of the solution at a point when half of $NH₄OH$ is neutralised is –
- **Q.3** Which of the following will produce a buffer solution $*$ when mixed in equal volumes? (A) 0.1 mol dm⁻³ NH₄OH and 0.1 mol dm⁻³ HCl (B) 0.05 mol dm⁻³ NH₄OH and 0.1 mol dm⁻³ HCl
	- (C) 0.1 mol dm⁻³ NH₄OH and 0.05 mol dm⁻³ HCl
	- (D) 0.1 mol dm⁻³ CH₄COONa and 0.1 mol dm⁻³ NaOH $*$
- **Q.4** HX is weak acid $(K_a = 10^{-5})$. It forms a salt NaX (0.1 M) on reacting with caustic soda. The degree of hydrolysis of NaX is – $(A) 0.01\%$ (B) 0.0001%

EXERIBIRENT CONSUMED SET ALL CONSUMED SET ALL CONSUMERS

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90.500 MM (B) HCl < NH₄Cl < NGC + NH₄Cl < NG **Q.5** The pH of 0.1 M solution of the following salts increases in the order of : (A) NaCl \leq NH₄Cl \leq NaCN \leq HCl (B) HCl \leq NH₄Cl \leq NaCl \leq NaCN (C) NaCN < NH4Cl < NaCl < HCl (D) HCl<NaCl< NaCN < NH4Cl What will be the pH of the buffer solution containing 0.15 moles of NH_4OH and 0.25 moles of NH_4Cl . K_b for NH₄OH is 1.8×10^{-5} . $(A) 9.08$ (B) 10.03 $(C) 9.05$ (D) 9.03 **ANSWERS (1)** (A) **(2)** 9.25 **(3)** (C) **(4)** (A) **(5)**(B) $(6)(D)$ HCI<NH₄CI<N(aCI<NCI<NCI<NCI<NCI

NaCN<NH₄CI<NCI<NCI

NECI<NACI<NCI<NCI

NECI<NACI<NCI<NCI

HCI<NACI<NACI<NCI

and will be the pH of the buffer solution containing

5moles of NH₄OH and 0.25 moles of NH₄CI.

for NH NH₄Cl

NH₄Cl

MH₄Cl

medient solution containing

d 0.25 moles of NH₄Cl.

0⁻⁵.

(B) 10.03

(D) 9.03

(D) 9.03

225

(3) (C)

(B) (6) (D)

(PRODUCT

temperature the mass of a solute

in its saturated solution is pH of the burter solution containing

DH and 0.25 moles of NH₄Cl.

.8 × 10⁻⁵.

(B) 10.03

(D) 9.03
 (S) (B) (6) (D)
 ILITY PRODUCT

stant temperature the mass of a solute

blyent in its saturated solution is calle 1.8 x 10⁻⁵.

(B) 10.03

(D) 9.03

(D) 9.03
 ANSWERS

(2) 9.25

(3) (C)

(5) (B) (6) (D)

BILITY PRODUCT

most temperature the mass of a solute

solvent in its saturated solution is called

estion, solubility of a subs

SOLUBILITY & SOLUBILITY PRODUCT

Solubility : At a constant temperature the mass of a solute present in 100g of solvent in its saturated solution is called its Solubility.

In numerical question, solubility of a substance is expressed in mole lit^{-1} it is number of moles of solute present in one litre of solvent.

Number of moles =
$$
\frac{\text{amount in grams}}{\text{mol wt.in grams}}
$$

Solubility =
$$
\frac{\text{Number of moles}}{\text{Litres of the solution}}
$$

Solubility Product : Solubility product of a sparingly soluble electrolyte is equal to the product of ionic concentrations in a saturated solution at a given temperature.

Ionisation of compound A_xB_y may be shown as

$$
A_x B_y \Leftrightarrow xA^{y+} + yB^{x-}
$$

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

Let solubility of compound A_xB_y is S mol lit⁻¹ equilibrium

$$
A_xB_y \Leftrightarrow xA^{y+} + yB^{x-}
$$

\nS xS yS
\n
$$
K_{sp} = (xS)^x (yS)^y
$$

\n
$$
K_{sp} = x^x y^y S^{x+y}
$$

- * In a saturated solution $K_{sp} = [A^+] [B^-]$
- ^{*} In an unsaturated solution of AB, $K_{sp} > [A^+] [B^-]$ i.e. more solute can be dissolved.
- * In a super saturated solution $K_{sp} < [A^+] [B^-]$ i.e precipitation will start to occur.
- * For AB type (in case of AgCl, PbSO₄, BaSO₄); $K_{sp} = S^2$
- * For AB_2 or A_2B type (in case of PbCl₂, CaCl₂, Na₂SO₄, $Pb(NO_3)_2) K_{sp} = 4S^3$
- * For A₃B or AB₃ type [AlCl₃, FeCl₃, Na₃PO₄)]; K_{sp} = 27S⁴
- * For A₃B₂ or A₂B₃ type $[Al_2(SO_4)_3, Ba_3(PO_4)_2]$; K_{sp} = 108S⁵ **Note :** K_{sn} increases with increase in temperature.

Example 22 :

Given the solubility product of F_3B_2 is 2×10^{-30} . What will be the solubility in moles/litre.

EXAMPLE 22 :	STUDY1				
aple 22 :	(3) Predicting precipitation in io Given the solubility product of F_3B_2 is 2×10^{-30} . What will be the solubility in moles/litre.	(4) Salting out action of soap :			
(A) 1.85×10^{-32}) ^{1/5}	(B) $\left(\frac{2 \times 10^{-30}}{108}\right)^{1/5}$	(C) $\left(\frac{2 \times 10^{-28}}{5400}\right)^{1/5}$	(D) All of these (E) $\left(\frac{5}{5400}\right)^{1/5}$	(D) All of these in the solution of a P. Then	(A) Salting out action of soap :
(B) $K_{\text{sp}} \circ f A_3B_2 = 2 \times 10^{-30}$	(D) All of these in the following	(E) In qualitative analysis of mixture of solubility product. Some in follows:			

(C)
$$
\left(\frac{2 \times 10^{-28}}{5400}\right)^{1/5}
$$
 (D) All of these

101FARNING

1.1 TO DEFARNING

2.1 The solubility product of F₃B₂ is 2 × 10⁻³⁰. What

the solubility in moles/litre.

8.5 × 10⁻³²)^{1/5}

8.5 × 10⁻³²)^{1/5}

8.5 × 10⁻³²)^{1/5}

8.5 × 10⁻³²)^{1/5}

9.10 multati **STUDY!**

STUDY!

STUDY!

STUDY!

STUDY!

STEDEN!

22:

the solubility product of F₃B₂ is 2 × 10⁻³⁰. What

the mediation in an ionic race

the solubility in moles/litre.

1.85 × 10⁻³²)^{1/5}

(B) $\left(\frac{2 \times 10^{-30}}{1$ **Sol. (B).** K_{sp} of A₃B₂ = 2 × 10⁻³⁰. $A_3^9B_2 \rightleftharpoons 3A^{+2} + 2B^{-3}$ Assume s is the solution of A_3B_2 then $K_{\text{sp}} = (3s)^3 (2s)^2 = 108 s^5$

$$
s = 5\sqrt{\frac{K_{sp}}{108}} = \left(\frac{2 \times 10^{-30}}{108}\right)^{1/5}
$$

Example 23 :

If the solubility product of BaSO₄ is 1.44×10^{-12} then find out the solubility of SO_4^{-2} . .

Sol. $BaSO_4 \rightleftharpoons Ba^{+2} + SO_4^{-2}$

$$
S = \sqrt{K_{sp}} = \sqrt{1.44 \times 10^{-12}} = 1.2 \times 10^{-6}
$$

Effect of Common ion on Solubility:

For a saturated solution product of concentration of ions should not exceed a constant called solubility product. Now suppose, extra common ions are added to increase concentration of ions it will tend to increase the value of $K_{\rm sn}$, but it can not increase so ions will react to form molecule and get precipitated e.g. let solubility of AB in water be s moles L^{-1} . It is dissolved in solution CB having a common ion B– let concentration of CB be C.

$$
AB \Leftrightarrow A^{+} + B^{-}
$$

\n
$$
S \t S \t S
$$

\n
$$
K_{\text{sp}} = s^{2} = [A^{+}] [B^{-}] \t S \t(1)
$$

(Product should not exceed $s²$ otherwise precipitation will start) Now. Let solubility of AB in CB be S_1 .

$$
CB \Leftrightarrow C^+ + BC
$$

\n
$$
C
$$

\n
$$
AB \Leftrightarrow A^+ + BC
$$

\n
$$
S_1 S_1 + S_1 S_1
$$

\n
$$
[A^+] = s_1; [B^-] = s_1 + C; K_{sp} = (s_1)(s_1 + C)
$$

\nIf $s_1 \ll C, K_{sp} = s_1C = s^2$, from eq. (1)
\nor $s_1 = s^2/C$. So, solubility gets reduced.

APPLICATION OF Ksp

(1) In purification of common salt :

In a saturated solution of NaCl $&$ imparities, by passing HCl gas through it increase the Cl⁻ ion concentration which shifts the equilibrium NaCl \Leftrightarrow Na⁺ + Cl⁻ to left & causes the precipitation of NaCl

(2) In preparation of NaHCO³ by solvay method : Precipitation of NaHCO₃ from it's saturated solution is done by addition of NH_4HCO_3 , HCO_3 as common ion.

(3) Predicting precipitation in ionic reaction :

Precipitation in an ionic reactions could be predicted by comparing K_{sp} to the ionic concentration product of ions

(4) Salting out action of soap :

STUDYMATERIAL: CHEMISTI

2 is 2×10^{-30} . What

² 10⁻³⁰. What

² 10⁻³⁰. What

² 10⁻³⁰. What

² 210⁻³⁰. What

² 210⁻³⁰ ^{1/5}

^{1/5}
 4 Salting out action of soap :
 4 Salting out action of so $\overline{108}$ (RCOONa) concentration of Na⁺ increases & causes to **STUDYMATERIAL: CHEMISTRY**

B₂ is 2×10^{-30} . What

Predicting precipitation in ionic reaction :

Precipitation in an ionic reactions could be predicted by

comparing K_{sp} to the ionic concentration product of ions
 STUDYMATERIAL: CHEMISTRY

B₂ is 2×10^{-30} . What

(3) Predicting precipitation in ionic reaction :

Precipitation in an ionic reactions could be predicted by

comparing $K_{\rm sp}$ to the ionic concentration product of When NaCl is added to saturated solution of soap precipitation of soap it is due to $[Na^+][RCOO^-] > K_{sp}$

(5) In qualitative analysis :

Qualitative analysis of mixtures is based on the principle of solubility product. Some important applications are as follows:

Precipitation of II group radicals : Group reagent is H_2S . In presence dilute HCl the ionic product of H_2S is suppressed due to common ion $(H⁺)$ and only ionic product of the sulphides of II radicals exceeds their solubility product and get precipitated.

––

II and IV precipitated H₂S \rightleftharpoons 2H⁺ \downarrow + s⁻² \downarrow common ion effect.

Precipitation of IV group radicals : Group reagent is H_2S in presence of $NH₄OH$. The presence of $NH₄OH$ enhances the dissociation of H_2S . Thus due to high ionisation of H_2S the concentrations of S^{-2} ions increases and ionic product of sulphides of IV group radicals exceeds corresponding solubility product and get precipitated.

Precipitation of III group radicals : Group reagent is $NH₄OH$ in presence of NH₄Cl the presence of NH₄Cl suppresses the ionisation of NH_4OH , NH_4^+ as common ion. Thus the ionic product of hydroxides of III group radicals exceeds corresponding solubility products and get precipitated.

Example 24 :

You are given 10^{-5} M NaCl solution and 10^{-8} M AgNO₃ solution, they are mixed in 1 : 1 volume ratio, predict whether AgCl will be precipitated or not, if solubility product of AgCl is 10^{-10} . **EVALUATE:**
 IMD
 EXERCUTE: $\frac{10^{-5} \text{ N NaCl solution and } 10^{-8} \text{ M AgNO}_3}{\text{EVALU}}$

The precipitated or not, if solubility

is 10⁻¹⁰.

In the precipitated or not, if solubility
 $K_{\text{In}} = \frac{[H^+][\text{In}^-]}{[\text{HIn}]}$; $[H^+] = \frac{K_{\text{In}} \times [HIn]}{[\text{In}^-$ **M**

² M NaCl solution and 10^{-8} M AgNO₃

² mixed in 1 : 1 volume ratio, predict

² l be precipitated or not, if solubility

² x $\frac{10^{-10}}{2}$.

² $\frac{10^{-8}}{2} = 25 \times 10^{-15} < K_{sp}$

² $\frac{10^{-8}}{2} = 25 \times 10^{-$ **EXECUTE AT A SUBDAMAGE SOLUTION ANCIENT SURVEY AND BY APPLYING EDUCATION BY APPLYING EDUCATION BY APPLYING EDUCATION BY APPLYING EDUCATION SECTION AND RESPONSIVE SURVEY AND RESPONSIVE SURVEY AND RESPONSIVED AT A SUBDAMAG**

Sol. Ionic product =
$$
\frac{10^{-5}}{2} \times \frac{10^{-8}}{2} = 25 \times 10^{-15} < K_{sp}
$$
 pH

Hence, no precipitation will take place.

Example 25 :

What $[H^+]$ must be maintained in a saturated H_2S (0.1 M) to precipitate CdS but not ZnS, if $[Cd^{2+}] = [Zn^{2+}] = 0.1$ initially?

$$
K_{\rm sp} \text{(CdS)} = 8 \times 10^{-27}, K_{\rm sp} \text{(ZnS)} = 1 \times 10^{-21},
$$

$$
K_a \text{(H}_2\text{S)} = 1.1 \times 10^{-21}
$$

Sol. In order to prevent precipitation of ZnS,
\n
$$
[Zn^{2+}][S^{2-}] < K_{sp} (ZnS) = 1 \times 10^{-21}
$$

\n(Ionic product)

or
$$
(0.1)
$$
 [S²⁻] < 1 × 10⁻²¹
[S²⁻] < 1 × 10⁻²⁰ M

This is the maximum value of $[S^2]$ before ZnS will precipitate. Let $[H^+]$ to maintain this $[S^2^-]$ be x. Thus for $H_2S \rightleftharpoons 2H^+ + S^{2-}$

$$
K_a = \frac{[H^+] [S^{2-}]}{[H_2S]} = \frac{x^2 (1 \times 10^{-20})}{0.1} = 1.1 \times 10^{-21}
$$

or $x = [H^+] = 0.105 M$

 \therefore No ZnS will precipitate at a concentration of H⁺ greater than 0.105 M.

ACID–BASE INDICATOR

Indicators are organic compound of complex structure that change colour in solution as the pH changes.

Hence acid-base indicator is useful in deciding end point in titration. Indicator does not change colour at particular pH but it changes for a specific pH range.

Suitable indicator for pH titration :

Colour change interval (sensitive range or useful range): Generally indicator are weak organic acid or weak organic base.

Consider the dissociation of the indicator HIn [Weak organic acid] $\text{HIn} \rightleftharpoons \text{H}^+ + \text{In}^-$

HIn – (different colour)

In– – (different colour)

By applying law of mass action

applying law of mass action
\n
$$
K_{In} = \frac{[H^+][In^-]}{[HIn]} \; ; \; [H^+] = \frac{K_{In} \times [HIn]}{[In^-]}
$$
\n
$$
n = \text{dissociation constant of indicator}
$$
\n
$$
pH = pK_{In} + \log \frac{[In^-]}{[HIn]}
$$
\n
$$
lution assumes colour of HIn, when
$$
\n
$$
\frac{[HIn]}{[In^-]} = 10, pH = pK_{In} - 1
$$

 K_{In} = dissociation constant of indicator

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

Solution assumes colour of HIn, when

$$
\frac{\text{[HIn]}}{\text{[In}^-]} = 10, \text{pH} = \text{pK}_{\text{In}} - 1
$$

Solution assumes colour of In⁻, when

2 2 20 [H][S] x (1 10) K 1.1 10 [H S] 0.1 [In] [HIn] [In] = 0.1, pH = pKIn ± 1 pH range = pKIn 1 **Indicator pH range Colour change** (a) Phenolphthalein 8.3 – 10.0 Colourless – Red (b) Methyl orange 3.1 – 4.4 Red – orange (c) Methyl red 4.2 – 6.3 Red – Blue (d) Litmus 5.0 – 8.0 Red – Blue (e) Phenol red 6.8 – 8.4 Yellow – Red

21 **OSTWALD'STHEORY**

Indicators are very weak organic acid or bases. There exists an equilibrium, e.g.

 $\text{HIn} \rightleftharpoons \text{In}^- + \text{H}^+, \text{InOH} \rightleftharpoons \text{In}^+ + \text{OH}^-$ One different One different colour colour colour colour **Phenolphthalein :** $HPh \rightleftharpoons H^+ + Ph^-$ **In acidic medium :** Due to common ion effect of H⁺ $[HPh] > [Ph^-]$, the solution would be colourless In basic medium : $OH^- + H^+ \rightarrow H_2O$ $[Ph^-]$ > $[HPh]$ The solution would be pink **Methyl orange :** $\text{MeOH} \rightleftharpoons \text{Me}^+ + \text{OH}^-$ **In basic medium :** Due to common ion effect of OH– : $[MeOH] > [Me^+]$ the solution would be light yellow **In acidic medium :** H^+ + OH⁻ \rightarrow [H₂O] Hence $[Me⁺]$ > $[MeOH]$ The solution would be red **pH of the indicator :** or $\text{HIn} \rightleftharpoons \text{H}^+ + \text{In}^$ quinorium, e.g.

HIn \rightleftharpoons In⁻+H⁺, InOH \rightleftharpoons In⁺+OH⁻

One different One different

colour colour colour colour colour
 nolphthalein: HPh \rightleftharpoons H⁺+ Ph⁻

[HPh] > [Ph⁻], the solution would be colour

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

Where K_{In} = dissociation constant of the indicator. $pH = pK_{In}$ Then, $[HIn] = [In^-]$ Thus indicator changes its colour at this pH.

QUINONOID & BENZENOID THEORY

Indicator is an equilibrium mixture of two tautomeric forms. One is of structure benzenoid and other of quinonoid structure.

Benzenoid form colourless or light colour

$$
\left\langle _____\text{B-form}
$$

Quinonoid form deep colour

$$
\leftarrow
$$
 \leftarrow Q-form

Equilibrium : Between two tautomeric forms B –from \rightleftharpoons O–from

Example :

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

- HPh B–form acidic medium (colourless)
- Ph– Q–form basic medium (pink)
	- $MeOH \rightleftharpoons Me^+ + OH^-$
- MeOH B–form...............basic medium (light yellow)
- Me+ Q–form acidic medium (red)
- **(i) Choice of indicator :** For a particular titration the indicator is chosen on reviewing two factors :
	- (a) pH near the end point in the neutralization curve [pH vs volume of alkali added graph]
	- (b) Colour change interval of the indicator HPh pH [8.3 – 10.0] MeOH pH[3.1 – 4.4] Ph Red $pH[6.8 - 8.4]$

Acid–Base Titration

Titration of strong acid with strong base : Sharp change of pH near the end point $[4 - 10]$ suitable indicator : HPh $(8.3 - 10)$ and MeOH $(3.1 - 4.4)$ **Titration of a strong acid with week base :** Sharp change of pH near the end point $[4 - 7]$ Suitable indicator : MeOH $(3.1 – 4.4)$ HPh can not be used $(8.3 - 10)$

Titration of weak acid with strong base :

Sharp change of pH near the end point $[7 - 10]$.

Suitable indicator : HPh ; MeOH can not be used $(3.1 - 4.4)$ **Titration of a weak acid with a weak base :**

No sharp change of pH $[6.5 - 7.5]$

For rough titration the indicator is : Phenol red $(6.8 - 8.4)$

Titration of sodium carbonate with HCl :

Suitable indicator : MeOH

Example 26 :

The pH at which an acid indicator with $K_a = 1 \times 10^{-5}$ changes colours when indicator concentration is 1×10^{-5} M is $$ ble indicator : MeOH (3.1 – 4.4) HPh can not be used

10)

10)

p change of pH near the end point [7 – 10].

ble indicator : HPh ; MeOH can not be used (3.1 – 4.4)

thion of a weak acid with a weak base :

harp change of of weak acid with strong base :

ange of pH near the end point [7 – 10].

Indicator: HPh; MeOH can not be used (3.1–4.4)

of a weak acid with a weak base :

change of pH [6.5–7.5]

titration the indicator is : Phenol red

(A) 4 (B) 5 (C) 6 (D) 3

Sol. (B). For acid indicator HIn

$$
\begin{array}{rcl} \n\text{HIn} & \rightleftharpoons & \text{H}^+ + \text{In}^- \\ \n\text{Colour (A)} & \text{Colour (B)} \n\end{array}
$$

 $I_n = \frac{[H - I]I_n}{[H I_n]}$ when $[I_n^-] = [H I_n]$ indicator c $^{-}$] = [HI_n] indicator changes

colour.

 $K_{In} = [H^+] = 1 \times 10^{-5}$: pH = 5

TRY IT YOURSELF-3

Q.1 Mg(OH)₂ has a solubility product equal to 1.2×10^{-11} . A solution contains 0.10 moles of magnesium ions per litre. The maximum pH of the solution from which $Mg(OH)_{2}$ will not precipitate is – (A) 4.96 (B) 9.04 (C) 12.13 (D) 6.94 **Q.2** Given the solubility product of $Pb_3 (PO_4)_2$ is 1.5×10^{-32} . Determine the solubility in gms/litre. **Q.3** If K_{sp} for HgSO₄ is 6.4 \times 10⁻⁵, then solubility of this substance in moler per m^3 is $-$ (A) 8×10^{-3} (B) 6.4×10^{-5} (C) 8×10^{-6} (D) None of these **Q.4** 50 litre of a solution containing 10^{-5} mole of is mixed with 50 litre of a 2×10^{-7} M HBr solution. [Ag⁺] in resultant solution is [Given : K_{sp} (AgBr) = 5×10^{-13}]
(A) 10^{-5} M (B) 10^{-6} M (A) 10^{-5} M (C) 10^{-7} M (D) None of these **Q.5** 0.1 M of HA is titrated with 0.1 M NaOH, Calculate the pH at end point. Given $K_a(HA) = 5 \times 10^{-6}$ and $\alpha \ll 1$ **Q.6** 2.5 mL of (2/5) M weak monoacidic base $(K_b = 1 \times 10^{-12})$ at 25°C) is titrated with (2/15) M HCl in water at 25°C. The concentration of $H⁺$ at equivalence point is $(K_2 = 1 \times 10^{-14} \text{ at } 25^{\circ}\text{C})$ (A) 3.7 × 10⁻¹³ M (B) 3.2 10⁻⁷ M (C) 3.2×10^{-2} M (D) 2.7×10^{-2} M

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IONIC EQUILIBRIUM

Q.7 In 1 L saturated solution of AgCl $[K_{\rm sn} \, (\rm AgCl) = 1.6 \times 10^{-10}]$, 0.1 mol of CuCl $[K_{\text{sp}}^{\text{P}}(CuCl) = 1.0 \times 10^{-6}]$ is added. The resultant concentration of Ag⁺ in the solution is 1.6×10^{-x} . Value of x is **Q.8** The K_{sp} of Ag₂CrO₄ is 1.1×10^{-12} at 298K. The solubility $(\text{in mol/L}) \text{ of } \text{Ag}_2\text{CrO}_4 \text{ in a } 0.1\text{M AgNO}_3 \text{ solution is } -$ (A) 1.1×10^{-11} (B) 1.1×10^{-10} (C) 1.1×10^{-12} (D) 1.1×10^{-9} **ANSWERS (1)** (B) **(2)** 2.92×10^{-7} **(3)**(D)

IMPORTANT POINTS

-
- * Strong acids : HCl, HNO_3 , H_2SO_4 , HClO₄, etc.
- * Weak acids : HCN, H_2CO_3 , H_3PO_4 , CH₃CHOOH, H₂S, etc.
-
- * Strong bases : NaOH, KOH, CsOH, Ba(OH)₂

* Weak bases : Mg(OH)₂, NH₄OH, Ca(OH)₂, LiOH, etc.
- Conjugate acid = Given species + H^+
- Conjugate base = Given species H^{+} .
- If the acid is strong its conjugate base will be weak and vice versa.
- Water posseses a high dielectric constant and high dipole moment. It is amphoteric in character.
- When water functions as an acid, its conjugate base is OH– . When it functions as a base, its conjugate acid is
- H_3O^+ ion.
* Strength of all strong acids (HCl, H_2SO_4 , HNO₃, HClO₄ * etc.) in water is same and become equal to the strength of
- H_3O^+ ion. This is called levelling effect of water.
The strength of hydracids increases as we go down a group because of increase in the size of the atom or anion. For example, HF < HCl < HBr < HI

$$
H_2O < H_2S < H_2Se < H_2Te.
$$

- and $H_2O < H_2S < H_2Se < H_2Te$.
In the case of oxyacids of the same element, the acidic nature increases as the oxidation number increases. $HClO₄ > HClO₃ > HClO₂ > HClO.$ $(+7)$ $(+5)$ $(+4)$ $(+1)$
- Among oxyacids of the same type formed by different elements. acidic nature increases as the electronegativity increases.
- $HIO_4 < HBrO_4 < HClO_4$
- (2.5) (2.8) (3.0)
- Among oxyacids of an element. acidic strength increases as the number of non-hydrogenated oxygen atoms increases or oxidation number increases.

 $HCIO₄ > HClO₃ > HClO₂ > HClO$ Oxidation No. $+7$ +5 $+3$ Non hydrogenated oxygen atoms $3 \t 2 \t 1 \t 0$

Acidic nature of the hydride increases as the electronegativeity of the central atom increases.

$$
CH_4 < NH_3 < H_2O < HF
$$

(2.5) (2.5) (2.5) (4.0)

- (2.5) (3.5) (3.5) (4.0)
- The basic nature of hydroxides and oxides of IA and IIA group elements increases with increase in electropositivity. LiOH < NaOH < KOH < RbOH < CsOH.
	- $Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(O$
- The basic nature of hydrides of same group decreases with the size of the central atom. This is because of decrease in the ability to donate the lone pair.

 $NH_3 > PH_3 > AsH_3 > SbH_3 > BiH_3$

- All Lewis bases are also the Bronsted Lowry bases, but all Lewis acids are not Bronsted Lowry acids. Lewis acids are electron deficient molecules such as BF₃, AlCl₃, GaCl₃ etc.
- Some examples of Lewis acids are Na⁺, Ag⁺, Cu²⁺, Al³⁺, Fe^{3+} , O = C = O, O = S = O, BF₃, AlCl₃, ZnCl₂, SnCl₂ etc. Some examples of Lewis bases are Cl⁻, Fe⁻, Br⁻, H₂O, N_{H₃,} ROH, R_2O , R_2S , CO, NO, C_2H_4 , C_2H_2 etc.

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

$$
K_{\rm w} = [H_3O^+][OH^-]
$$

$$
K_w^{\text{w}} = [H_3^3O^+]^2 = [H^+]^2 = [OH^-]^2
$$

- Heat of neutralisation :
	- 1. $S_A + S_B =$ Maximum heat

2.
$$
S_A + W_B = \text{Less}
$$

3. $W_A + W_B =$ Lessed

4. ^H ⁺ + OH¯ H2O + 13.6 Kcal

- \Rightarrow 13.6 × No. of Equivalent of water
- Ionic product of water is the product of concentration of H+ ions and OH– ions in pure water. It is constant at constant temperature.
- The ionic product of water increases with increase in temperature
- Salt hydrolysis is reverse of neutralisation
- The aqueous solution of salts of weak acid and strong base is alkaline in nature ($pH > 7$).
- The aqueous solution of strong acid and weak base is acidic in nature ($pH < 7$).
- The nature of salts of weak acid and weak base on hydrolysis depends upon the relative hydrolysis of the cation or anion of the salt.
- * The aqueous solution of salts of strong acid and strong base is neutral
- Solubility product is defined as the product of ionic concentration of sparingly soluble electrolyte in a saturated solution and is constant at constant temperature.

- A solution is saturated when ionic product is equal to solubility product.
- When ionic product exceeds the solubility product the precipitation takes place.
- HF is a strong acidic solvent. A strong acid in HF behave like weak base.

* Connection of [Salt] should in
$$
\frac{1}{1}
$$

- * Strength of acid $\text{HClO}_4 \geq \text{H}_2\text{SO}_4 \geq \text{HCl} \geq \text{HNO}_3 \geq \text{H}_3\text{O}^+ \geq \text{HSO}_4 \geq \text{H}_3\text{PO}_4$ > CH₃COOH > H₂CO₃ > H₂S > NH₄⁺ > HCN > C₆H₅OH
> H - OH > R - OH > NH₃ > R - NH₂ > H₂ > CH₄
- Solubility of sparingly soluble in acids or bases : There are many salts of weak acids or weak bases which are almost insoluble in water but their solubility increases with decreasing P^H . The salts containing Carbonates, Phosphates, Fluorids acetates etc. are soluble at low P^H . The salts of weak bases similarly dissolve at higher P^H (low acidity) CaF_2 is insoluble in water but it goes into **Exam** solution when some acid is added.

ADDITIONAL EXAMPLES

Example 1 :

Explain Ca C_2O_4 is insoluble in CH₃COOH but soluble in dilute HCl.

Sol. CH₃COOH being weaker acid than oxalic acid does not decompose CaC_2O_4 ; on the other hand HCl being stronger acid forms Ca^{+2} and oxalate ion to pass CaC_2O_4 in solution state. $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$

Example 2 :

Find out solubility of AgCl_(Ksp) in solution of NaCl (conc. C) **Sol.** AgCl \rightleftharpoons Ag⁺ + Cl^{\leftarrow}

First, S S S
\nK_{SP} =
$$
[Ag^+][Cl^-]
$$
; K_{SP} = S²
\nAgCl $\rightleftharpoons Ag^+ + Cl^-$ NaCl $\rightleftharpoons Na^+ + Cl^-$ (ii
\nS¹ (S¹ + C) C C
\nK_{SP} = [S¹] [S¹ + C]
\nK_{SP} = S¹C (ii
\nK_{SP} = S¹C
\nK_{SP} = S¹C
\nExamp
\nF
\nF
\nK
\nE
\nE
\nE
\nE
\nE
\nE

Example 3 :

Find out solubility of KCl (K_{SP}) in solution of CaCl₂ (conc. C) **Sol.** $KC \rightleftharpoons K^+ + Cl^-$

$$
K_{SP} = S^2
$$

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

\n
$$
C \quad C \quad 2C
$$

\n
$$
KCl \rightleftharpoons K^+ + Cl^-
$$

\n
$$
S^1 \quad S^1 + 2C
$$

\n
$$
K_{SP} = [S^1][S^1 + 2C]
$$

$$
K_{SP} = S12 + 2S1C
$$

$$
S1 = \frac{K_{sp}}{2C}
$$

Example 4 :

Mole What should be the concentration of H^+ and OH^- ions in litre a solution having 4g of NaOH dissolve 100ml of water at 25°C? $\frac{1}{\sqrt{1 + \frac{1}{n}}}$
 $\frac{1}{\sqrt{1 + \frac{1}{n}}}$ **CHEMISTRY**

and OH⁻ ions in

bml of water at
 $\frac{4}{40 \times 0.1} = 1 \text{ N}$

ion/litre **STUDYMATERIAL: CHEMISTRY**

= $S^{1^2} + 2S^1C$

= $\frac{K_{sp}}{2C}$

hould be the concentration of H⁺ and OH⁻ ions in

on having 4g of NaOH dissolve 100ml of water at
 $y = \frac{Weight of substance}{Equivalent weight} \times Volume = \frac{4}{40 \times 0.1} = 1 N$

ormality is 1N, STUDYMATERIAL: CHEMISTRY

= $S^{1^2} + 2S^1C$

= $\frac{K_{sp}}{2C}$

d be the concentration of H⁺ and OH⁻ ions in

aving 4g of NaOH dissolve 100ml of water at

Weight of substance
 $\frac{1}{2}$ wing integrals and $\frac{4}{40 \times 0.1$ SP = $S^2 + 2S^1C$
 $= \frac{K_{sp}}{2C}$
 \therefore

Should be the concentration of H⁺ and OH⁻ ions in

tion having 4g of NaOH dissolve 100ml of water at
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2$

Sol. Normality =
$$
\frac{\text{Weight of substance}}{\text{Equivalent weight} \times \text{Volume}} = \frac{4}{40 \times 0.1} = 1 \text{ N}
$$

When normality is 1N,

[OH–] = 1 gram equivalent ion/litre

$$
[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{1} = 10^{-14} \text{ gram mole ion/litre}
$$

Example 5 :

Find out the pOH of 10^{-3} M NH₄OH if K_b = 10^{-5}

Sol. [OH⁻] =
$$
\sqrt{K_b \times C} = \sqrt{10^{-5} \times 10^{-3}} = \sqrt{10^{-8}} = 10^{-4}
$$

pOH = $-\log$ [OH⁻] = $-\log$ [10⁻⁴] = 4 log 10 = 4

Example 6 :

Find out the K_h of centi normal $[10^{-2}N]$ solution of NH₄Cl [SA-WB) if dissociation constant of $NH₄OH$ is 10⁻⁶ and $K_{\text{w}} = 10^{-14}$. Find out degree of hydrolysis and also find $[H^+]$, pH of solution ? Given: $K_w = 10^{-14}$, $K_b = 10^{-6}$ ien normality is 1N,
 H^{-}] = 1 gram equivalent ion/litre
 F_1] = $\frac{K_w}{[OH^-]}$ = $\frac{10^{-14}}{1}$ = 10^{-14} gram mole ion/litre

5:

S.

d out the pOH of 10^{-3} M NH₄OH if K_b = 10^{-5}
 H^{-}] = $\sqrt{K_b \times C} = \sqrt{10^{-5}$ Equivalent weight × Volum 40×0.1

normality is 1N,

= 1 gram equivalent ion/litre
 $\frac{K_w}{[OH^-]} = \frac{10^{-14}}{1} = 10^{-14}$ gram mole ion/litre

tt the pOH of 10⁻³ M NH₄OH if $K_b = 10^{-5}$

= $\sqrt{K_b \times C} = \sqrt{10^{-5} \times 10^{-3}} = \sqrt{10^{-$

 -8

Sol. (i)
$$
K_h = \frac{K_w}{K_b} = \frac{10^{-14}}{10^{-6}} = 10^{-8}
$$

$$
[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{1} = 10^{-14} \text{ gram mole ion/litre}
$$
\n
$$
n\text{ple 5:}
$$
\nFind out the pOH of 10⁻³ M NH₄OH if K_b = 10⁻⁵\n
$$
[OH^-] = \sqrt{K_b \times C} = \sqrt{10^{-5} \times 10^{-3}} = \sqrt{10^{-8}} = 10^{-4}
$$
\n
$$
pOH = -\log [OH^-] = -\log [10^{-4}] = 4 \log 10 = 4
$$
\n
$$
n\text{ple 6:}
$$
\nFind out the K_h of centi normal [10⁻² N] solution of NH₄Cl\n[SA-WB) if dissociation constant of NH₄OH is 10⁻⁶ and\nK_w = 10⁻¹⁴. Find out degree of hydrolysis and also find\n[H^+], pH of solution?\nGiven: K_w = 10⁻¹⁴, K_b = 10⁻⁶\n(i) K_h = $\frac{K_w}{K_b} = \frac{10^{-14}}{10^{-6}} = 10^{-8}$ \n(ii) $h = \sqrt{\frac{K_h}{K_C}} = \sqrt{\frac{10^{-8}}{10^{-2}}} = \sqrt{10^{-6}} = 10^{-3}$ \n(iii) [H^+] = Ch = 10⁻² × 10⁻³ = 10⁻⁵\n(iv) pH = -\log [H^+] = -\log [10^{-5}] = +5 \log 10 = +5 × 1 = 5\n
$$
n\text{ple 7:}
$$

Example 7 :

Find out pH, h and [OH–] of milli molar solution of KCN [10^{-3} M]. If the dissociation constant of HCN is 10^{-7} .

Example 6:
\nFind out the K_h of centi normal [10⁻² N] solution of NH₄Cl
\n[SA-WB) if dissociation constant of NH₄OH is 10⁻⁶ and
\nK_w = 10⁻¹⁴. Find out degree of hydrolysis and also find
\n[H⁺], pH of solution?
\nGiven: K_w = 10⁻¹⁴, K_b = 10⁻⁶
\n**Sol.** (i) K_h =
$$
\frac{K_w}{K_b} = \frac{10^{-14}}{10^{-6}} = 10^{-8}
$$

\n(ii) $h = \sqrt{\frac{K_h}{K_C}} = \sqrt{\frac{10^{-8}}{10^{-2}}} = \sqrt{10^{-6}} = 10^{-3}$
\n(iii) [H⁺] = Ch = 10⁻² × 10⁻³ = 10⁻⁵
\n(iv) pH = -log [H⁺] = -log [10⁻⁵] = +5 log 10 = +5 × 1 = 5
\n**Example 7:**
\nFind out pH, h and [OH⁻] of milli molar solution of
\nKCN [10⁻³ M]. If the dissociation constant of HCN is 10⁻⁷.
\n**Sol.** (i) pH = $7 + \frac{1}{2} pK_a + \frac{1}{2} log C = 7 + \frac{1}{2} × 7 + \frac{1}{2} log 10^{-3}$
\n= $7 + \frac{7}{2} - \frac{3}{2} log 10 = \frac{14 + 7 - 3}{2} = \frac{21 - 3}{2} = \frac{18}{2} = 9$
\n(ii) $h = \sqrt{\frac{K_h}{c}} = \sqrt{\frac{K_w}{K_a \times C}} = \sqrt{\frac{10^{-14}}{10^{-7} \times 10^{-3}}}$

EXEC EQUILIBRIUM
\n
$$
= \sqrt{10^{-14} \times 10^{10}} = \sqrt{10^{-4}} = 10^{-2}
$$

\n(iii) [OH⁻] = $\sqrt{\frac{K_w \times C}{K_a}} = \sqrt{\frac{10^{-14} \times 10^{-3}}{10^{-7}}}$
\n
$$
= \sqrt{10^{-17} \times 10^{+7}} = \sqrt{10^{-10}} = 10^{-5}
$$

\n
$$
= \sqrt{10^{-17} \times 10^{+7}} = \sqrt{10^{-10}} = 10^{-5}
$$

\n
$$
= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
$$

\n
$$
= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
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= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
$$

\n
$$
= \sqrt{10^{-17} \times 10^{-7}} = \sqrt{10^{-10}} = 10^{-5}
$$

Example 8 :

Hydrolysis constant of salts, M_1X and M_2X formed from strong acid and weak base, are 10^{-7} and 10^{-4} , respectively. For the base, M_3OH , $K_b = 10^{-4}$. What will be the decreasing order of strength of the base ? $(1) M_3OH, M_1OH, M_2OH$ (2) M₂OH, M₃OH, M₁OH

- (3) M₁OH, M₂OH, M₃OH (4) Equal for all
- **Sol.** (1). Strength of base $\propto K_b$

$$
K_{h} = \frac{K_{w}}{K_{b}}
$$
 then $K_{b} = \frac{K_{w}}{K_{h}}$
\n
$$
K_{h} = \frac{10^{-14}}{10^{-7}} = 10^{-7}
$$

\n
$$
K_{sp} = 4s^{3}; 4 \times 10^{-12}
$$

\n
$$
K_{sp} = 4s^{3}; 4 \times 10^{-12}
$$

\nExample 11:
\nIf K_{sp} of saturated so
\nfind out its pH.
\n
$$
K_{sp} = 4s^{3}; 4 \times 10^{-12}
$$

(ii)
$$
M_2x = K_b = \frac{10^{-14}}{10^{-4}} = 10^{-10}
$$

\n(iii) $M_3x = K_b = 10^{-4}$
\nTherefore, M₃OH > M₁OH > M₂OH
\n $pH = 1$

Example 9 :

What should be the percentage of hydrolysis and pH of

$$
\frac{N}{100}
$$
 KCN, if the dissociation constant of
HCN = 1.3 × 10⁻⁹ and K_w = 1 × 10⁻¹⁴ ?

Sol.
$$
x = \sqrt{\frac{K_w}{K_a \times C}}
$$
; $K_w = 10^{-14}$, $K_a = 1.3 \times 10^{-9}$, $C = \frac{N}{100} = 0.01$

K_b K_h Lxaup
\nif K_g of Stautated solution of Mg(OH)₂ is 4 × 1
\nfind out its pH.
\n(a) M₁x = K_b =
$$
\frac{10^{-14}}{10^{-7}} = 10^{-7}
$$

\n(b) M₂x = K_b = $\frac{10^{-14}}{10^{-4}} = 10^{-10}$
\n(cii) M₂x = K_b = $\frac{10^{-14}}{10^{-4}} = 10^{-10}$
\n(diii) M₃x = K_b = 10⁻⁴
\n $\frac{5}{8} = (10^{-12})^{1/3} = 10^{-4}$
\n $\frac{100}{100} = -\log(2 \times 10^{-4}) = -\log 2 + 4 = -0.3010$
\nTherefore, M₃OH > M₁OH > M₂OH
\n**Example 9:**
\nWhat should be the percentage of hydrolysis and pH of
\nHCl = 1.3 × 10⁻⁹ and K_w = 1 × 10⁻¹⁴?\n\n
$$
x = \sqrt{\frac{K_w}{K_a \times C}}; K_w = 10^{-14}, K_a = 1.3 × 10^{-9}, C = \frac{N}{100} = 0.01
$$
\n
$$
x = \sqrt{\frac{10^{-14}}{1.3 \times 10^{-9} \times 10^{-2}}} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 10^{-2}
$$
\n
$$
K_a = 1.3 × 10^{-9} = 2.77 × 1
$$

Example 10 :

Solubility product of AgCl is 2.8×10^{-10} at 25°C. Calculate solubility of the salt in 0.1 MAgNO₃ solution

- (C) 3.2×10^{-9} mole/litre (A) 2.8 \times 10⁻⁹ mole/litre (B) 2.8×10^{-10} mole/litre (D) 3.2×10^{-12} mole/litre
- **Sol. (A).** In 0.1 M AgNO₃ $\text{AgNO}_3 \rightleftharpoons \text{Ag}^+ + \text{NO}_3^ AgCl \rightleftharpoons Ag^+ + Cl^ K_{sp} = [Ag^{+}] [Cl^{-}]$ **10**
 10 Example 10 COMMOVANGED LEARNING

10:

bility product of AgCl is 2.8 × 10⁻¹⁰ at 25°C. Calculate
 2.8×10^{-9} mole/litre (B) 2.8 × 10⁻¹⁰ mole/litre
 1.2×10^{-9} mole/litre (D) 3.2 × 10⁻¹² mole/litre

10 **EDENTAINMENTAINS**

FORMADVANCED LEARNING

of the salt in 0.1 MAgNO₃ solution
 10^{-9} mole/litre (B) 2.8 × 10⁻¹⁰ mole/litre
 0^{-9} mole/litre (D) 3.2 × 10⁻¹² mole/litre
 M AgNO₃
 $D_3 \rightleftharpoons Ag^+ + NO_3^-$,
 $\rightleftharpoons Ag$ **Solution Analytical Set of AgCl is 2.8** × 10⁻¹⁰ at 25°C. Calculate
ty of the salt in 0.1 MAgNO₃ solution
 $\times 10^{-9}$ mole/litre (B) 2.8 × 10⁻¹⁰ mole/litre
 $\times 10^{-9}$ mole/litre (D) 3.2 × 10⁻¹² mole/litre
0.1 MAgNO

Now $[Ag^+]$ can be taken as $[AgNO_3]$ while [Cl–] is the solubility of AgCl

$$
\therefore \quad \text{Cl} = \frac{\text{K}_{\text{sp}}}{\lceil \text{Ag}^+ \rceil} = \frac{2.8 \times 10^{-10}}{0.1}
$$

 \therefore Solubility of AgCl = 2.8 \times 10⁻⁹ mole/litre

 K_h If K_{sp} of saturated solution of Mg(OH)₂ is 4×10^{-12} , then find out its pH.

 $\frac{10^{-14}}{10^{-7}} = 10^{-7}$ **Sol.** Mg(OH)₂ \rightleftharpoons Mg⁺² + 2OH⁻¹ (i) $M_1x = K_b = \frac{10^{-14}}{10^{-7}} = 10^{-7}$

(ii) $M_2x = K_b = \frac{10^{-14}}{10^{-4}} = 10^{-10}$

(iii) $M_3x = K_b = 10^{-4}$

Therefore, $M_3OH > M_1OH > M$ 1 s 2s $K_{\text{sp}} = 4s^3$; $4 \times 10^{-12} = 4s^3$; $10^{-12} = s^3$ $s = (10^{-12})^{1/3} = 10^{-4}$ $OH^- = 2s = 2 \times 10^{-4}$ $pOH = -\log(2 \times 10^{-4}) = -\log 2 + 4 = -0.3010 + 4 = 3.7$ $pH = 14 - 3.7 = 10.3$ turated solution of Mg(OH)₂ is 4×10^{-12} , then

pH.
 \Rightarrow Mg⁺² + 2OH⁻
 \Rightarrow Mg⁺² + 2OH⁻
 s 2s
 $10^{-12} = 4s^3$; $10^{-12} = s^3$
 $10^2 = 10^{-4}$
 2×10^{-4}
 $g(2 \times 10^{-4}) = -\log 2 + 4 = -0.3010 + 4 = 3.7$
 $\therefore 7 = 10.3$

Example 12 :

The pH of blood stream is maintained by a proper balance of H_2CO_3 and NaHCO₃ concentrations. What volume of 5M NaHCO₃ solution should be mixed with 10mL of blood which is 2M in H_2CO_3 in order to maintain pH of 7.4. K_a for H_2CO_3 in blood is 7.8 \times 10⁻⁷ ? $4 \times 10^{-12} = 4s^3$; $10^{-12} = s^3$
 $/3 = 10^{-4}$
 2×10^{-4}
 2×10^{-4}) $= -\log 2 + 4 = -0.3010 + 4 = 3.7$
 $7 = 10.3$

alood stream is maintained by a proper balance

and NaHCO₃ concentrations. What volume of
 $\frac{1}{3}$ solut ined by a proper balance
rations. What volume of
ixed with 10mL of blood
maintain pH of 7.4.
 \neg^{-7} ?
on mixed = x mL
i x mL of 5M NaHCO₃
mL of 2M H₂CO₃
(0.005x)
ion

Sol. Let the volume of NaHCO₃ solution mixed = x mL Number of moles of NaHCO₃ in x mL of 5M NaHCO₃

solution =
$$
\frac{5 \times x}{1000} = 0.005 \times \text{mol}
$$

Number of moles of $\rm H_2CO_3$ in 10 mL of 2M $\rm H_2CO_3$

solution =
$$
\frac{2 \times 10}{1000}
$$
 = 0.02 mol

pH of the solution = 7.4 K_a for H₂CO₃ = 7.8 \times 10⁻⁷ According to Henderson's equation

$$
pH = -\log K_a + \log \frac{[salt]}{[acid]}
$$

7.4 = -\log (7.8 × 10⁻⁷) + \log \left(\frac{0.005x}{0.02}\right)
7.4 = (7 - 0.892) + \log (0.25x)

 $7.4 = 6.108 + log(0.25x)$ $log(0.25x) = 7.4 + 6.108 = 1.292$

$$
0.25 \text{ x} = \text{Antilog } 1.292 = 19.59 \quad \therefore \quad \text{x} = \frac{19.59}{0.25} = 78.36
$$

Thus, the volume of 5M NaHCO₃ solution to be mixed $= 78.36$ mL.

Example 13 :

Find the pH of a 2 litre solution which is 0.1 M each with respect to CH_3COOH and $(CH_3OO)_2Ba$.

Sol. pH = pK_a + log
$$
\frac{[CH_3COO^-]}{[CH_3COOH]}
$$

\n $pK_a = -log(1.8 \times 10^{-5}) = 4.7447$
\n[CH₃COO⁻] = 2 × [(CH₃COO)₂Ba) = 0.2 M
\n[CH₃COOH] = 0.1 M

$$
pH = 4.7447 + \log \frac{0.2}{0.1} = 5.046
$$

Example 14 :

When a decinormal solution of acetic acid gets 1.3%
ionized than find the value of ionization constant ionised, then find the value of ionisation constant.

Sol. Ionisation of acetic acid
$$
\frac{1.3}{100} \times 0.1 = 0.0013
$$

 $(Decimal = 0.1 N)$

 $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
C 0 0 Initial conc. C 0 0 0 Conc. at equ. $(1 - 0.0013) \text{ C} (0.0013) \text{ C} (0.0013) \text{ C}$ 0.9987 C 0.0013 C 0.0013 C

$$
K_{a} = \frac{[H^{+}][CH_{3}COO^{-}]}{[CH_{3}COOH]}
$$
; $K_{a} = \frac{0.0013 \times 0.1 \times 0.0013}{0.9987}$ Here, $AccO^{-1}$

Example 15 :

Hydrolysis of NaCl does not take place, because –

- (A) Na⁺ ions are surrounded by negative ends of water molecules.
- (B) Na^+ and Cl⁻ are present in ionic state.
- (C) It is a monovalent salt
- (D) $Na⁺$ is a weak acid and Cl⁻ is a weak base.
- **Sol.** (4). The weak acid Na^+ will not react with negative OH⁻. Similarly, the weak base Cl⁻ will not react with H^+ . Therefore, hydrolysis will not occur.

Example 16 :

KCN in solution gets 3.7% hydrolysed at 100 litre dilution at 100 litre dilution. If the dissociation constant of HCN is 7.2×10^{-10} , find the value of K_w.

Sol. KCN is formed from KOH (strong base) & HCN (weak

STUDYMATERIAL: CHEMISTRY
\nlog (0.25x)
\n
$$
7.4 + 6.108 = 1.292
$$

\nlog 1.292 = 19.59 $\therefore x = \frac{19.59}{0.25} = 78.36$
\n $= 78.36$ **28.3**
\n $= 78.36$ **29.3**
\n $= 78.36$ **20.3**
\n $= 78.36$ **21.2**
\n $= 78.36$ **22.3**
\n $= 78.36$ **23.4**
\n $= 78.36$ **24.5**
\n $= 78.36$ **25.6**
\n $= 78.36$ **26.6**
\n $= 78.36$ **27.2**
\n $= 78.36$ **28.10**
\n $= 78.36$ **29.1**
\n $= 78.36$ **20.1**
\n $= 78.36$ **21.10**
\n**20.1**
\n $= 78.36$ **21.10**
\n**21.10**
\n**22.10**
\n**23.10**
\n**24.10**
\n**21.10**
\n**2**
\n**22.10**
\n**23.10**
\n**24.10**
\n**25.10**
\n**26.10**
\n**27.10**
\n**29.1**
\n**20.10**
\n**21.10**
\n**21.10**
\n**21.10**
\n**21.10**
\n**21.10**
\n**21.10**
\n**21.10**
\n**2**

Example 17 :

A solution is 0.1 molar in relation to AcOH and is 0.2 molar in relation to AcONa. What should be the concentration of Ac⁻¹, if $x = 0.2$?

Thus, the volume of 5M NalfCO₃ solution to be mixed
\n=78.36 mi.
\n
$$
K_a = 7.2 \times 10^{-10}, v = 100, x^2 = \frac{K_w \times 100}{K_a} = 0.03772 = \frac{K_w \times 100}{K_w \times 100} = 0.03772 = \frac{K_w \times 100}{7.2 \times 10^{-10}}
$$
\n
$$
H = pK_a + \log \frac{[CH_3COOH]}{[CH_3COOH]} = \frac{K_w \times 100 - (0.037)^2 \times 7.2 \times 10^{-10}}{100} = 0.986 \times 10^{-14}
$$
\n
$$
M = -\log (1.8 \times 10^{-5}) = 4.7447
$$
\n
$$
CH_3COOH = 0.1 M
$$
\n
$$
M = -\frac{1}{2} = 5.046
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.1} = 5.046
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.1} = 5.046
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.1} = 5.046
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H = 4.7447 + \log \frac{0.2}{0.1} = 5.046
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H = 4.7447 + \log \frac{0.2}{0.1} = 5.046
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$$
H = 4.7447 + \log \frac{0.2}{0.01} = 5.046
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$$
H = 4.7447 + \log \frac{0.2}{0.01} = 5.046
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.01} = 5.046
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.01} = 0.0013
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.01} = 0.0013
$$
\n
$$
H = 4.7447 + \log \frac{0.2}{0.01} =
$$

 AcO^{-1} obtained from $AcONa = 0.2$

Total AcO^{-1} obtained = 0.202

Example 18 :

 K_a for acid HA is 2.5×10^{-8} calculate for its decimolar solution at 25°C. (i) % dissociation (ii) pH (iii) OH⁻ ion concentration

Sol. HA
$$
\rightleftharpoons
$$
 H⁺ + A⁻

$$
\begin{array}{ccc}\nC & 0 & 0 \\
C(1-\alpha) & C\alpha & C\alpha\n\end{array}
$$

$$
K_a = \frac{[H^+][A^-]}{[HA]} \Rightarrow \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)} = C\alpha^2
$$

AcoH ⇒ Na⁺ + AcO⁻¹
\n0.2 0.2
\nHere, ACONa will ionise completely.
\nAcO⁻¹ obtained from AcOH = 0.002
\nAcO⁻¹ obtained from AcONa = 0.2
\nTotal AcO⁻¹ obtained = 0.202
\n
\n**ple 18 :**
\nK_a for acid HA is 2.5 × 10⁻⁸ calculate for its decimolar
\nsolution at 25°C. (i) % dissociation (ii) pH (iii) OH⁻ ion
\nconcentration
\nHA ⇒ H⁺ + A⁻
\nC 0 0
\nC (1-α) Cα Cα
\nK_a =
$$
\frac{[H^+][A^-]}{[HA]} \Rightarrow \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)} = C\alpha^2
$$

\n(i) $\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2.5 \times 10^{-8}}{1/10}}$ (C=1/10 M)
\n= 5 × 10⁻⁴ = 0.05%

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\n(ii) [H⁺] = Cα =
$$
\frac{1}{10} \times 5 \times 10^{-4}
$$
 mol/L. So, pH=5-log 5 = 4.30

\n(iii) [H⁺] [OH⁻] = 1 × 10⁻¹⁴

\n(b) (A) 8.5

\n(c) 6.5

\n(d) 8.5

\n(e) 6.5

\n(f) 6.5

\n(g) 6.5

\n(h) 6.5

\n(iii) [H⁺] [OH⁻] = $\frac{10^{-14}}{5 \times 10^{-4}} = 2 \times 10^{-10}$ mol/L

\nSo, A) 6.4

\n(a) 9.5

\n(b) 6.6

\n(c) 6.5

\n(d) 9.5

\n(e) 1.8 × 10⁻⁵ ; pK_a = 4.75

 $4 - 2 \times 10$ more E

Example 19 :

The degree of dissociation of pure water at 25°C is found to be 1.8×10^{-9} . Find K_w and K_d at 25°C. (A) 3.24 \times 10⁻¹⁸; 5.83 \times 10⁻²⁰ (B) 1 \times 10⁻¹⁴; 1.8 \times 10⁻¹⁵ (C) 1.8×10^{-16} ; 1×10^{-14} (D) 1×10^{-14} ; 1×10^{-14} E **Sol. (B).** Since $\alpha = 1.8 \times 10^{-14}$ $\frac{10^{-14}}{5 \times 10^{-4}} = 2 \times 10^{-10} \text{ mol/L}$ (C)6.5

Sol. (A). pH = 7 + $\frac{1}{2}$ pK_a + $\frac{1}{2}$ log

If K_a = 1.8 × 10⁻⁵; pK_a

sociation of pure water at 25°C is found

Find K_w and K_d at 25°C.

5.83 × 10⁻²⁰ (B)

and for water $C = 1000/18 = 55.56$ $[H^+] = [OH^-] = C\alpha = 55.56 \times 1.8 \times 10^{-9} = 1 \times 10^{-7} M$ Sol. NH₂C $K_w = [H^+] \times [OH^-] = (1 \times 10^{-7})^2 = 10^{-14}$

$$
K_{d} = \frac{[H^{+}][OH^{-}]}{[H_{2}O]} = \frac{K_{w}}{[H_{2}O]} = \frac{10^{-14}}{55.56} = 1.8 \times 10^{-16}
$$
 So, $h = \sqrt{\frac{K_{w}}{K_{b} \cdot C}} = \sqrt{\frac{1000 \times 10^{-16}}{1000}} = 1.8 \times 10^{-16}$

Example 20 :

Calculate pH of the mixture $(25 \text{ mL of } 0.1 \text{ M NH}_4 \text{ OH} + 25$ mL of 0.1 M CH₃COOH)

Given that $K_a = 1.8 \times 10^{-5}$, and $K_b = 1.8 \times 10^{-5}$

Sol. $NH_4OH + CH_3COOH \rightarrow CH_3COONH_4 + H_2O$ Initially milli moles 25×0.1 25×0.1 0 0 $= 2.5$ $= 2.5$

Final milli moles 0 0 2.5 2.5 As salt is formed (salt of weak acid and weak base) and pH will be decided by salt hydrolysis

1.8 × 10⁻¹⁶ (10⁻¹⁴ m) m. W and N₄ are 25 × 0.1 × 2 × 10⁻⁵
\n1.8 × 10⁻¹⁶; 1 × 10⁻¹⁴ (D) 1 × 10⁻¹⁴; 1 × 10⁻¹⁵
\n1.8 × 10⁻¹⁶; 1 × 10⁻¹⁴ (D) 1 × 10⁻¹⁴; 1 × 10⁻¹⁵
\n1.8 × 10⁻¹⁶; 1 × 10⁻¹⁴ (D) 1 × 10⁻¹⁴; 1 × 10⁻¹⁴
\nand for water C = 1000/18 = 55.56
\n[H⁺] = [OH⁻] = Cα = 55.56 × 1.8 × 10⁻⁹ = 1 × 10⁻⁷ M
\nK_w = [H⁺] × [OH⁻] = (1 × 10⁻⁷)² = 10⁻¹⁴
\nK_d =
$$
\frac{[H^+][OH^-]}{[H_2O]} = \frac{K_w}{[H_2O]} = \frac{10^{-14}}{55.56} = 1.8 × 10^{-16}
$$

\n $K_h = \frac{K_w}{K_b} = \frac{10^{-14}}{1.5 × 10^{-14}} = 6.667 × 10^{-14} × 1$
\nor h = 0.816 > 0.1, so we use actual rela
\n20:
\n1.8 h² = 1.8 × 10⁻⁵, and K_h = 1.8 × 10⁻⁵
\n $NH_4OH + CH_3COOH \rightarrow CH_3COONH_4 + H_2O$
\n $1.5 h^2 + h - 1 = 0 ⇒ h = 0.55$
\n1.15 h² + h - 1 = 0 ⇒ h = 0.55
\n1.16 moles 0 0 2.5 2.5 1.5 m
\n1.17 h² = 1.8 ×

Example 21 :

Find out the concentration of $[H^+]$ in 0.1 M CH₃COONa solution $[K_a = 10^{-5}]$

Sol. Salt is [WA – SB]

$$
[H^+] = \sqrt{\frac{K_W \times K_a}{C}} = \sqrt{\frac{10^{-14} \times 10^{-5}}{10^{-1}}}
$$

= $\sqrt{10^{-19} \times 10^{+1}} = \sqrt{10^{-18}} = 10^{-9}$ pF

Example 22 :

Calculate degree of hydrolysis of a mixture containing 0.1 N NH₄OH and 0.1 N HCN.

Sol. Salt is [WA – WB]

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

Example 23 :

What should be the pH of 0.018 M sodium acetate, its

$$
K_a = 1.8 \times 10^{-5} ?
$$

(A) 8.5
(B) 7.5
(C) 6.5
(D) 9.5

¹⁰ [OH] 2 10 mol / L 5 10 [H][OH] 10 ^K K 1.8 10 [H O] [H O] 55.56 **Sol. (A).** pH = ^a 1 1 7 pK log C 2 2 If K^a = 1.8 × 10–5 ; pK^a = 4.75 ⁼ 1 1 ² 7 4.75 log1.8 10 2 2 = 7 + 2.37 – 0.87 = 8.5 K C 1.5 10 1 ^K ¹⁰ K 6.667 10 ^K 1.5 10 ^K Ch 1 ^K K 1 h 1.5

Example 24 :

Calculate degree of hydrolysis, K_h and pH of 1 M urea hydrochloride solution in water, K_b^{th} (Urea) = 1.5×10^{-14} . **Sol.** NH₂CONH₃Cl is a salt of $(SA + WB)$

So,
$$
h = \sqrt{\frac{K_w}{K_b \cdot C}} = \sqrt{\frac{10^{-14}}{1.5 \times 10^{-14} \times 1}}
$$

or $h = 0.816 > 0.1$, so we use actual relation

$$
K_{h} = \frac{K_{w}}{K_{b}} = \frac{10^{-14}}{1.5 \times 10^{-14}} = 6.667 \times 10^{-1}
$$

$$
K_{h} = \frac{K_{w}}{1.5 \times 10^{-14}} = 6.667 \times 10^{-1}
$$

$$
K_{h} = \frac{K_{w}}{K_{b}} = \frac{CH}{1-h} = \frac{1}{1.5}
$$

1.5 h² + h - 1 = 0 \Rightarrow h = 0.55
[H⁺] = ch = 0.55 M ; pH = 0.26

Example 25 :

Calculate degree of hydrolysis (h) and pH of solution obtained by dissolving 0.1 mole of $CH₃COONa$ in water to get 100 lt. of solution (Take K_a of acetic acid = 2×10^{-5})

and for water C = 1000/18 = 55.56
\nby drochlorides solution in water, K_b (Uren) = 1.5 × 10⁻¹⁴,
\nK_q = [H⁺] × [OH⁻] = (x × 16⁻⁷)² = 10⁻¹⁴
\nK_q = [H⁺] × [OH⁻] =
$$
\left[\frac{K_w}{[H_2O]}\right]
$$
 = $\frac{K_w}{[H_2O]}$ = $\frac{10^{-14}}{55.56}$ = 1.8 × 10⁻¹⁶
\n $K_d = \frac{[H^+][OH^-]}{[H_2O]} = \frac{10^{-14}}{[H_2O]}$ = $\frac{10^{-14}}{55.56}$ = 1.8 × 10⁻⁵
\n $\frac{10^{-14}}{100 \times 10^{-14}} = 1.8 \times 10^{-5}$
\n $\frac{10^{-14}}{10^{-14}} = 1.6 \times 10^{-$

HI

The correct statement is

forward direction.

backward direction.

direction.

either direction.

 $Time \longrightarrow$ Equilibrium \longleftarrow Time

(A) Chemical equilibrium in the reaction,

(B) Chemical equilibrium in the reaction,

(C) Chemical equilibrium in the reaction,

(D) Chemical equilibrium in the reaction,

Concentration

HI

 $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ can be attained from either

 $H_2 + I_2$ $H_2 + I_2$

 $2XY \rightleftharpoons X_2 + Y_2$ is 81, what is the value of equilibrium constant for the reaction

$$
XY \rightleftharpoons \frac{1}{2}X_2 + \frac{1}{2}Y_2
$$

(A) 81 (B) 9
(C) 6561 (D) 40.5
For the reaction
$$
{}^{250}C^{(0)} + O^{(0)} \rightarrow {}^{250}C^{(0)}
$$

- $H_2(g) + I_2(g) \rightleftharpoons 2HI$ (g) can be attained from (A) **Q.9** For the reaction, $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ What is K_c when the equilibrium concentration of $[SO_2] = 0.60 M, [O_2] = 0.82 M$ and $[SO_3] = 1.90 M$? (A) $12.229 \text{ L mol}^{-1}$ (B) 24.5 L mol^{-1} (C) 36.0 L mol^{-1} (D) $2.67 \times 10^3 \text{ L mol}^{-1}$
- $H_2(g) + I_2(g) \rightleftharpoons 2HI$ (g) can be attained from **Q.10** The following concentrations were obtained for the formation of NH_3 from N_2 and H_2 at equilibrium at 500K. $[N_2] = 1.5 \times 10^{-2} M$,

 $[H_2] = 3.0 \times 10^{-2}$ M and $[NH_3] = 1.2 \times 10^{-2}$ M. Calculate equilibrium constant.

direction.
\nChemical equilibrium in the reaction,
\n
$$
H_2(g) + I_2(g) \rightleftharpoons 2HI (g) can not be attained from
\n
$$
H_2(g) + I_2(g) \rightleftharpoons 2HI (g) can not be attained from
\n
$$
H_2(g) + I_2(g) \rightleftharpoons 2HI (h) + I_2(g) \rightleftharpoons 2HI (i) + I_2(g) \rightleftharpoons 2III (ii) + I_2(g) \rightleftharpoons 2III (iii) + I_2(g) \rightleftharpoons 2III (iv) + I_2(g) \rightleftharpoons 2III (v) + I_2(g) \rightleftharpoons 2III (v) + I_2(g) \rightleftharpoons 2III (v) + I_2(g) \rightleftharpoons 2III (vi) + I_2(g) \rightleftharpoons 2III (v) + I_2(g) \rightleftharpoons 2III (v
$$
$$
$$

122

Q.11 At a certain temperature only 50% HI is dissociated Q.21 into H_2 and I_2 at equilibrium. The equilibrium constant is

- **Q.12** If the value of equilibrium constant K_c for the reaction, $N_2 + 3H_2 \rightleftharpoons 2NH_3$ is 7. The equilibrium constant for \blacksquare the reaction $2N_2 + 6H_2 \rightleftharpoons 4NH_3$ will be – (A) 49 (B) 7 (C) 14 (D) 28
- **Q.13** At equilibrium, the concentrations of $N_2 = 3.0 \times 10^{-3} M$, $O_2 = 4.2 \times 10^{-3} M$ and $\overline{NO} = 2.8 \times 10^{-3}$ M in a sealed vessel at 800K. What will be K_c for the reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $(A) 0.912$ (B) 0.122 $(C) 0.622$ (D) 0.214

PART-4 : HOMOGENEOUS EQUILIBRIA

- **Q.14** For which of the following reactions, $K_p = K_c$ $(A) PCl_3(g)+Cl_2(g) \rightleftharpoons PCl_5(g)$ $(B) H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ $(C) N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ (D) CaCO₃ (s) \rightleftharpoons CaO (s) + CO₂ (g)
- **Q.15** In a homogeneous system, all the reactants and products are in the (A) different phase (B) same phase (C) gaseous state only (D) liquid and gas state
- **Q.16** For the equilibrium, $2 \text{ NOCl}(g) \rightleftharpoons 2 \text{ NO}(g) + \text{Cl}_2(g)$ the value of the equilibrium constant, K_c is 3.75×10^{-6} at 1069 K. Calculate the K_p for the reaction at this temperature? $(A) 0.112$ (B) 0.033 $(C) 0.33$ (D) 0.098 Solution and and state of the continent and \mathbf{K}_k (D) injudind gas state of the equilibrium. 2 NOCI (g) \Rightarrow 2NO (g) + C(g) = \mathbf{K}_k (S) = $\frac{[\text{Cu}(\text{NO})_2]_2\text{S}]^2}{\frac{(\text{Cu}(\text{NO})_2)\text{S}^2}}$

(B) $\mathbf{K}_2 = \frac{[\text{Cu$
- **Q.17** In the system $X + 2Y \rightleftharpoons Z$, the equilibrium concentrations are, $[X] = 0.06$ mol L^{-1} $[Y] = 0.12 \text{ mol } L^{-1}$, $[Z] = 0.216 \text{ mol } L^{-1}$ Find the equilibrium constant of the reaction. (A) 250 (B) 500 (C) 125 (D) 273
- **Q.18** Which of the following is an example of homogeneous equilibrium? $(A) 2SO₂(g) + O₂(g) \rightleftharpoons 2SO₃(g)$

(B) C(s) + H₂O(g)
$$
\rightleftharpoons
$$
 CO(g) + H₂(g)
(C) CaCO₃(s) \rightleftharpoons CaO(s) + CO₂(g)
(D) NH₄HS (s) \rightleftharpoons NH₃(g) + H₂S(g)

- **Q.19** What is the relation between K_p and K_c for a general reaction, $aA + bB \rightleftharpoons cC + dD$? $(A) K_p = K_c (RT)^{\Delta n}$ $(B) K_c = K_p (RT)^{\Delta n}$ (C) $K_p = \frac{1}{K_c (RT)^{\Delta n}}$ (D) $K_c = \frac{1}{K_p (RT)^{\Delta n}}$ temperature.

IV. The equilibrium constant for the
- **Q.20** 5 moles of PCI_5 are heated in a closed vessel of 5 litre capacity. At equilibrium 40% of PCl_5 is found to be dissociated. What is the value of K_c ?

Q.21 For a reaction, $NH_4COONH_2(s) \rightleftharpoons 2NH_3(g) + CO_2(g)$, the equilibrium pressure is 3 atm. K_p for the reaction will be – $(A) 27$ (B) 4 (C) 3 (D) 9 **EDIMADVANCED LEARNING**
 EDIMADVANCED LEARNING
 n, NH₄COONH₂ (s) \rightleftharpoons 2NH₃ (g) + CO₂(g),

(B) 4

(D) 9
 EROGENEOUS EQUILIBRIA

rium constant for the following reaction

s)+ 5O₂ (g) \rightleftharpoons P₄O₁₀ (**EXECUTE ARABUTE 10**
 EXECUTE:
 EXECUT EXECUTE:

Interaction, NH₄COONH₂(s) $\rightleftharpoons 2NH_3(g) + CO_2(g)$,

equilibrium pressure is 3 atm. K_p for the reaction

be-

P27 (B)4

(D)9
 EXETEROGENEOUS EQUILIBRIA

equilibrium constant for the following reaction

be **EXECUTE ANTAINMAN COONLEASE (2)**

Interior, NH₄COONH₂(s) $\rightleftharpoons 2NH_3(g) + CO_2(g)$,

Ilibrium pressure is 3 atm. K_p for the reaction

(B) 4

(D) 9
 ETEROGENEOUS EQUILIBRIA

Ililibrium constant for the following react ENTADVANCED LEARNING

ENTADVANCED LEARNING
 $2NH_3(g) + CO_2(g)$,
 C_p for the reaction

(DUILIBRIA
 $\frac{1}{[O_2]^5}$
 $=[O_2]^5$

tion(s) show(s) For a reaction, NH₄COONH₂ (s) \implies 2NH₃ (g) + CO₂(g),

the equilibrium pressure is 3 atm. K_p for the reaction

will be –

(A) 27 (B) 4

(C) 3 (D) 9

(C) 3 (D) 9

(C) 3 (D) 9

(C) 4 (C) 3 (D) 9

(A) K_c = $\frac{[$

PART-5 : HETEROGENEOUS EQUILIBRIA

Q.22 The equilibrium constant for the following reaction will be : $P_4(s)$ + 5O₂ (g) \rightleftharpoons $P_4O_{10}(s)$

(A)
$$
K_c = \frac{[P_4][O_2]^5}{[P_4O_{10}]}
$$
 (B) $K_c = \frac{1}{[O_2]^5}$

(C)
$$
K_c = \frac{[P_4O_{10}]}{[P_4][O_2]^5}
$$
 (D) $K_c = [O_2]^5$

- **EXECUTE ARISED**

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ium pressure is 3 atm. K_p for the reaction

(B) 4

(D) 9
 EROGENEOUS EQUILIBRIA

rium constant for the following reaction

(s) + 5O₂(g) \rightleftharpoons P₄O₁₀(s)

²4^{1[O₂]⁵} **EXECUTE ANTIFICATE SET AND CONDUCT AND CONTAINMARY SUBARDINATE SUBARDINATE SUB-

DOM ADVANCED LEARNING

DEVIDED IN THE CONDUCT (B)

27

(B) 4

3

EXETEROGENEOUS EOUILIBRIA**

equilibrium constant for the following reacti **EXECUTE ANTIVE SET AN ART AN ART AND SET AN ART AN ART AND THE SCHOON H₂ (s)** \rightleftharpoons **2NH₃ (g) + CO₂(g),

(B) 4

(B) 4

(D) 9

ETEROGENEOUS EQUILIBRIA**

(D) 9
 ETEROGENEOUS EQUILIBRIA
 $\left[\frac{P_4(s) + 5O_2(g) \rightleftharpoons P$ **Q.23** Which of the following reaction(s) show(s) heterogeneous equilibria? (A) N₂(g) + 3H₂(g) \rightleftharpoons 2NH₃(g) $(B) H₂O(\ell) \rightleftharpoons H₂O(g)$ (C) CaCO₃ (s) \rightleftharpoons CaO (s) + CO₂ (g) (D) Both (B) and (C) GENEOUS EOUILIBRIA

onstant for the following reaction
 $2(g) \rightleftharpoons P_4O_{10}(s)$
 $\frac{5}{5}$
 $\frac{1}{[O_2]^5}$
 $\frac{1}{[O_2]^5}$
 $\frac{1}{[O_2]^5}$
 $\frac{1}{[O_2]^2}$
 $\frac{1}{[O_2]}$
 $\frac{1}{[O_2]}$
 $\frac{1}{[O_2]}$
 $\frac{1}{[O_2]}$
 $\frac{1}{[O_2]}$ **IEOUS EQUILIBRIA**

and for the following reaction
 $\Rightarrow P_4O_{10}(s)$

(B) $K_c = \frac{1}{[O_2]^5}$

(D) $K_c = [O_2]^5$

wing reaction(s) show(s)
 a ?
 $NH_3(g)$
 $s) + CO_2(g)$

ilibrium constant, K_c for the
 $(s) + 4NO_2(g) + O_2(g)$
 $\frac{2(g$ nt for the following reaction
 $\Rightarrow P_4O_{10}(s)$

(B) K_c = $\frac{1}{[O_2]^5}$

(D) K_c = $[O_2]^5$

wing reaction(s) show(s)

a?

2NH₃(g)

s) + CO₂(g)

uilibrium constant, K_c for the

0(s) + 4NO₂(g) + O₂(g)

<u>2(g)</u>] equilibrium pressure is 3 atm. K_p for the reaction

be - P(B)4

(B) 4

(Cu) (B) 4
 $\text{EDEROGENEOLIS EOLILLIBRIA}$
 $\text{equilibrium constant for the following reaction}$
 $\text{Re} : P_4(s) + 50_2(g) \rightleftharpoons P_4O_{10}(s)$
 $K_c = \frac{[P_4][O_2]^5}{[P_4O_{10}]}$ (B) $K_c = \frac{1}{[O_2]^5}$
 $K_c = \frac{[$ (B)4

(D)9

ETEROGENEOUS EOUILIBRIA

uilibrium constant for the following reaction
 $P_4(s)+5Q_2(g) \rightleftharpoons P_4Q_{10}(s)$
 $= \frac{[P_4][Q_2]^5}{[P_4Q_{10}]}$ (B) $K_c = \frac{1}{[Q_2]^5}$
 $= \frac{[P_4Q_{10}]}{[P_4][Q_2]^5}$ (D) $K_c = [Q_2]^5$

of the fo $\frac{O_2 1^5}{O_{10} 1}$ (B) $K_c = \frac{1}{[O_2]^5}$
 $\frac{(D) K_c = [O_2]^5}{O_2 1^5}$ (D) $K_c = [O_2]^5$

e following reaction(s) show(s)

equilibria?
 $H_2 O(g) \Rightarrow 2NH_3(g)$
 $\Rightarrow CaO(s) + CO_2(g)$
 $d(C)$

on for equilibrium constant, K_c for the

io (B) K_c = $\frac{1}{[O_2]^5}$

(D) K_c = $[O_2]^5$

Ilowing reaction(s) show(s)

ibria?
 \Rightarrow 2NH₃(g)

(g)

(0(s) + CO₂(g)

equilibrium constant, K_c for the

:

CuO (s) + 4NO₂(g) + O₂(g)

NO₂(g)]⁴[O₂(g)]

NO₃ (D) 9

HETEROGENEOUS EQUILIBRIA

equilibrium constant for the following reaction

be : $P_4(s) + 5O_2(g) \rightleftharpoons P_4O_{10}(s)$
 $K_c = \frac{[P_4][O_2]^5}{[P_4O_{10}]}$ (B) $K_c = \frac{1}{[O_2]^5}$
 $\zeta_c = \frac{[P_4O_{10}]}{[P_4][O_2]^5}$ (D) $K_c = [O_2]^5$ ETEROGENEOUS EOUILIBRIA

ilibrium constant for the following reaction
 $P_4(s) + 5O_2(g) \rightleftharpoons P_4O_{10}(s)$
 $= \frac{[P_4][O_2]^5}{[P_4O_{10}]}$ (B) $K_c = \frac{1}{[O_2]^5}$
 $= \frac{[P_4O_{10}]}{[P_4][O_2]^5}$ (D) $K_c = [O_2]^5$

of the following rea The equilibrium constant for the following reaction

will be: $P_4(s)+5O_2(g) \rightleftharpoons P_4O_{10}(s)$

(A) $K_c = \frac{[P_4][O_2]^5}{[P_4O_{10}]}$ (B) $K_c = \frac{1}{[O_2]^5}$

(C) $K_c = \frac{[P_4O_{10}]}{[P_4O_{10}]}$ (D) $K_c = [O_2]^5$

(C) $K_c = \frac{[P_4O_{10}$ following reaction(s) show(s)

uilibria?

2) $\rightleftharpoons 2NH_3(g)$
 $O(g)$

CaO (s) + CO₂ (g)

C

C)

or equilibrium constant, K_c for the

is
 $2CuO(s) + 4NO_2(g) + O_2(g)$
 $\frac{1^2[NO_2(g)]^4[O_2(g)]}{u(NO_3)_2(s)]^2}$
 $\frac{g}{1^2}$
 $\frac{1}{2^$ = $\frac{[P_4][O_2]^5}{[P_4O_{10}]}$ (B) K_c = $\frac{1}{[O_2]^5}$

= $\frac{[P_4O_{10}]}{[P_4][O_2]^5}$ (D) K_c = $[O_2]^5$

of the following reaction(s) show(s)

eneous equilibria?
 $y()$, $y=3H_2(g) \rightleftharpoons 2NH_3(g)$
 $y()$, $y=15Q_0(g)$
 $O_3(s)$
- **Q.24** The expression for equilibrium constant, K_c for the following reaction is

$$
2Cu(NO_3)_2(s) \rightleftharpoons 2CuO(s) + 4NO_2(g) + O_2(g)
$$

(A)
$$
K_c = \frac{[CuO(s)]^2 [NO_2(g)]^4 [O_2(g)]}{[Cu(NO_3)_2(s)]^2}
$$

(B)
$$
K_c = \frac{[NO_2(g)]^4 [O_2(g)]}{[Cu(NO_3)_2(s)]^2}
$$

(C)
$$
K_c = [NO_2(g)]^4 [O_2(g)]
$$

(D)
$$
K_c = \frac{[CuO(s)]^2}{[Cu(NO_3)_2(s)]^2}
$$

PART-6 : APPLICATIONS OF EQUILIBRIUM CONSTANTS

- **Q.25** Important features of equilibrium constants are as follows:
- id and gas state
 K_p (or NC) $K_c = \frac{[NO_2(g)]^4 [O_2(g)]}{[C u(NO_3)_2(s)]^2}$
 K_p for the reaction
 K_c is $\frac{[NO_2(g)]^4 (O_2(g)]}{[C u(NO_3)_2(s)]^2}$

(C) $K_c = [NO_2(g)]^4 (O_2(g)]$
 $\frac{[C u(NO_3)_2(s)]^2}{[C u(NO_3)_2(s)]^2}$
 $\frac{[C u(NO_3)_2(s)]^2}{[C u$ I. Expression for the equilibrium constant is not applicable only when concentrations of the reactants and products have attained constant value at equilibrium state.
	- II. The value of equilibrium constant is dependent on initial concentration of the reactants and products.
	- III. Equilibrium constant is temperature dependent having one unique value for a particular reaction represented by a balanced equation at a given temperature.
	- $p(RT)^{\Delta n}$ IV. The equilibrium constant for the reverse reaction is directly proportional to the equilibrium constant for the forward reaction.

Choose the correct statement(s) is/are

Q.26 Study the figure given below and mark the correct statement about K, and dependence of extent of reaction on it.

- (A) X-Reaction does not occur Y-Reaction proceeds to completion Z-Reaction does not occur
- (B) X-Reaction completes Y-Reaction does not occur Z-Reactants and products are at equilibrium.
- (C) X-Reaction hardly occurs Y-Reactants and products are at equilibrium. Z-Reaction proceeds to completion
- (D) X-Reaction proceeds to completion Y-Reactants and products are at equilibrium. Z-Reaction hardly occurs
- **Q.27** Consider the following graph which shows the Q.32 prediction the direction of the reaction.

- Point out the correct statement(s) for the above graph.
- (A) If $Q_c < K_c$, net reaction goes from left to right (forward reaction).
- (B) If $Q_c > K_c$ net reaction goes from right to left Q (reverse reaction)
- (C) If $Q_c = K_c$ no net reaction occurs
- (D) All of the above
- **Q.28** In the following reaction:

 $2NO(g) + Cl₂(g) \rightleftharpoons 2NOCl(g)$

it is observed that equilibrium is not attained and the rate of forward reaction is greater than rate of backward reaction. Which of the following is true for the reaction?

(A)
$$
K_p = Q_p
$$

\n(C) $Q_p < K_p$
\n(B) $Q_p > K_p$
\n(D) $Q_p = 0$

- **Q.29** The reaction quotient Q is used to
	- (A) predict the extent of a reaction on the basis of its magnitude.
	- (B) predict the direction of the reaction
	- (C) calculate equilibrium concentrations
	- (D) calculate equilibrium constant
- **Q.30** Which of the following is/are important application(s) of equilibrium constant?
	- (A) To predict the extent of a reaction on the basis of its magnitude
	- (B) To predict the direction of the reaction
	- (C) To calculate equilibrium concentrations
	- (D) All of the above

PART-7 : RELATIONSHIP BETWEEN EQUILIBRIUM CONSTANT K, REACTION QUOTIENT Q AND GIBBS ENERGY G

- **Q.31** Equilibrium constant (K) is directly related to the thermodynamics of the reaction and to the change in Gibbs energy, ΔG . If
	- I. ΔG is negative, then the reaction is spontaneous and proceeds in the forward direction.
- II. ΔG is positive, then reaction is considered non– spontaneous. Instead, as reverse reaction would have a negative ΔG , the products of the forward reaction shall be converted to the reactants. **EXERCIST AND GIBBS ENERGY G**
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	- III. ΔG is 0, reaction has achieved equilibrium; at this point, there is no longer any free energy left to drive the reaction.

Choose the correct statement(s).

- (A) I and II (B) II and III
- (C) I and III (D) I, II and III
- The correct relationship between free energy change in a reaction and the corresponding equilibrium constant, K_c is

(A)
$$
\Delta G = RT \ln K_c
$$

\n(B) $-\Delta G = RT \ln K_c$
\n(C) $\Delta G^\circ = RT \ln K_c$
\n(D) $-\Delta G^\circ = RT \ln K_c$

(C) $\Delta G^{\circ} = RT \ln K_c$ (D) $-\Delta G^{\circ} = RT \ln K_c$
 Q.33 Which of the following equation represent relationship between reaction quotient, Q, and Gibbs energy, G? (A) $\Delta G^{\Theta} = \Delta G + RT \ln Q$ (B) $\Delta G = \Delta G^{\Theta} + RT \ln Q$

(C)
$$
\Delta G = \frac{1}{\Delta G^{\Theta} + RT \ln Q}
$$
 (D) $\Delta G^{\Theta} = \frac{1}{\Delta G + RT \ln Q}$

PART-8 : FACTORS AFFECTING EQUILIBRIA

- **Q.34** Le-Chatelier's principle is applicable to all (A) physical equilibria (B) chemical equilibria (C) physical or chemical equilibria (D) physical and chemical equilibria
- **Q.35** Which of the following reactions will not get affected on increasing the pressure?
	- $(A) 2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$
	- (B) 4NH₃(g) + 5O₂(g) \rightleftharpoons 4NO(g) + 6H₂O(g)

$$
(C) CH4(g) + 2S2(g) \rightleftharpoons CS2(g) + 2H2S(g)
$$

(D)
$$
PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)
$$

Q.36 Which of the following statements is incorrect?

- (A) In equilibrium mixture of ice and water kept in perfectly insulated flask, mass of ice and water does not change with time.
- (B) The intensity of red colour increases when oxalic acid is added to a solution containing iron (III) nitrate and potassium thiocyanate.
- (C) On addition of catalyst, the equilibrium constant value is not affected.
- (D) Equilibrium constant for a reaction with negative ΔH value decreases as the temperature increases.

- **Q.37** Consider the following statements,
	- I. In general, the temperature dependence of the equilibrium constant depends on the sign of ΔH for the reaction.
	- II. The equilibrium constant for an exothermic reaction (negative ΔH) decreases as the temperature increases.
	- III. The equilibrium constant for an endothermic reaction (positive ΔH) increases as the temperature increases.
	- IV. Temperature changes affect the equilibrium Q.43 constant and rates of reactions.

Choose the correct statement(s).

- (C) I, II and III (D) I, II, III and IV
- **Q.38** On increasing the pressure, in which direction will the gas phase reaction proceed to reestablish equilibrium, is predicted by applying the Le Chatelier's principle. Consider the reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ Which of the following is correct, if the total pressure at which the equilibrium is established, is increased without changing the temperature?
	- (A) K will remain same.
	- (B) K will decrease.
	- (C) K will increase.
	- (D) K will increase initially and decrease when pressure 0.44 is very high.
- **Q.39** When I_2 dissociates to its atomic form the following reaction occurs: $I_2(g) \rightleftharpoons 2I(g)$; $\Delta_f H^\circ = + 150 \text{ kJ}$ The reaction is favoured at
	- (A) low temperature
	- (B) high temperature
	- (C) no change with temperature
	- (D) high pressure.
- **Q.40** Which of the following statement(s) is/are correct?
	- (A) Addition of hydrogen in the formation of HI at equilibrium results in the value of Q_c being less than K_c . Thus, in order to attain equilibrium again \mathbf{Q} . reaction moves in the forward direction.
	- (B) In case of manufacture of ammonia, ammonia is liquified and removed from the reaction mixture so that reaction keeps moving in forward direction.
	- (C) In the large scale production of CaO (used as important building material) from $CaCO₃$, constant removal of CO_2 from the kiln drives the reaction to $\sqrt{40}$ completion.
	- (D) All of the above
- **Q.41** The reaction $2SO_2 + O_2 \rightleftharpoons 2SO_3 + Heat$, will be favoured by
	- (A) high temperature and low pressure.
	- (B) low temperature and high pressure.
	- (C) high temperature and high pressure.
	- (D) low temperature and low pressure.
- **Q.42** Effect of a catalyst on a equilibrium reaction.
	- I. A catalyst increases the rate of the chemical reaction by making available a new low energy

pathway for the conversion of reactants to products.

- II. It increases the rate of forward and reverse reactions that pass through the same transition state and does not affect equilibrium. **EXERCUADE THEAT SAMPLE SET AND SET A EXECUTE:**
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- III. It lowers the activation energy for the forward and reverse reactions by exactly the same amount.

Which of the above statement(s) is (are) correct? (A) Only I (B) I and II

$$
(C) I, II \text{ and III} \qquad (D) II \text{ and III}
$$

Q.43 The following reaction is at equilibrium, Fe^{3+} (aq) + SCN (aq) \rightleftharpoons [Fe(SCN)]²⁺ (aq)

Yellow Colourless Deep red

$$
K_c = \frac{[Fe(SCN)]^{2+}}{[Fe^{3+}][SCN^-]}
$$

In the above reaction, colour intensity of red colour can be increased by

- (A) addition of KSCN
- (B) addition of oxalic acid which reacts with $Fe³⁺$ ions
- (C) addition of Hg^{2+} ions which react with SCN⁻ ion.
- (D) red colour intensity cannot be changed.

PART-9 : IONIC EQUILIBRIUM IN SOLUTION

- (X) classified electrolytes into strong and weak electrolytes (Y) electrolytes on dissolution in water are ionised almost completely, while the (Z) electrolytes are only partially dissociated. Correct words for X, Y, Z is $-$
	- (A) Faraday, weak, strong
	- (B) Arrhenius, weak, strong
	- (C) Arrhenius, strong, weak
	- (D) Faraday, strong, weak

PART-10 : ACIDS, BASES AND SALTS

Q.45 Sodium chloride (common salt) is an important component of our diet and is formed by the reaction between

(A) dichlorine and sodium hydride

- (B) hydrochloric acid and sodium hydride
- (C) hydrochloric acid and sodium hydroxide
- (D) dichlorine and sodium hydroxide
- **Q.46** Fill in the blanks in the given table with the appropriate choice.

Species Conjugate acid Conjugate base $HCO_3^ -p$ CO_3^2 2– $HSO₄$ –
– H_2SO_4 q $NH₃$ r S H_2O t $\rm \tilde{O}H^-$ (A) $p=H_2CO_{3,}q=SO_4^{2-}$, $r=NH_4^+$, $s=NH_2^-$, $t=H_3O^+$ (B) $p = HCO_3^{-3}$, $q = H_2SO_3$, $r = NH_2^{-3}$, $s = NH_4^{4}$, $t = H_3O^+$ (C) $p=H_2CO_3$, $q=HSO_4^-$, $r=NH_4^+, s=NH_2^-, t=H_2O$ (D) $p = HCO_3^{-3}$, $q = H_2SO_4$, $r = NH_2^{+}$, $s = NH_2^{-}$, $t = OH^{-}$

Q.47 Point out the species which acts as Lewis acids. (A) Mg^{2+} , H_2O , NH_3 ($(B) NH_3$, OH^- , Co^{3+} (C) $\text{AlCl}_3, \text{Co}^{3+}$, $, \overline{Co}^{3+}, Mg^{2+}$ (D) AlCl₃, H₂O, Co³⁺

Q.48 Acidity of BF_3 can be explained on the basis of which of the following concepts?

- (A) Arrhenius concept
- (B) Bronsted–Lowry concept
- (C) Lewis concept
- (D) Bronsted–Lowry as well as Lewis concept

Q.49
$$
\begin{array}{ccc}\n & \text{adds proton} \\
 & \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \\
 & \text{Py} & \text{Conjugate} \\
 & & \text{edge} & \text{noise} \\
 & & \text{loss} & \text{proton} \n\end{array}
$$

 (X) and (Y) in the above figure are (A) neutral, acid (B) acid, base (C) base, acid (D) base, neutral **Q.50** According to Lewis concept, an acid is a/an (A) proton donor (B) electron pair donor (C) proton acceptor (D) electron pair acceptor.

- **Q.51** "An acid is a substance that is capable of donating a hydrogen ion H^+ and bases are substances capable of accepting a hydrogen ion, H^{+} ". The above statement is justified by (A) Arrhenius concept
	- (B) Bronsted–Lowry theory
	- (C) Lewis concept
	- (D) All of these
- **Q.52** Point out the conjugate acids for the following Bronsted bases : NH_2^- , NH_3 and HCOO⁻ (A) NH_2^- , NH_4^+ and HCOOH (B) NH_3 , NH_4^+ and HCOOH (C) N^{3-} , NH^{2–} and HCOOH (D) NH_3 , NH_2^- and HCOOH

PART-11 : IONIZATION OF ACIDS AND BASES

- **Q.53** Strong acid dissociates completely in water, the resulting base formed would be very weak. The reason is that
	- (A) strong acids have strong conjugate bases.
	- (B) strong acids have strong conjugate acids.
	- (C) strong acids have very weak conjugate bases.
	- (D) strong acids have very weak conjugate acids.
- **Q.54** The pH of neutral water at 25°C is 7.0. As the temperature increases, ionisation of water increases, however, the concentration of H^+ ions and OH⁻ ions Q.6 are equal. What will be the pH of pure water at 60° C? (A) Equal to 7.0 (B) Greater than 7.0 (C) Less than 7.0 (D) Equal to zero
- **Q.55** Which of the following salts does not show its correct nature mentioned against it? (A) KBr solution – Neutral (B) NaCN solution – Acidic $(C) NH₄NO₃$ solution – Acidic
	- (D) KF solution Basic

Q.56 For a polybasic acid, the dissociation constants have a different values for each step, e.g.,

$$
H_3A \rightleftharpoons H^+ + H_2A^-; K = K_{a_1}
$$

\n
$$
H_2A^- \rightleftharpoons H^+ + HA^{2-}; K = K_{a_2}
$$

\n
$$
HA^{2-} \rightleftharpoons H^+ + A^{3-}; K = K_{a_3}
$$

DYMATERIAL: CHEMISTRY
the dissociation constants have
each step, e.g.,
 $K = K_{a_1}$
 $K = K_{a_2}$
 $K = K_{a_3}$
trend of dissociation constants
(B) $K_a = K_a$ = K_a What is the observed trend of dissociation constants in successive stages?

STUDYMATERIAL: CHEMISTRY
\nFor a polybasic acid, the dissociation constants have
\na different values for each step, e.g.,
\n
$$
H_3A \rightleftharpoons H^+ + H_2A^-
$$
; $K = K_{a_1}$
\n $H_2A^- \rightleftharpoons H^+ + HA^{2-}$; $K = K_{a_2}$
\n $HA^{2-} \rightleftharpoons H^+ + A^{3-}$; $K = K_{a_3}$
\nWhat is the observed trend of dissociation constants
\nin successive stages?
\n(A) $K_{a_1} > K_{a_2} > K_{a_3}$ (B) $K_{a_1} = K_{a_2} = K_{a_3}$
\n(C) $K_{a_1} < K_{a_2} < K_{a_3}$ (D) $K_{a_1} = K_{a_2} + K_{a_3}$
\nWhat is the value of pK_w of water?
\n(A) 7 (B) 10
\n(C) 12 (D) 14
\nMatch the Column I with Column II and choose the
\ncorrect option from the codes given below.
\nColumn I
\nColumn II
\n(Aqueous solutions) (Relativevalue of

- **Q.57** What is the value of pK_w of water? (A) 7 (B) 10
(C) 12 (D) 14 (D) 14
- **Q.58** Match the Column I with Column II and choose the correct option from the codes given below.

Q.60 The ionisation of weak base MOH can be represented by equation MOH (aq) \rightleftharpoons M⁺ (aq) + OH⁻(aq) $K_b = [M^+] [OH^-] / [MOH]$

- Where K_b is represented as
- (A) base ionisation constant
- (B) acid ionisation constant
- (C) ionic product of water
- (D) Both (A) and (B)

 (A)

- If the pH of a solution is 2, the hydrogen ion concentration in moles per litre is
	- (A) 1×10^{-14} (B) 1×10^{-2}
	- (C) 1×10^{-7} (D) 1×10^{-12}

Q.62 Point out the correct relation between
$$
K_a
$$
, K_b and K_w .
\n(A) $K_a + K_b = K_w$ (B) $K_a - K_b = K_w$

.

(C)
$$
\frac{K_a}{K_b} = K_w
$$
 (D) $K_a \times K_b = K_w$

Q.63
$$
H_2O(\ell) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)
$$

acid (X) conjugate acid (Y)
Here, (X) and (Y) refer to
(A) base and conjugate acid
(B) acid and conjugate base
(C) base and conjugate base
(D) acid and conjugate acid

Q.64 Hydronium ion concentration in molarity is more Q.73 conveniently expressed on a logarithmic scale known as the – (A) pH scale (B) pOH scale

(C) ionic product of water (D) Both (A) and (B)

Q.65 What will be the pH of a soft drink if hydrogen ion concentration in sample is 3.8×10^{-3} M?

 $(A) 3.8$ (B) 5.04

 (C) 2.42 (D) 9.2

PART-12 : BUFFER SOLUTIONS

- **Q.66** An acidic buffer solution can be prepared by mixing the solution of
	- (A) sodium acetate and acetic acid
	- (B) ammonium acetate and ammonium hydroxide
	- (C) sodium chloride and sodium hydroxide
	- (D) potassium sulphate and sulphuric acid.
- **Q.67** Which of the following will produce a buffer solution when mixed in equal volumes?
	- (A) 0.1 mol dm⁻³ NH₄OH and 0.1 mol dm⁻³ HCl
	- (B) 0.05 mol dm⁻³ NH₄OH and 0.1 mol dm⁻³ HCl
	- (C) 0.1 mol dm⁻³ NH₄OH and 0.05 mol dm⁻³ HCl
	- (D) 0.1 mol dm^{-3} CH₃COONa and 0.1 mol dm^{-3} NaOH
- **Q.68** The solutions which resist change in pH on dilution or with the addition of small amounts of acid or alkali are called –
	- (A) homogeneous solutions
	- (B) heterogeneous solutions
	- (C) ionic solutions
	- (D) buffer solutions

PART-13 : SOLUBILITY EQUILIBRIA OF SPARINGLY SOLUBLE SALTS

- **Q.69** The solubility of salts of weak acids like phosphates increases at –
	- (A) higher pOH (B) lower pOH
	- (C) higher pH (D) lower pH
- **Q.70** Solubility product expression of salt MX_4 which is sparingly soluble with a solubility 's' can be given as – (A) 256s⁵ $(B) 16s³$ (C) 5s (D) $25s^4$
- **Q.71** Match the column I with column II and mark the appropriate choice.

(A) 143.5 (B) 108 (C) 1.57×10^{-8} (D) 1.79×10^{-3}

- What is the range of solubility of slightly soluble salts? $(A) 0.001 M - 0.01 M$ (B) 0.01 M – 0.1 M $(C) 0.1 M - 1.0 M$ (D) $1.0 M - 10.0 M$
- **Q.74** For a reaction, $A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$, K_{sp} can be represented as (A) $[A^{y+}]^x$ $[B^{x-}]^y$ (B) $[A]^y$ $[B]^x$
	- (B) $[A]^x [B]^y$ $(D) [A]^{x+y} [B]^{x-y}$
- **Q.75** The values of K_{sp} of two sparingly soluble salts Ni(OH)₂ and AgCN are 2.0×10^{-15} and 6×10^{-17} respectively. Which salt is more soluble? (A) Ni(OH)_{2} is more soluble than AgCN. (B) AgCN is more soluble than $Ni(OH)_2$. (C) Both $\text{Ni}(\text{OH})_2$ and AgCN soluble to same extent. (D) Ni(OH)_2 is soluble but AgCN is insoluble. -1.0 M (D) 1.0 M – 10.0 M

tion, $A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$, K_{sp} can be

d as

d as
 $(B^{x-y}$ (B) $[A]^y [B]^x$

(D) $[A]^{x+y} [B]^{x-y}$

s of K_{sp} of two sparingly soluble salts

and AgCN are 2.0 × 10⁻¹⁵ and

sepectively. W
- Solubility product of radium sulphate is 4×10^{-11} . What will be the solubility of Ra²⁺ in 0.10 M $Na₂SO₄?$ (A) 4 × 10⁻¹⁰ M (B) 2 × 10⁻⁵ M

(C)
$$
4 \times 10^{-5}
$$
 M
(D) 2×10^{-10} M

$$
Q.77 \quad \text{BaSO}_4\left(s\right) \xrightarrow{\text{solution}} \text{Ba}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)}.
$$

The solubility product of the above reaction is (A) K $[BaSO_4] = [Ba^{2+}] [SO_4^{2-}]$ (B) $K [Ba^{2+}] = [BaSO_4]$

- (C) $K [Ba^{2+}] [SO_4^{2-}] = [BaSO_4]$
- (D) $K[SO_4^{2-}] = [BaSO_4]$
- **Q.78** The solubility product of AgCl is 1.8×10^{-10} . Precipitation of AgCl will occur by mixing which of the following solutions in equal volumes? (A) 10^{-8} M Ag⁺ and 10^{-8} M Cl⁻ ions
	- (B) 10^{-3} M Ag⁺ and 10^{-3} M Cl⁻ ions
	- (C) 10^{-6} M Ag⁺ and 10^{-6} M Cl⁻ ions
	- (D) 10^{-10} M Ag⁺ and 10^{-10} M Cl⁻ ions
- **Q.79** From Le-Chatelier's principle that if we (i) the concentration of anyone of the ions, it should combine with the ion of its __(ii)__charge and some of the salt will be precipitated till once again $K_{sp} = Q_{sp}$. Similarly, if the concentration of one of the ions is decreased, more salt will dissolve to increase the concentration of both the ions till once again (iii) . Correct words for (i) , (ii) and (iii) are $-$
	- (A) decrease, same, $K_{sp} > Q_{sp}$
	- (B) increase, opposite, $K_{sp} = Q_{sp}$
	- (C) increase, same, $K_{sp} > Q_{sp}$
	- (D) decrease, opposite, $K_{sp} > Q_{sp}$
- **Q.80** Let the solubilities of AgCl in pure H_2O , 0.01 M CaCl₂, 0.01 M NaCl and 0.05 M AgNO₃ be s_1 , s_2 , s_3 , s_4 M respectively. The correct relationship between these quantities would be

(A)
$$
s_1 > s_2 > s_3 > s_4
$$

\n(B) $s_1 > s_2 = s_3 > s_4$
\n(C) $s_1 > s_3 > s_2 > s_4$
\n(D) $s_4 > s_2 > s_3 > s_1$

spontaneous reaction or the reaction which proceeds in the forward direction to such an extent that the products are present predominantly.

(B) If $\Delta G^{\Theta} > 0$, then $-\Delta G^{\Theta}$ / RT is negative, and

$$
e^{-\Delta G^{\Theta}/RT} < 1
$$
, that is, K < 1, which implies a

non-spontaneous reaction or a reaction which proceeds in the forward direction to such a small Q.14 degree that only a very minute quantity of product is formed.

- (C) Both (A) and (B)
- (D) None of the above
- (A) pH increases and pOH decreases with rise in
- (B) pH decreases and pOH increases with rise in
- (C) both pH and pOH increases with rise in temperature.

Point out of the correct statement(s) for the above

- , products predominate over reactants, is very large, the reaction proceeds nearly
- is very small, the reaction proceeds rarely.
- , appreciable concentrations of both reactants and products are
- bar^{-1} at equilibrium. What will be the value of K_c at at $(A) 7.4 \times 10^{11}$ L mol⁻¹ $(B) 8715 \times 10^{10}$ L mol⁻¹ $(C) 0.08$ L mol⁻¹ (D) 8.715×10^{11} L mol⁻¹

(A)
$$
\left(\frac{K_{sp}}{x^y \cdot y^x}\right)^{1/x+y}
$$
 (B) $\left(\frac{K_{sp}}{x^x \cdot y^y}\right)^{x+y}$
(C) $\left(\frac{K_{sp}}{x^x \cdot y^y}\right)^{1/x+y}$ (D) $\left(\frac{K_{sp}}{x^y \cdot y^x}\right)^{x+y}$

For the reaction $a + b \rightleftharpoons c + d$, initially concentrations of a and b are equal and at equilibrium the concentration of d will be twice of that of a. What will be the equilibrium constant for the reaction?

(A) 2 (B) 9 (C) 4 (D) 3

Q.15 At 473 K, K_c for the reaction PCl₅ (g) \rightleftharpoons PCl₃(g)+Cl₂ (g) is 8.3×10^{-3} . What will be the value of K'_c for the formation of PCI_5 at the same temperature? (A) 8.3 \times 10³ (B) 120.48 $(C) 8.3 \times 10^{-3}$ (D) 240.8 **Q.16** Correct statements for Sulphuric acid : (A) It is a strong acid in the first dissociation step (is very large). (B) It is a weak acid in the second step. (C) It is a strong acid in the second step. (D) Both (A) and (B) **Q.17** Production of ammonia according to the reaction, N_2 $(g) + 3H_2(g) \rightleftharpoons 2NH_3(g);$ $\Delta H = -92.38 \text{ kJ} \text{ mol}^{-1}$ is an exothermic process. At low temperature, the reaction shifts in (A) forward direction (B) backward direction (C) either forward or backward direction (D) None of the above **Q.18** For the reaction, $H_2(g) + I_2(g) \rightleftharpoons 2HI (g);$ K_c = 57.0 at 700 K The molar concentration of $[H_2]_t = 0.10 \text{ M}, [I_2]_t = 0.20 \text{ M} \text{ and } [H_1]_t = 0.40 \text{ M}$ Calculate the reaction quotient, Q_c of the reaction. $(A) 10.0$ (B) 7.0 $(C) 8.0$ (D) 12.0 **Q.19** The value of K_c for the following equilibrium is $CaCO₃(s) \rightleftharpoons CaO(s) + CO₂(g)$ Given $K_p = 167$ bar at 1073 K. (A) 1.896 mol L^{-1} (B) $4.38 \times 10^{-3} \text{ mol L}^{-1}$ (C) 6.3×10^4 mol L⁻¹ $(D) 6.626$ mol L⁻¹ **Q.20** Formation of CIF₃ from Cl₂ and F₂ is an exothermic process. The equilibrium system can be represented as $Cl_2(g) + 3F_2(g) \rightleftharpoons 2ClF_3$; $\Delta H = -329$ kJ Which of the following will increase quantity of $CIF₃$ in the equilibrium mixture? (A) Increase in temperature, decrease in pressure, addition of Cl_2 . (B) Decrease in temperature and pressure, addition of $CIF_3.$ 4.8×10^{-9} respectively. What will be the order of their

- (C) Increase in temperature and pressure, removal of Cl_2 .
- (D) Decrease in temperature, increase in pressure, addition of $F₂$
- **Q.21** Nucleophiles are <u>while electrophiles are \equiv </u> (A) Lewis bases, Lewis acids (B) Lewis acids, Lewis bases
	- (C) Bronsted acids, Bronsted bases
	- (D) Lewis acids, Bronsted bases
- **Q.22** In which of the following reactions the increase in pressure will favour the increase in products? (A) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ (B) $\text{PCl}_3(g) + \text{Cl}_2(g) \rightleftharpoons \text{PCl}_5(g)$
	-
	- (C) $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$ (D) $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$
- **Q.23** Equimolar solutions of HF, HCOOH and HCN at 298 K have the values of K_a as 6.8×10^{-4} , 1.8×10^{-4} and

acidic strength? $(A) HF > HCN > HCOOH$ (B) $HF > HCOOH > HCN$ (C) HCN > HF > HCOOH (D) HCOOH > HCN > HF **Q.24** In which of the following solvents is silver chloride most soluble? (A) 0.1 mol dm^{-3} AgNO₃ solution (B) 0.1 mol dm^{-3} HCl solution (C) H₂O (D) Aqueous ammonia **Q.25** Which of the following salts with a concentration 0.1 M will give a basic solution? (A) Ammonium acetate (B) Ammonium chloride (C) Ammonium sulphate (D) Sodium acetate **Q.26** Consider the equilibrium set up: $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$ What will be the effect of the following on the equilibrium of the reaction? (i) Addition of H_2
(iii) Removal of CO (ii) Addition of $CH₃OH$ (iv) Removal of $CH₃OH$ (A) (i)-Forward, (ii)-Backward, (iii)-Backward, (iv)-Forward **Q.24**

In which of the blowling solvents is since choosing the blowling (A) 0.1 mol dm⁻³ AgNO₃ solution

(B) Angleous ammonia

(C) H₂O

(D) Aqueous ammonia

M will give a basic solution?

A Anmonium macette (B) Amm Which of the following salts with a concentration 0.1

M will give a basic solution?

(A) Ammonium acetate (B) Ammonium chloride

(C) Ammonium sure to:

(C) Ammonium sure to:

(C) Ammonium sure to:
 $2H_3(g) + CO(g) \rightleftharpoons CH_3OH(g$ equilibrium of the reaction?

(i) Addition of CH₃OH

(iii) Removal of CCI₃OH

(iii) Removal of CCI₃OH

(A) (i)-Forward, (iii)-Backward, (iii)-Backward,

(iv)-Forward,

(iv)-Forward, (iii)-Backward, (iii)-Forward,

(ch of the following salts with a concentration 0.1
Ill give a basic solution?

Mmmonium acetate (B) Ammonium chloride

Mmmonium sactate (B) Ammonium chloride

(g) + CO(g) \Rightarrow CH₃OH(g)

(g) + CO(g) \Rightarrow CH₃OH(g)

(g) h a concentration 0.1

Ammonium chloride

Sodium acetate

ne following on the

Addition of CH₃OH

Removal of CH₃OH

(iii)-Forward,

(iii)-Forward,

pective ionisation

pective ionisation

pective ionisation

nions:

1 (iii) Removal of CO² (iv) Removal of CH₃OH

(A) (i)-Forward, (ii)-Backward, (iii)-Backward,

(iv)-Forward, (iii)-Backward, (iii)-Forward,

(iv)-Forward, (ii)-Forward, (iii)-Forward,

(iv)-Backward, (iii)-Forward, (iii Ammonium acetate (B) Ammonium chloride

Mmonium sulphate (D) Sodium acetate

ider the equilibrium set up:

(g) + CO(g) \rightleftharpoons CH₃OH(g)

tt will be the effect of the following on the

Holdition of H₂ (ii) Addition of The following on the

Addition of CH₃OH

Removal of CH₃OH

iii)-Backward,

(iii)-Forward,

(iii)-Forward,

ii)-Forward,

pective ionisation

tions:

n K_{a₁}, K<sub>a₂ and K<sub>a₃ is

K_{a3} = K_{a1} / K<sub>a₂

K_{a3} = K_a</sub></sub></sub>

- (B) (i)-Backward, (ii)-Backward, (iii)-Forward, (iv)-Forward,
- (C) (i)-Forward, (ii)-Forward, (iii)-Backward, (iv)-Backward
- (D) (i)-Backward, (ii)-Forward, (iii)-Forward, (iv)-Backward

constants for the following reactions:

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

react:

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

(A)
$$
K_{a_3} = K_{a_1} \times K_{a_2}
$$
 (B) $K_{a_3} = K_{a_1} + K_{a_2}$

Q.28 At 500 K, equilibrium constant, K_c , for the following

ion is 5.
$$
\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons HI(g)
$$

What would be the equilibrium constant K_c for the reaction: $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)?$

- $(A) 0.04$ (B) 0.4 $(C) 25$ (D) 2.5
- **Q.29** 1 mole N_2 and 3 mol H_2 are placed in a closed container at a pressure of 4 atm. The pressure falls to 3 atm at the same temperature when the following equilibrium is attained. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. The equilibrium constant K_p for dissociation of NH₃ is – 1.5×10^{-1}
 1.5×10^{-1 \Rightarrow 2H⁺ + S²⁻

ect relationship between K_{a₁}, K_{a₂} and K_{a₃} is

= K_{a1} × K<sub>a₂ (B) K_{a3} = K_{a1} + K<sub>a₂

5 (D) K_{a3} = K_{a1} / K<sub>a₂

5 (S, equilibrium constant, K_c, for the following

is 5. $\frac{1}{2}$ $K_{a_3} = K_{a_1} \times K_{a_2}$ (B) $K_{a_3} = K_{a_1} + K_{a_2}$
 $K_{a_3} = K_{a_1} - K_{a_2}$ (D) $K_{a_3} = K_{a_1} / K_{a_2}$

200 K, equilibrium constant, K_c , for the following

ion is 5. $\frac{1}{2} H_2(g) + \frac{1}{2} I_2(g) \rightleftharpoons H1(g)$

t would be the = $K_{a_1} \times K_{a_2}$ (B) $K_{a_3} = K_{a_1} + K_{a_2}$

= $K_{a_1} - K_{a_2}$ (D) $K_{a_3} = K_{a_1} / K_{a_2}$

K, equilibrium constant, K_c , for the following

is 5. $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons H1(g)$

ould be the equilibrium constant $\frac{1}{3} = K_{a_1} + K_{a_2}$
 $= K_{a_1} / K_{a_2}$
 \therefore for the following
 $2(g) \rightleftharpoons HI(g)$

onstant K_c for the
 \therefore

a a closed container

falls to 3 atm at the

ing equilibrium is

g). The equilibrium
 $\frac{1}{3}$ is $-\frac{1}{2}$ $K_{a_3} = K_{a_1} / K_{a_2}$
 $K_{a_3} = K_{a_1} / K_{a_2}$
 K_c , for the following
 $\frac{1}{2} I_2(g) \rightleftharpoons HI(g)$

(constant K_c for the (g)?

0.4

0.5

d in a closed container

rue falls to 3 atm at the owing equilibrium is
 $I_3(g)$. Th

(A)
$$
\frac{1}{0.5} \times (1.5)^3 \text{atm}^{-2}
$$
 (B) $0.5 \times (1.5)^3 \text{atm}^2$

(C)
$$
\frac{0.5 \times (1.5)^3}{3 \times 3} \text{atm}^2
$$
 (D) $\frac{3 \times 3}{0.5 \times (1.5)^3} \text{atm}^{-2}$

129

(C) 0.145 (D) 1.45×10^{-4}

 $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$ will be

(C) 1.8×10^{-4} (D) 0.55×10^{-4}

 $(A) 1.8 \times 10^3$

Q.35 PCl₅, PCl₃ and Cl₂ are at equilibrium at 500 K in a closed container and their concentrations are 0.8×10^{-3} mol

respectively. The value of K_c for the reaction:

 L^{-1} ,1.2 × 10⁻³ mol L^{-1} and 1.2 × 10⁻³ mol L^{-1}

(B) 1.8×10^{-3}

Q.36 A 20.0 litre vessel initially contains 0.50 mole each of H_2 and I_2 gases. These substances react with finally reach an equilibrium condition. Calculate the equilibrium concentration of HI if K_{eq} = 49 for the reaction

 $Q.37$ If CO₂ be passed into water, what will remain present in solution –

 $(A) CO_2, H_2CO_3, CO_3^{2-}$, $HCO_3^ (B)$ H₂CO₃, CO₂ $(C) HCO₃⁻, CO₃²⁻$

- $(D) CO₂$, $HCO₃⁻$
- **Q.38** What will be the solubility of AgCl in 0.05 M NaCl aqueous solution if solubility product of AgCl is 1.5×10^{-10} ?

9 Consider following reactions in equilibrium with equilibrium concentration 0.01 M of every species

(I) $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$.

(II) 2HI (g)
$$
\rightleftharpoons
$$
 H₂ (g) + I₂ (g).

 $(III) N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g).$

Extent of the reactions taking place is –

Q.40 50 litre of a solution containing 10^{-5} mole of is mixed with 50 litre of a 2×10^{-7} M HBr solution. [Ag⁺] in resultant solution is [Given : $K_{\rm sn}$ (AgBr) = 5 × 10⁻¹³]

EQUILIBRIUM QUESTION BANK

EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)

NOTE : The answer to each question is a NUMERICAL VALUE.

Q.1 A sample of HI(g) is placed in flask at a pressure of 0.2 atm. At equilibrium, the partial pressure of HI(g) is 0.04 atm. What is K_p for the given equilibrium ?

 $2\text{H}(g) \rightleftharpoons H_2(g) + I_2(g)$

- **Q.2** In a gaseous system of the type, $AB(g) \rightleftharpoons A(g) + B(g)$, at a given temperature 50% of AB is dissociated at equilibrium. The value of P at equilibrium in terms of the $p.$ Find the value of x.
- **Q.3** 25 mL of H_2 and 18 mL of I_2 vapours were heated in a sealed glass tube at 465ºC and at equilibrium 30.8 mL of HI was formed. Calculate the approximate percentage degree of dissociation of HI at 465ºC. and B and As a pressure of 0.2

and the actual of the second right and K_n₂ for the priori trequilibrium, the partial pressure of HI(g) is 0.04

and the reaction: $NA \rightharpoonup A^2 \rightharpoonup H^+$; $R(H(g) \rightharpoonup H_2(g) + I_2(g)$

action HI(g) is placed in lask at a pressure of 0.2

infiltrium, the partial pressure of HI(g) is 0.04
 $H_2A \rightleftharpoons HA^- + H^+$;
 $\Rightarrow H_2(g) + I_2(g)$
 $\Rightarrow H_2(g) + I_2(g)$
 $\Rightarrow A(g) + B(g)$,
 $\Rightarrow A^2 - + H^+$;
 $\Rightarrow H_2(g) + I_2(g)$
 $\Rightarrow A(0.0)$ ml of a sol **EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)**
 EXERCISE - 3 (NUMERICAL VALUE 0.6 A dipote and H₂A undergoes the

mple of HI(g) is placed in fask at a pressure of 0.2 on A dipote acid H₂A undergoes the

At equilib The answer to each question is a NUMERICAL VALUE. 0.6 A diprotic acid H₂A undergoes the sample of HI(g) is placed in flask at a pressure of 0.2 tion reaction :

turn. At equilibrium, the particle priori equilibrium ?

t anset to calculate the approximate of H_2 and B and H_3 is given as at a pressure of 0.2 iton reaction:

At equilibrium, the partial pressure of $H_1(g)$ is 0.04 $H_2A \rightleftharpoons HA^- + H^+$; $K_1 =$

What is K_p for the gi

Q.4 For the reaction: NH₃(g)
$$
\Leftrightarrow \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)
$$
. **Q.7** A so

Degree of dissociation of $NH₃$ is given as :

$$
\alpha = \left[1 + \frac{3\sqrt{3}}{4} \frac{P}{K_p}\right]^{-1/2}, \text{ where 'P' is equilibrium } \textbf{Q.8} \text{ In all } \frac{3}{8} \text{ and } \frac{1}{8} \text{ is 3}
$$

pressure. If K_p of the above reaction is 78.1 atm at 400°C, Q K_c is $(A + 0.413)$ moles litre⁻¹. Find the value of A.

Q.5 The pressure of iodine gas at 1273 K is found to be 0.112atm whereas the expected pressure is 0.074 atm. The increased pressure is due to the dissociation of

 $I_2 \rightleftharpoons 2I$. K_p is 15.75×10^{-A} . Find the value of A.

A diprotic acid H_2A undergoes the following dissociation reaction :

$$
H_2A \rightleftharpoons HA^- + H^+;
$$
 $K_1 = 1.0 \times 10^{-7}$
\n $NA \rightleftharpoons A^{2-} + H^+;$ $K_2 = 1.0 \times 10^{-11}$

A 20.0 ml of a solution containing a mixture of Na₂A and NaHA is titrated with 0.300 M hydrochloric acid, the progress of the titration is followed with a glass electrode pH meter.

Calculate the total volume (in ml) of HCl required to reach the second equivalence point.

- $2^{2^{2} \times 2^{12}} 2^{12} \times 2^{12}$ concentration of S^{2-} is 1.44×10^{-4} M. Find the value of $\frac{3}{2}$ H₂(g). **Q.7** A solution contains 0.1 M H₂S and 0.3 M HCl. The A. Given K_{a_1} and K_{a_2} for H_2S are 10^{-7} and 1.3×10^{-13} respectively.
	- **Q.8** In above quesiton concentration of HS⁻ ions in solution is 3.3×10^{-4} M. Find the value of A.
	- **Q.9** 50 mL of 0.1 M NaOH is added to 75 mL of 0.1 M NH₄Cl to make a basic buffer. If pK_a of NH_4^+ is 9.26, Calculate the integral part of pH value.
	- **Q.10** The dissociation constants for aniline and acetic acid at 25°C are 3.83×10^{-10} M and 1.75×10^{-5} M respectively and K_w of water at 25°C is 1.008×10^{-14} M². The degree of hydrolysis of anilinium acetate in a decinormal solution is 5.5/A. Find the value of A.
	- **Q.11** Calculate the integral part of pH value of a solution that is 0.050 M in ammonium ion.

EXERCISE - 4 [PREVIOUS YEARS AIEEE / JEE MAIN QUESTIONS]

SECTION-A (CHEMICAL EQUILIBRIUM)

Q.1 For the reaction
$$
CO + \frac{1}{2}O_2 \rightarrow CO_2
$$
, the value of $\frac{K_P}{K_C}$ **Q.9**

is **[AIEEE-2002]**

(A)
$$
\frac{1}{RT}
$$
 (B) \sqrt{RT} (C) $\frac{1}{\sqrt{RT}}$ (D) RT (A)
 (B) Inc

- **Q.2** Which of the following equilibria is not affected by change in volume of the flask – **[AIEEE-2002]** $(A) PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ $(B) N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $(C) N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $(D) SO₂Cl₂(g) \rightleftharpoons SO₂(g) + Cl₂(g)$ **Q.3** For the reaction equilibrium $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ the
- concentrations of N_2O_4 and NO_2 at equilibrium are 4.8×10^{-2} and 1.2×10^{-2} mol L⁻¹ respectively. The value of K_C for the reaction is –
[AIEEE-2003] of K_C for the reaction is –
(A) 3×10^{-3} mol L⁻¹ (B) 3×10^{3} mol L⁻¹ (B) 3×10^3 mol L⁻¹ (C) 3.3 ×10² mol L⁻¹ (D) 3 ×10⁻¹ mol L⁻¹
- **Q.4** Consider the reaction equilibrium $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; $\Delta H^{\circ} = -198 \text{ kJ}$ On the basis of Le Chatelier's principle, the condition favourable for the forward reaction is - **[AIEEE-2003]** (A) Lowering the temperature and increasing the pressure.
	- (B) Any value of temperature and pressure.
	- (C) Lowering of temperature as well as pressure.
	- (D) Increasing temperature as well as pressure.

Q.5 What is the equilibrium expression for the reaction P4(S) + 5O2(g) P4O10(s) ? **[AIEEE-2004]** (A) K^C = [P4O10] [P⁴] [O²]⁵ (B) K^C = [P4O10] 5 [P⁴] [O²] (C) K^C = [O²]⁵ (D) K^C = 1 [O²]⁵

- **Q.6** For the reaction $CO_{(g)} + Cl_{2(g)} \rightleftharpoons COCl_{2(g)}$ the K_p/K_c is equ is equal to - **[AIEEE-2004]** $(A) 1/RT$ (B) RT (C) \sqrt{RT} (D) 1.0
- **Q.7** The equilibrium constant for the reaction $N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$ at temperature T is 4×10^{-4} . The value of K_C for the reaction

- **Q.8** For the reaction : $2 \text{ NO}_{2(g)} \rightleftharpoons 2 \text{ NO}_{(g)} + \text{ O}_{2(g)}$,
 $(\text{K}_{\text{c}} = 1.8 \times 10^{-6} \text{ at } 184^{\circ}\text{C}) (\text{R} = 0.0831 \text{ kJ/(mol.K)})$ When K_p and K_c are compared at 184°C it is found that (A) K_p is less than K_c **[AIEEE-2005]** (B) K_p is greater than K_c
	- (C) Whether K_p is greater than, less than or equal to K_c

depends upon the total gas pressure.

- $(D) K_p = K_c$
- $O_2 \rightarrow CO_2$, the value of $\frac{K_P}{K_C}$ **C**. The exothermic formation of CIF₃ is represented by the equation : $Cl_{2(g)} + 3F_{2(g)} \rightleftharpoons 2 ClF_{3(g)}$; $\Delta_r H = -329$ kJ (D) $K_p = K_c$
 K_P **Q.9** The exothermic formaton of ClF₃ is represented by the $\frac{1}{\sqrt{2}}$ (A) Removing Cl₂ Which of the following will increase the quantity of ClF_3 in an equilibrium mixture of Cl_2 , F_2 and ClF_3 ? **[AIEEE-2005]**
	- \overline{RT} (D) RT (B) Increasing the temperature (C) Adding F_2
		- (D) Increasing the volume of the container
		- **Q.10** A schematic plot of $\ln K_{eq}$ versus inverse of temperature
for a reaction is shown below [AIEEE-2005] for a reaction is shown below

The reaction must be

(A) endothermic

(B) exothermic

- (C) highly spontaneous at ordinary temperature
- (D) one with negligible enthalpy change
- **Q.11** An amount of solid $NH₄HS$ is placed in a flask already containing ammonia gas at a certain temperature and 0.50atm pressure. Ammonium hydrogen sulphide decomposes to yield NH_3 and H_2S gases in the flask. When the decomposition reaction reaches equilibrium, the total pressure in the flask rises to 0.84 atm ? The equilibrium constant for $NH₄HS$ decomposition at this temperature is-
[AIEEE-2005] [AIEEE-2005] $(A) 0.18$ (B) 0.30 $(C) 0.11$ (D) 0.17 cothermic

ighly spontaneous at ordinary temperature

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mount of solid NH₄HS is placed in a flask already

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[AIEEE-2005]
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P
- **Q.12** Phosphorus pentachloride dissociates as follows, is a closed reaction vessel, $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$ If total pressure at equilibrium of the reaction mixture is P and degree of dissociation of PCl_5 is x, the partial pressure of PCl₃ will be $-$ [AIEEE 2006]

(A)
$$
\left(\frac{2x}{1-x}\right)P
$$

\n(B) $\left(\frac{x}{x-1}\right)P$
\n(C) $\left(\frac{x}{1-x}\right)P$
\n(D) $\left(\frac{x}{x+1}\right)P$

Q.13 The equilibrium constant for the reaction $SO_3(g) \rightleftharpoons SO_2(g) + \frac{1}{2} O_2(g)$ is $K_c = 4.9 \times 10^{-2}$. The value of K_c for the reaction $2SO_2(g)+O_2(g) \rightleftharpoons 2SO_3(g)$ will be **[AIEEE 2006]** $(A) 2.40 \times 10^{-3}$ (B) 9.8×10^{-2} $(C) 4.9 \times 10^{-2}$ (D) 416 **Q.14** For the following three reactions a, b and c, equilibrium

constants are given - **[AIEEE 2008]** (a) $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g); K_1$ (b) $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g); K_2$ $\text{CO}_4(g) + 2\text{H}_2\text{O}(g) \rightleftharpoons \text{CO}_2(g) + 4\text{H}_2(g); K_3$ Which of the following relations is correct ?

EQUILIBRIUM QUESTION BANK

(A)
$$
K_2 K_3 = K_1
$$

\n(B) $K_3 = K_1 K_2$
\n(C) $K_3 K_2^3 = K_1^2$
\n(D) $K_1 \sqrt{K_2} = K_3$

(B) K₃ = K₁K₂ 2.22 Consider the following reversible chemical reactions :

(D) K₁ $\sqrt{K_2} = K_3$ A₂(g) + Br₂(g) = K₁ 2AB(g)(1)

1 and Kp₂for the reactions

bectively are in the ratio of

sectively are **Q.15** The equilibrium constants Kp_1 and Kp_2 for the reactions $X \rightleftharpoons 2Y$ and $Z \rightleftharpoons P + Q$, respectively are in the ratio of 1 : 9. If the degree of dissociation of X and Z be equal then the ratio of total pressures at these equilibria is - (A) 1 : 1 (B) 1 : 3 **[AIEEE 2008]** $(C) 1 : 9$

$$
(D) 1:36
$$

- **Q.16** A vessel at 1000 K contains CO_2 with a pressure of **Q.**2 0.5atm. Some of the CO_2 is converted into CO on the addition of graphite. If the total pressure at equilibrium is 0.8 atm, the value of K is $-$ [AIEEE 2011] (A) 1.8 atm (B) 3 atm (C) 0.3 atm (D) 0.18 atm Free equinoring constant $X \ne 2$ and $X \ne$ **12**
 12 (B) $K_1 = K_1$

(B) $K_2 = K_2$

(D) **E** $K_1 \sqrt{K_2} = K_3$

(D) **E** $K_1 \sqrt{K_2} = K_3$

(D) **E** $K_1 \sqrt{K_2} = K_3$
 $\sqrt{2(g) + Br_2(g)} = \frac{K_{1-2}}{2}$

(D) **E** $\sqrt{K_2} = K_3$

and $K_2 = P + Q$, respectively are in the ratio of
 Q.18 For the reaction, \log_2 and \log_2 (C) 1:9

Also A ressel at 1000 K contains CO₂ is converted into CO on the given

addition of graphite. If the total pressure at equilibrium

addition of graphite. If the total \Rightarrow P+Q, respectively are in the ratio of

Solid pressures at these diabread in the ratio of

ord pressures at these equilibria is-

(B) 1:36 (ALS_E 3008)

(B) CALS_E 3, (B) EX MAIN (JAN

(D) 1:36 (C) K₁K₂=3 (D) K₁
- **Q.17** The equilibrium constant (K_c) for the reaction $N_2(g) + O_2(g) \rightarrow 2NO(g)$ at temperature T is 4×10^{-4} . The value of K_c for the reaction [AIEEE 2012]

NO(g)
$$
\rightarrow \frac{1}{2}N_2(g) + \frac{1}{2}O_2(g)
$$
 at the same temperature is
\n(A) 0.02
\n(C) 4 × 10⁻⁴
\n(B) 2.5 × 10²
\n(A) 50.0

Q.18 For the reaction,
$$
SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g)
$$
, If

 $Kp = K_C (RT)^x$ where the symbols have ususal meaning then value of x is (assuming ideality) **[JEE MAIN 2014]** $(A) 1/2$ (B) 1 $(D) -1/2$

Q.19 The standard Gibbs energy change at 300 K for the reaction $2A \rightleftharpoons B + C$ is 2494.2 J. At a given time, the composition of the reaction mixture is $[A] = 1/2$, $[B] = 2$ and $[C] = 1/2$. The reaction proceeds in the :

[R = 8.314 J/K/mol, e = 2.718] **[JEE MAIN 2015]**

- (A) Reverse direction because $Q > K_C$
- (B) Forward direction because $Q \leq K_C$
- (C) Reverse direction because $Q \leq K_C$
- (D) Forward direction because $Q > K_C$
- **Q.20** The equilibrium constant at 298 K for a reaction $A + B \rightleftharpoons C + D$ is 100. If the initial concentration of all the four species were 1M each, then equilibrium concentration of D (in mol L^{-1}) will be [JEE MAIN 2016] (A) 0.818 (B) 1.818 (C) 1.182 (D) 0.182
- **Q.21** Which of the following lines correctly show the temperature dependence of equilibrium constant, K, for an exothermic reaction? **[JEE MAIN 2018]**

 (A) c and d (B) a and d (C) a and b (D) b and c **Q.22** Consider the following reversible chemical reactions :

$$
A_2(g) + Br_2(g) \xrightarrow{K_1} 2AB(g) \qquad \dots(1)
$$

$$
6AB(g) \xrightarrow{K_2} 3A_2(g) + 3B_2(g) \qquad \dots (2)
$$

The relation between K_1 and K_2 is :

[
$$
[\mathbf{JEE}\mathbf{MAIN}(\mathbf{JAN})\mathbf{2019}]
$$

(
$$
\mathbf{R} \times \mathbf{R} = \mathbf{K}^{-3}
$$

(A)
$$
K_2 = K_1
$$

\n(B) $K_2 = K_1^{-3}$
\n(C) $K_1 K_2 = 3$
\n(D) $K_1 K_2 = 1/3$

Q.23 For the following reactions, equilibrium constants are given $S(s) + O_2(g) \rightleftharpoons SO_2(g); K_1 = 10^{52}$

 $2S(s) + 3O_2(g) \rightleftharpoons 2SO_3(g); K_2 = 10^{129}$

The equilibrium constant for the reaction,

$$
2SO2(g) + O2(g) \Leftrightarrow 2SO3(g) is\n[JEE MAIN 2019 (APRIL)]\n(A) 10181\n(B) 10154\n(C) 1025\n(D) 1077
$$

Q.24 For the reaction, $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$, $\Delta H = -57.2 \text{kJ} \text{ mol}^{-1} \text{ and } K_c = 1.7 \times 10^{16}.$ Which of the following statement is INCORRECT?

[JEE MAIN 2019 (APRIL)]

- (A) The equilibrium constant is large suggestive of reaction going to completion and so no catalyst is required.
- (B) The equilibrium will shift in forward direction as the pressure increase.
- (C) The equilibrium constant decreases as the temperature increases.
- (D) The addition of inert gas at constant volume will not affect the equilibrium constant.

Q.25 In which one of the following equilibria,
$$
K_p \neq K_c
$$
?
\n[**JEE MANI** 2019 (APRIL)]

 $(A) NO₂(g) + SO₂(g) \rightleftharpoons NO(g) + SO₃(g)$

 (B) 2 HI $(g) \rightleftharpoons H_2(g) + I_2(g)$ (C) 2NO $(g) \rightleftharpoons N_2(g) + O_2(g)$

$$
(D) 2C(s) + O2(g) \rightleftharpoons 2CO(g)
$$

Q.26 In the figure shown below reactant A (represented by square) is in equilibrium with product B (represented by circle). $A \rightleftharpoons B$.

The equilibrium constant is : **[JEE MAIN 2020 (JAN)]**

SECTION-B (IONIC EQUILIBRIUM)

 (C) 8

Q.2 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]**

Q.3 The molar solubility (in mol L^{-1}) of a sparingly soluble salt MX_4 is 's'. The corresponding solubility product is K_{SP} . 's' is given in terms of K_{SP} by the relation -

 [AIEEE-2004]

WANGEDILARINING	STUDYM		
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]	Cauchy B	Cauchy S	Q.10 In aqueous solution the ioniza acid are K ₁ = 4.2 × 10 ⁻⁷ and K S
(A) 1×10^{-15}	(B) 1×10^{-10}	(A) The concentration of CO ₂ S	
The molar solubility (in mol L ⁻¹) of a sparingly soluble salt MX ₄ is 's'. The corresponding solubility product is K _{SP} . 's' is given in terms of K _{sp} by the relation [A] 1×10^{-19}	(B) The concentration of CO ₂ of H ⁺ & of the equal.		
(A) $s = (K_{SP} / 128)^{\frac{1}{4}}$	(B) $s = (128K_{SP})^{\frac{1}{4}}$	Q.11 Solubility product of silver br quantity of potassium bromide of mol ⁻¹) to be added to 1 li silver nitrate to start the preci and do 1 li and X ₂ , in water is 4 × 10 ⁻¹² . The concentration of M ²⁺ ions in the aqueous solution of the salt is [AIEEE-2005]	(A) 1.2×10^{-10} g (A) 1.0×10^{-4} M B) 2.0×10^{-6} M C) 1.6×10^{-4} M Q.12 At 25°C, the solubility product of a sublity product of a

Q.4 The solubility product of a salt having 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]**

Q.5 The first and second dissociation constants of an acid H₂A are 1.0×10^{-5} & 5.0×10^{-10} respectively. The overall dissociation constant of the acid will be -**[AIEEE-2007]** (A) 5.0 × 10⁻⁵ (B) 5.0 × 10¹⁵

(C) 5.0 × 10–15 (D) 0.2 × 10⁵

Q.6 The pK_a of a weak acid (HA) is 4.5. The pOH of an aqueous buffered solution of HA in which 50% of the acid is ionized is- **[AIEEE-2007]**

Q.7 In a saturated solution of the sparingly soluble strong electrolyte AgIO₃ (molecular mass = 283) the equilibrium (C) which sets in is : AgIO_{3(s)} \Longleftrightarrow Ag⁺_(aq) + IO₃⁻_(aq) If the solubility product K_{sp} of AgIO₃ at a given temperature is 1.0×10^{-8} , what is the mass of AgIO₃ contained in 100 ml of its saturated solution ?

[AIEEE-2007]

Q.8 The p K_a of a weak acid, HA, is 4.80. The p K_b of a weak base, BOH, is 4.78. The pH of an aqueous solution of the corresponding salt, BA, will be - **[AIEEE-2008]**

Q.9 Solid Ba(NO₃) is gradually dissolved in a 1.0×10^{-4} M solution Na_2CO_3 solution. At what concentration of Ba²⁺ will a precipitate begin to form ? (K_{sp} for Ba CO₃ = 5.1 \times 10⁻⁹)

[AIEEE-2009]

 $Q.10$ **Q.10** In aqueous solution the ionization constants for carbonic acid are K₁ = 4.2 \times 10⁻⁷ and K₂ = 4.8 \times 10⁻¹¹. .

Select the correct statement for a saturated 0.034 M solution of the carbonic acid. **[AIEEE 2010]**

- (A) The concentration of CO_3^2 ⁻ is 0.034 M.
- (B) The concentration of CO_3^2 ⁻ is greater than that of HCO_3^-
- (C) The concentration of $H^+ \& HCO_3^-$ are approximately equal.
- 1 (D) The concentration of H^+ is double that of CO_3^2 ⁻
- $128K_{SP}$ $\overline{)4}$ Q.11 Solubility product of silver bromide is 5.0×10^{-13} . The $\frac{1}{5}$ of mol⁻¹) to be added to 1 litre of 0.05 M solution of **CUESTION BANK**

STUDYMATERIAL: CHEMISTRY

ulube salt AB₂ is **Q.10** In aqueous solution the ionization constants for carbonic

int will be $-$

[AIEEE-2003] Select the correct statement for a saturated 0.034 M

solution quantity of potassium bromide (molar mass taken as 120g silver nitrate to start the precipitation of AgBr is

[AIEEE 2010]

(A)
$$
1.2 \times 10^{-10} \text{ g}
$$

\n(B) $1.2 \times 10^{-9} \text{ g}$
\n(C) $6.2 \times 10^{-5} \text{ g}$
\n(D) $5.0 \times 10^{-8} \text{ g}$

- **Q.12** At 25^oC, the solubility product of $Mg(OH)_2$ is
- 1.0×10^{-11} . At which pH, will Mg²⁺ ions start precipitating in the form of $Mg(OH)_2$ from a solution of 0.001 M Mg2+ ions ? **[AIEEE 2010]** $(A) 9$ (B) 10 $(C) 11$ (D) 8 (D) The concentration of H¹ is double that of CQ_3 ⁻⁻
Solubility product of silver bromide is 5.0×10^{-13} . The
quantity of potassium bromide (molar mass taken as 120g
of mol⁻¹)</sup> to be added to 1 litre of 0.05 M Fine concentration of H⁺ is double that of CO_3 ²-

The concentration of H⁺ is double that of CO_3 ²-

bility of potassium bromide (molar mass taken as 120g

holdron of nitrate to start the precipitation of AgBr i ncentration of H⁺ is double that of CO₃²⁻

roduct of silver bromide is 5.0 × 10⁻¹³. The

botassium bromide (molar mass taken as 120g

be added to 1 litre of 0.05 M solution of

to start the precipitation of AgBr i Is double that of CO₃⁻⁻

omide is 5.0 × 10⁻¹³. The

(molar mass taken as 120g

tre of 0.05 M solution of

pitation of AgBr is

[AIEEE 2010]

(B) 1.2 × 10⁻⁹ g

(D) 5.0 × 10⁻⁸ g

to f Mg(OH)₂ is

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of 0.05 M solution of
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[AIEEE 2010]
1.2 × 10⁻⁹ g
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[AIEEE hat of CO_3^{2-}
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AgBr is

[AIEEE 2010]
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 0^{-8} g
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 2^{2+} ions start

pm a solution of

[AIEEE 2010]

eak electrolyte,

the expression: of mol⁻¹) to be added to 1 litre of 0.05 M solution of
silver nitrate to start the precipitation of AgBr is
(A) 1.2×10^{-10} g (B) 1.2×10^{-9} g (D) 5.0×10^{-8} g (D) 5.0×10^{-8} g (D) 5.0×10^{-8} g (D) $5.0 \times$ te to start the precipitation of AgBr is

[AIEEE 2010]
 0^{-10} g (B) 1.2×10^{-9} g
 $(D) 5.0 \times 10^{-8}$ g

en solubility product of Mg(OH)₂ is
 $e^{1.1}$. At which pH, will Mg²⁺ ions start

rang in the form of Mg(OH)₂ potassium bromide (molar mass taken as 120g

potassium bromide (molar mass taken as 120g

to be added to 1 litre of 0.05 M solution of

te to start the precipitation of AgBr is

[AIEEE 2010]
 0^{-10} g (B) 1.2×10^{-9} g nity of potassium bromide (molar mass taken as 120g

not⁻¹) to be added to 1 litre of 0.05 M solution of

er nitrate to start the precipitation of AgBr is
 1.2×10^{-10} g (B) 1.2×10^{-9} g
 6.2×10^{-5} g (D) $5.0 \$ tre of 0.05 M solution of
pitation of AgBr is

[AIEEE 2010]

(B) 1.2×10^{-9} g

(D) 5.0×10^{-8} g

et of Mg(OH)₂ is

I, will Mg²⁺ ions start

g(OH)₂ from a solution of

[AIEEE 2010]

(B) 10

(D) 8

(a) of a weak AgBr is

[AIEEE 2010]
 0^{-9} g
 0^{-8} g
 H)₂ is
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[AIEEE 2010]

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the expression:

[AIEEE 2011]
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 $\frac{1}{i-1}$

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e acid 10×10^{-1} . The
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[AIEEE 2010]
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[AIEEE 2010]
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[AIEEE 2011]
 $\frac{i-1}{i-1}$
i $+y+1$
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tion of AgBr is

[AIEEE 2010]

1.2 × 10⁻⁹ g

5.0 × 10⁻⁸ g

[Mg(OH)₂ is

ill Mg²⁺ ions start

H₂ from a solution of

[AIEEE 2010]

10

8

f a weak electrolyte,

(i)
- **Q.13** The degree of dissociation (α) of a weak electrolyte, A_xB_y is related to van't Hoff factor (i) by the expression:

[AIEEE 2011]

(A)
$$
\alpha = \frac{i-1}{(x+y-1)}
$$

\n(B) $\alpha = \frac{i-1}{x+y+1}$
\n(C) $\alpha = \frac{x+y-1}{i-1}$
\n(D) $\alpha = \frac{x+y+1}{i-1}$

Q.14 The pH of a 0.1 molar solution of the acid HQ is 3. The value of the ionization constant, K_a of the acid is :

[AIEEE 2012]

Q.15 How many litres of water must be added to 1 litre an aqueous solution of HCl with a pH of 1 to create an aqueous solution with pH of 2 ? **[JEE MAIN 2013]** $(A) 0.1 L$ (B) 0.9 L

$$
(C) 2.0 L \t\t (D) 9.0 L
$$

- **Q.16** pK_a of 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 2017]** $(A) 1.0$ (B) 7.2 $(C) 6.9$ (D) 7.0
- **Q.17** Which of the following salts is the most basic in aqueous solution? **[JEE MAIN 2018]** (A) FeCl₃ $(B) Pb(CH_3COO)$ (C) Al (CN) ₃ (D) CH₃COOK

Q.18 An alkali is titrated against an acid with methyl orange as indicator, which of the following is a correct combination? **[JEE MAIN 2018]**

Q.19 An aqueous solution contains an unknown concentration of Ba^{2+} . When 50 mL of a 1 M solution of $Na₂SO₄$ is added, BaSO₄ just begins to precipitate. The Q.25 final volume is 500 mL. The solubility product of $BaSO₄$ is 1×10^{-10} . What is the original concentration of Ba²⁺

[JEE MAIN 2018]

Q.20 An aqueous solution contains $0.10 M H₂S$ and $0.20M$ HCl. If the equilibrium constants for the formation of HS⁻ from H₂S is 1.0×10^{-7} and that S²⁻ from HS⁻ ions is 1.2×10^{-13} then the concentration of S^{2–} ions in aqueous solution is : **[JEE MAIN 2018]** blance is Solomic. The solohibity product of IBSO on L 200 min. 2010

10-10. What is the original concentration of Ba² solution is now taken in another container.

16-10. What is the original concentration of Ba² solu 1 × 1 0⁻² M

x 10⁻⁹ M

(D) 1.0 × 10⁻⁹ M

(D) 1.0 × 10⁻⁹ M

(D) 2 × 10⁻⁹ M

Fig. (S)

Q.21 20 mL of 0.1 MH₂SO₄ solution is added to 30 mL of 0.2M $_{\text{Q,27}}$ The solution $NH₄OH$ solution. The pH of the resulatant mixture is :

Q.22 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

[JEE MAIN 2019 (APRIL)]

(A)
$$
S = \left(\frac{K_{sp}}{929}\right)^{1/9}
$$
 (B) $S = \left(\frac{K_{sp}}{216}\right)^{1/7}$

(C)
$$
S = \left(\frac{K_{sp}}{144}\right)^{1/6}
$$
 (D) $S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$

Q.23 The pH of a 0.02M NH₄Cl solution will be [Given $K_b(NH_4OH) = 10^{-5}$ and $log 2 = 0.301$]

[JEE MAIN 2019 (APRIL)]

- **Q.24** Two solutions A and B, each of 100 L was made by dissolving 4g of NaOH and 9.8 g of H_2SO_4 in water, respectively. The pH of the resultant solution obtained from mixing 40 L of solution A and 10 L of solution B is________. **[JEE MAIN 2020 (JAN)**]
- **Q.25** 3 gram of acetic acid is mixed in 250 mL of 0.1 M HCl. This mixture is now diluted to 500 mL. 20mL of this solution is now taken in another container.

(1/2) mL of 5 M NaOH is added to this. Find the pH of this solution. ($\log 3 = 0.4771$, $pK_a = 4.74$)

[JEE MAIN 2020 (JAN)]

Q.26 The K_{sp} for the following dissociation is 1.6×10^{-5}

$$
PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)
$$

Which of the following choices is correct for a mixture of 300 mL 0.134 M $Pb(NO₃)₂$ and 100mL 0.4 M NaCl ?

[JEE MAIN 2020 (JAN)]

Q.27 The solubility product 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 2020 (JAN)]

EXERCISE - 5 (PREVIOUS YEARS AIPMT/NEET EXAM QUESTIONS)

Choose one correct response for each question.

- **Q.1** What is the correct relationship between the pHs of isomolar solutions of sodium oxide (pH_1) , sodium (A) sulphide (pH_2) , sodium selenide (pH_3) and sodium (C) telluride (pH_4)?) ? **[AIPMT 2005]** (A) $pH_1 > pH_2 > pH_3 > pH_4$ (B) $p H_1 > p H_2 \approx p H_3 > p H_4$ (C) $pH_1 < pH_2 < pH_3 < pH_4$
	- (D) $pH_1 < pH_2 < pH_3 \approx pH_4$
- **Q.2** At 25°C, the dissociation constant of a base, BOH, is 1.0×10^{-12} . The concentration of hydroxyl ions in 0.01M aqueous solution of the base would be –**[AIPMT 2005]** (A) 1.0×10^{-5} mol L⁻¹ (B) 1.0×10^{-6} mol L⁻¹ (C) 2.0×10^{-6} mol L⁻¹ (D) 1.0×10^{-7} mol L⁻¹
- **Q.3** H₂S gas when passed through a solution of cations containing HCl precipitates the cations of second group of qualitative analysis but not those belonging to the fourth group. It is because **[AIPMT 2005]** (A) Presence of HCl decreases the sulphide ion concentration. (A) Pivape Point and the Real Monte of the Columbia (A) 100 (A) 2000

(A) L(1) The Concentration of the base would be $[-4\text{IPMT} 2005]$

(A) L(3) II and K₂ ii and C₁ and 10 x (i) π (i) π
	- (B) Sulphides of group IV cations are unstable in HCl.
	- (C) Solubility product of group II sulphides is more than that of group IV sulphides.
	- (D) Presence of HCl increases the sulphides ion concentration.
- **Q.4** Equilibrium constants K_1 and K_2 for the following **Q.11** equilibria : **[AIPMT 2005]**

(i) $NO(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2(g), K_1$

(ii) $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g), K_2$

The equilibrium constants K_1 and K_2 are related as –

(A)
$$
K_2 = 1/K_1
$$

\n(B) $K_2 = K_1^{1/2}$
\n(C) $K_2 = 1/K_1^2$
\n(D) $K_2 = K_1^2$

- **Q.5** For the reaction : $CH_4(g)+2O_2(g) \rightleftharpoons CO_2(g)+2H_2O(\ell)$ $\Delta_r H = -170.8 \text{ kJ} \text{ mol}^{-1}$ [AIPMT 2006] Which of the following statement is not true – stants K₁ and K₂ for the following **Q.11** The value of equilibity

[AIPMT 2005]

(g) \Rightarrow NO₂(g), K₁ HI(g) \Rightarrow $\frac{1}{2}$ H₂(g) \Rightarrow $\frac{1}{2}$ H₂(g) \Rightarrow (A) 1/8

constants K₁ and K₂ are related as -

(B) K
	- (A) The equilibrium constant for the reaction is given

by
$$
K_p = \frac{[CO_2]}{[CH_4][O_2]}
$$
.

- (B) Addition of CH₄(g) or O_2 (g) at equilibrium will cause a shift to the right.
- (C) The reaction is exothermic.
- (D) At equilibrium, the conentration of $CO₂(g)$ and $H₂O(\ell)$ are not equal.
- **Q.6** The hydrogen ion concentrated of a 10–8 M HCl aqueous solution at 298 K ($K_w = 10^{-14}$) is [AIPMT 2006] Q. (A) 11×10^{-8} M (B) 9.525×10^{-8} M
(C) 1.0×10^{-8} M (D) 1.0×10^{-6} M (D) 1.0×10^{-6} M
- **Q.7** Which of the following pairs constitutes a buffer
- **[AIPMT 2006]** (A) NaOH and NaCl (B) HNO₃ and NH₄NO₃ (C) HCl and KCl (D) HNO₂ and NaNO₂ **Q.8** A weak acid, HA, has a K_a of 1.00×10^{-5} if 0.100 mole of this acid dissolved in one litre of water, the percentage of acid dissociated at equilibrium is closest to – (A) 1.00% (B) 99.9% **[AIPMT 2007]** $(C) 0.100\%$ (D) 99.0% **ADEA ASSAURE 14 (B)** HNO₃ and NH₄NO₃

Cl and KCl (B) HNO₂ and NH₄NO₃

Cl and KCl (D) HNO₂ and NaNO₂

kacid, HA, has a K_a of 1.00 × 10⁻⁵ if 0.100 mole of

id dissolved in one litre of water, the percen **STUDYMATERIAL: CHEMISTRY**
 EET EXAM QUESTIONS
 EET EXAM QUESTIONS

An OH and NaCl
(B) HNO₃ and NH₄NO₃

HCl and KCl
(D) HNO₂ and NaNO₂

HCl and KCl
(D) HNO₂ and NaNO₂

ack acid, HA, has a K_a of 1.00 × [AIPMT 2006]
NO₃ and NH₄NO₃
NO₂ and NaNO₂
 10^{-5} if 0.100 mole of
ater, the percentage
s closest to –
9% [AIPMT 2007]
.0%
s are given
xidation of NH₃ by
[AIPMT 2007]
 $\frac{2K_3}{K_1}$ **EXEAL: CHEMISTRY**
 NS)

stitutes a buffer

[AIPMT 2006]

HNO₃ and NH₄NO₃

HNO₂ and NaNO₂

× 10⁻⁵ if 0.100 mole of

water, the percentage

is closest to –

99.9% [AIPMT 2007]

99.0%

mts are given

oxidation k acid, HA, has a K_a of 1.00×10^{-5} if 0.100 mole of

cid dissolved in one litre of water, the percentage

d dissociated at equilibrium is closest to –

00% (B) 99.9% [AIPMT 2007]

00% (D) 99.0%

100% (D) 99.0%

11-1 ch of the following pairs constitutes a buffer

[AIPMT 2006]

HCl and KCl

HCl and KCl

(B) HNO₂ and NH₄NO₃

HCl and KCl

(B) HNO₂ and NaN₀

card acid dissolved in one litre of vater, the percentage

cid dissoci 0⁻⁵ if 0.100 mole of

ater, the percentage

closest to -

9% [AIPMT 2007]

0%

s are given

idation of NH₃ by

[AIPMT 2007]
 $\frac{2}{5}K_3$
 K_1
 $2K_3^3$
 K_1

5°C that contains

30⁺.[AIPMT 2007] stitutes a buffer

[AIPMT 2006]

HNO₃ and NH₄NO₃

HNO₂ and NaNO₂
 $\times 10^{-5}$ if 0.100 mole of

water, the percentage

is closest to –

9.9% [AIPMT 2007]

99.0%

mts are given

oxidation of NH₃ by

[AIPMT 2007]
	- The following equilibrium constants are given

 $N_2 + 3H_2 \rightleftharpoons 2NH_3$; K₁ $N_2 + O_2 \rightleftharpoons 2NO$; K_2 $H_2 + \frac{1}{2}O_2 \rightleftharpoons H_2O$; K₃

The equilibrium constant for the oxidation of $NH₃$ by oxygen to give NO is – **[AIPMT 2007]**

(A)
$$
\frac{K_2 K_3^2}{K_1}
$$
 (B) $\frac{K_2^2 K_3}{K_1}$

(C)
$$
\frac{K_1 K_2}{K_3}
$$
 (D) $\frac{K_2 K_3^3}{K_1}$

- **Q.10** Calculate the pOH of a solution at 25^oC that contains 1×10^{-10} M of hydronium ions, i.e., H₃O⁺.[AIPMT 2007]
(A) 4.000 (B) 9.0000 (B) 9.0000 (C) 1.000 (D) 7.000 (D) 99.0%

equilibrium constants are given
 $2NH_3$; K₁

NO ; K₂

H₂O ; K₃

m constant for the oxidation of NH₃ by

[AIPMT 2007]

(B) $\frac{K_2^2 K_3}{K_1}$

(D) $\frac{K_2 K_3^2}{K_1}$

(D) $\frac{K_2 K_3^2}{K_1}$

pOH of a s (D) 99.0%

n constants are given

for the oxidation of NH₃ by

[AIPMT 2007]

(B) $\frac{K_2^2 K_3}{K_1}$

(D) $\frac{K_2 K_3^3}{K_1}$

lution at 25°C that contains

ons, i.e., H₃O⁺.[AIPMT 2007]

(B) 9.0000

(D) 7.000

constan
- **Q.11** The value of equilibrium constant of the reaction

$$
HI(g) \xrightarrow{\ } \frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \text{ is 8.0.}
$$

The equilibrium constant of the reaction

 $H_2(g) + I_2(g) \xrightarrow{\longrightarrow} 2HI(g)$ will be: [AIPMT 2008] (A) 1/8 (B) 1/16 (C) 1/64 (D) 16

- **Q.12** Equal volumes of three acid solutions of pH 3, 4 and 5 are mixed in a vessel. What will be the H^+ ion concentration in the mixture ? **[AIPMT 2008]** (A) 1.11×10^{-3} M (B) 1.11×10^{-4} M
- (C) 3.7×10^{-4} M (D) 3.7×10^{-3} M
- **Q.13** The values of Kp_1 and Kp_2 for the reactions $X \longrightarrow Y + Z$ (i) and

ordination of group IV cations are unstable in HCl.

(C) K₃

(D) T

ordination of group II sulphides is more than

(D) T sulphides in the pOH of a solution at 2

ordination of HCl increases the sulphides ion

(A) 44.000 **EVALUATE SET AND THE SURVEY ON A CONSTANT AND CONSTANT (B) 11.11

SUP IN SUPRIMEND (B) 2008 (D) 7.000 (D) 7.000 (D) 7.000** $A \rightleftharpoons 2B$ (ii) are in ratio of 9:1. If degree of dissociation of X and A be equal, then total pressure at equilibrium (i) and (ii) are in the ratio : **[AIPMT 2008]** $(A) 1 : 1$ (B) 3 : 1 $(C) 1 : 9$ (D) 36 : 1

> Q.14 If the concentration of OH⁻ ions in the reaction Fe $(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$ is decreased by $1/4$ times, then equilibrium concentration of Fe^{3+} will increase by: **[AIPMT 2008]** (A) 4 times (B) 8 times (C) 16 times (D) 64 times

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 $(C) NH₃$

 (D) $BF₃$

initial value of 1.00 M. When equilibrium is reached, the concentration of D is measured and found to be 0.25 M. The value for the equilibrium constant for this reaction

Q.30 Equimolar solutions of the following substances were prepared separately. Which one of these will record the highest pH value ? **[AIPMT (PRE) 2012]** (A) $BaCl₂$
(C) LiCl (B) AlCl₃

(D) BeCl₂

- **Q.31** Buffer solutions have constant acidity and alkalinity because : **[AIPMT (PRE) 2012]**
	- (A) these give unionised acid or base on reaction with added acid or alkali.
	- (B) acids and alkalies in these solution are shielded from attack by other ions.
	- (C) they have large excess of H^+ or OH⁻ ions
	- (D) they have fixed value of pH.
- **Q.32** Given that the equilibrium constant for the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ has a value of 278 at a particular temperature. What is the value of the equilibrium constant for the following reaction at the same temperature? $SO_3(g) \rightleftharpoons SO_2(g) + \frac{1}{2}O_2(g)$

[AIPMT (MAINS) 2012] (A) 1.8×10^{-3} (B) 3.6×10^{-3}

 $(C) 6.0 \times 10^{-2}$ (D) 1.3×10^{-5}

Q.33 Given the reaction between 2 gases represented by A_2 and B_2 to give the compound AB (g).

 $A_2(g) + B_2(g) \rightleftharpoons 2 AB(g).$

At equilibrium, the concentration

of $A_2 = 3.0 \times 10^{-3}$ M of $\overline{B_2}$ = 4.2 × 10⁻³ M

of $\overline{AB} = 2.8 \times 10^{-3}$ M. lf the reaction takes place in a sealed vessel at 527°C, then the value of K_C will be : **[AIPMT (MAINS) 2012]** $(A) 2.0$ (B) 1.9

- $(C) 0.62$ (D) 4.5 **Q.34** Which of these is least likely to act as a Lewis base? (A) PF₃ (B) CO **[NEET 2013]** $(C) F^ (D)$ BF₃
- **Q.35** Which of the following salts will give highest pH in water? **[AIPMT 2014]** (A) KCl (B) NaCl (C) Na₂CO₃ (D) CuSO_{$₄$}</sub>

Q.36 For the reversible reaction,

 $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + Heat$ The equilibrium shifts in forward direction

[AIPMT 2014]

- (A) By increasing the concentration of $NH₃(g)$
- (B) By decreasing the pressure
- (C) By decreasing the concentrations of $N_2(g)$ and $H₂(g)$
- (D) By increasing pressure and decreasing temperature.
- **Q.37** For a given exothermic reaction, K_p and K'_p are the equilibrium constants at temperatures T_1 and T_2 , respectively ($T_2 > T_1$). Assuming that heat of reaction is constant in temperature range between T_1 and T_2 , it is readily observed that – **[AIPMT 2014]**

(A)
$$
K_p > K'_p
$$

\n(C) $K_p = K'_p$
\n(B) $K_p < K'_p$
\n(D) $K_p = 1/K'_p$

- **Q.38** If the value of an equilibrium constant for a particular reaction is 1.6×10^{12} , then at equilibrium the system will contain? **[AIPMT 2015]**
	- (A) mostly reactant.
	- (B) mostly products
	- (C) similar amounts of reactants and products.

(D) all reactants.

- **Q.39** The K_{sp} of Ag_2CrO_4 , AgCl, AgBr andAgI are respectively, 1.1×10^{-12} , 1.8×10^{-10} , 5.0×10^{-13} , 8.3×10^{-17} . Which one of the following salts will precipitate last if $AgNO_3$ solution is added to the solution containing equal moles of NaCl, NaBr, NaI and $Na_2CrO₄$? ? **[AIPMT 2015]** (A) AgCl (B) AgBr
	- (C) Ag₂CrO₄ (D) AgI
- **Q.40** If the equilibrium constant for $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is K, the equilibrium constant for $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g)$ will be

for
$$
\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g)
$$
 with be
\n(A) K
\n(B) K² [RE-AIPMT 2015]
\n(C) K^{1/2} (D) K/2

- **Q.41** Which one of the following pairs of solution is not an acidic buffer ? **[RE-AIPMT 2015]** (A) H_2CO_3 and Na_2CO_3 (B) H_3PO_4 and Na_3PO_4 (C) HClO₄ and NaClO₄ (D) CH₃COOH & CH₃COONa **Q.42** Aqueous solution of which of the following compounds
- is the best conductor of electric current? **[RE-AIPMT 2015]**
	- (A) Ammonia, NH₃

(B) Fructose, $C_6H_{12}O_6$

-
- (C) Acetic acid, $C_2H_4O_2$
- (D) Hydrochloric acid, HCl
- **Q.43** What is the pH of the resulting solution when equal volumes of 0.1 M NaOH and 0.01 M HCl are mixed ?

[RE-AIPMT 2015]

- $(A) 7.0$ (B) 1.04 (C) 12.65 (D) 2.0
- **Q.44** MY and NY_3 , two nearly insoluble salts, have the same K_{SP} values of 6.2 \times 10⁻¹³ at room temperature. Which statement would be true in regard to MY and NY_3 ?

[NEET 2016 PHASE 1]

- (A) The molar solubility of MY $&$ NY₃ in water are identical.
- (B) The molar solubility of MY in water is less than that of NY_3 .
- (C) The salts MY and NY₃ are more soluble in 0.5 M KY than in pure water.
- (D) The addition of the salt of KY to solution of MY and NY_3 will have no effect on their solubilities.

EQUILIBRIUM QUESTION BANK

Q.50 Which of the following statements is not correct?

[NEET 2017]

- (A) The value of equilibrium constant is changed in the presence of a catalyst in the reaction at equilibrium.
- (B) Enzymes catalyse mainly bio-chemical reactions.
- (C) Coenzymes increase the catalytic activity of enzyme.
- (D) Catalyst does not initiate any reaction.
- **Q.51** A 20 litre container at 400 K contains $CO₂(g)$ at pressure 0.4 atm and an excess of SrO (neglect the volume of solid SrO). The volume of the container is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of $\rm CO_2$ attains its maximum value, will be (Given that :

 $SrCO₃(s) \rightleftharpoons SrO(s) + CO₂(g), K_p = 1.6atm)[NEET 2017]$ (A) 10 litre (B) 4 litre (C) 2 litre (D) 5 litre

Q.52 Following solutions were prepared by mixing different volumes of NaOH and HCl of different concentrations

(a) 60 mL
$$
\frac{M}{10}
$$
 HCl + 40 mL $\frac{M}{10}$ NaOH [NEET 2018]

(b) 55 mL
$$
\frac{M}{10}
$$
 HCl + 45 mL $\frac{M}{10}$ NaOH

$$
\frac{\text{m r}}{\text{d}T} = \frac{20 \text{m}}{\text{R} \text{T}^2}
$$
 (c) 75 mL $\frac{\text{M}}{10}$ HCl + 25 mL $\frac{\text{M}}{5}$ NaOH

(d) 100 mL
$$
\frac{M}{10}
$$
 HC1+100 mL $\frac{M}{10}$ NaOH

pH of which one of them will be equal to 1? $(A) d$ (B) a

 $(C) b$ (D) c **Q.53** The solubility of $BaSO₄$ in water is

 2.42×10^{-3} gL⁻¹ at 298 K. The value of its solubility product $(K_{\rm SD})$ will be (Given molar mass of BaSO₄= 233 g mol⁻¹) [NEET 2018] (A) 1.08×10^{-14} mol²L⁻² (B) 1.08×10^{-12} mol²L⁻²

(C)
$$
1.08 \times 10^{-10}
$$
 mol²L⁻² (D) 1.08×10^{-8} mol²L⁻²

Q.54 Which one of the following conditions will favour maximum formation of the product in the reaction, $A_2(g) + B_2(g) \Longrightarrow X_2(g) \Delta_r H = -X kJ$? [NEET 2018] (A) High temperature and high pressure (B) Low temperature and low pressure

-
- (D) High temperature and low pressure
- **Q.55** pH of a saturated solution of $Ca(OH)_2$ is 9. The solubility product (K_{sp}) of Ca(OH)₂ is:
(A) 0.5×10^{-15} is: **[NEET 2019]** $(B) 0.25 \times 10^{-10}$ $(C) 0.125 \times 10^{-15}$ (D) 0.5×10^{-10}
- **Q.56** Conjugate base for Bronsted acids H_2O and HF are :
(A) OH⁻ and H_2F^+ , respectively. [NEET 2019] (A) OH⁻ and H₂F⁺, respectively. **[NEET 2019]** (B) H_3O^+ and F⁻, respectively. (C) OH⁻ and F⁻, respectively. (D) H_3O^+ and H_2F^+ , respectively.

Q.57 Which will make basic buffer? **[NEET 2019]**

- (A) 50 mL of 0.1 M NaOH + 25 mL of 0.1 M CH₃COOH
- (B) $100 \text{ mL of } 0.1 \text{ M CH}_3 \text{COOH} + 100 \text{ mL of } 0.1 \text{ M NaOH}$
- (C) $100 \text{ mL of } 0.1 \text{ M }$ HCl + 200 mL of 0.1 M NH₄OH.
- (D) 100 mL of 0.1 M HCl + 100 mL of 0.1 M NaOH

ANSWER KEY

CHEMICAL EQUILIBRIUM TRY IT YOURSELF-1

(1) $K_p = K_c (RT)^{\Delta n}$ $K_p = (6.02 \times 10^{-2} \text{ L}^{-2} \text{ mol}^{-2}) \times (0.082 \text{ L atm K}^{-1} \text{ mol}^{-1})$ \times 773 K)⁻²

.

$$
= 1.5 \times 10^{-5} \,\text{atm}^{-2}.
$$

- **(2)** (D)
- **(3)** The reaction is reversed and then multiplied by 1/2, therefore relation between K_c and K_c' will be

$$
K'_{c} = \sqrt{\frac{1}{K_{c}}} = \sqrt{\frac{1}{2.5 \times 10^{-5}}} = \frac{1}{5 \times 10^{-3}} = 2 \times 10^{2}
$$

(4) Given, $[SO_2] = [SO_3]$, let O_2 is x mole at equilibrium $V = 5$ litre, $K_c = 10$

$$
K_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}
$$

\n
$$
K_c = \frac{1}{[O_2]} \text{ [Since, } [SO_2] = [SO_3]]
$$

\n
$$
10 = \frac{1}{x/5} \Rightarrow x = 0.5 \text{ mole}
$$

(5) (A). Total moles = $2 + 2 + 2 = 6$

P_{PCl₃} =
$$
\frac{2}{6}
$$
 × 3, P_{PCl₅} = $\frac{2}{6}$ × 3, P_{Cl₂} = $\frac{2}{6}$ × 3
\nK_p = $\frac{P_{PCl_3} × P_{Cl_2}}{P_{PCl_5}} = \frac{1 × 1}{1} = 1$ atm
\n**(6)** (A). N₂ + O₂ ⇒ 2NO (12) (B).
\n3 2 0 (12) (B).
\n3 - x 2 - x 2x 2
\n∴ 2x = 1; x = 0.5
\n[N₂] = $\frac{3-0.5}{2}$ = 1.25
\n(7) (A). H₂(g) + S(s) ⇒ H₂S(g)
\nConc. at equ. 0.5 - x - x
\nK_c = $\frac{[H_2S]}{[H_2]}$ ⇒ 7 × 10⁻² = $\frac{x}{0.5-x}$; x = 0.0327 (1) (B)
\n $P_{H_2S} = \left(\frac{n_{H_2S}}{V}\right)$ RT
\n⇒ 0.0327 × 0.0821 × 360 = 0.966 atm
\n**(8)** Q = $\frac{1 × 1}{1 × 1}$ = 1 ∴ R

$$
K_c = \frac{1429 \text{ J}}{[H_2]} \Rightarrow 7 \times 10^{-2} = \frac{x}{0.5 - x} \; ; \; x = 0.0327
$$
\n(1)\n(B)\n
\n
$$
P_{H_2S} = \left(\frac{n_{H_2S}}{V}\right) RT
$$
\n
$$
\Rightarrow 0.0327 \times 0.0821 \times 360 = 0.966 \text{ atm}
$$
\n(2)

(8)
$$
Q = \frac{1 \times 1}{1 \times 1} = 1
$$

 \therefore Q > K_c so reaction will proceed in backward direction $A_2(g) + B_2(g) \implies C_2(g) + D_2(g)$

 \therefore K'_c is very high, so $2x \approx 0.8 \Rightarrow x \approx 0.4$ \therefore 0.8 – 2x \approx y

$$
10^5 = \frac{0.4}{(y)^2} \quad ; \quad y = (0.4 \times 10^{-5})^{1/2} = 2 \times 10^{-3}
$$

$$
\frac{[A]}{[A_2]} = \frac{y}{x} = \frac{2 \times 10^{-3}}{0.4} = \frac{5}{1000} = \frac{1}{200}
$$

- **(3) (D).** Addition of F_2 shits equilibrium in forward direction.
- **(4) (C).** Formation of diamond is endothermic reaction and at high pressure density increases.
- **(5) (B).** Ni combines with CO and produced Ni(CO)_4 so concentration Cl_2 increases and eq. in Ist reaction shifts in backward direction.
	- \therefore Concentration PCl₃ decreases.

(6) (A).
$$
\alpha = \frac{D-d}{d(n-1)} \times 100 = \frac{46-30}{30(2-1)} \times 100 = 53.3\%
$$

- **(6) (A).** $\alpha = 3.314 \text{ K}^{-1} = 3.30 \text{ K}^{-1} = 48 \text{ K} = 1.8 \times 10^{-7}$

(a). $\alpha = 1.24 \text{ K} = 1.24 \text{ K}$ **EXECUTE ALTERAL CHEMISTI**
 $\frac{[A]}{[A_2]} = \frac{y}{x} = \frac{2 \times 10^{-3}}{0.4} = \frac{5}{1000} = \frac{1}{200}$

(7) HCl
 $\frac{1}{[A_2]} = \frac{y}{x} = \frac{2 \times 10^{-3}}{0.4} = \frac{5}{1000} = \frac{1}{200}$

(7) HCl

Comparison of diamond is encodibiremine reaction an **EXECUTE 100**
 $\frac{1}{1000} = \frac{2 \times 10^{-3}}{0.4} = \frac{5}{1000} = \frac{1}{200}$

(7) HCl Cr
 $\frac{1}{1000} = \frac{5}{200}$

of F₂ shits equilibrium in forward direction.

sure density increases.

sure density increases.

times with CO and (7) We know that $\Delta G^{\circ} = -2.303$ RT log K Now, $R = 8.314$ JK⁻¹, T = 298 K, K = 1.8×10^{-7} Putting these value in the above expression, we have $\Delta G^{\circ} = -2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})$ $=-5705.85 \times (-6.7447) = 38484 \text{ J} = 38.484 \text{ kJ}$ 10 and produced virtory₄ so

inc.

increases and eq. in Is reaction shifts
 1_3 decreases.
 $1_4 + 1_5$ (1-1) and the state of t Coreases. $[OH^-] = K_b \frac{[CH_3NH_2]}{[CH_3NH_3^+]}$
 $\frac{46-30}{30(2-1)} \times 100 = 53.3\%$
 $\frac{46-30}{30(2-1)} \times 100 = 53.3\%$
 $[OH^-] = \frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}$
 $\frac{1}{0.008}$ RK $K = 1.8 \times 10^{-7}$
 $\frac{1}{1000}$
 $\frac{1}{1000$ $\frac{3^3}{11.5 \times 10^{-3}}$
 $\frac{3^3}{11.5 \times 10^{-3}}$
 $= 1.5 \times 10^{-3}$
 $= 1.5 \times 10^{-3}$
 $= 1.5 \times 10^{-3}$
 $= 1.5 \times 10^{-3} + 0.1215 = 0.13 \text{ M}$
 $= 1.5 \times 10^{-3} + 0.1215 = 0.13 \text{ M}$
 $= 1.5 \times 10^{-3} + 0.1215 = 0.13 \text{ M}$
 $= 1.5 \times 10^{-3} +$ = $\frac{D-d}{d(n-1)} \times 100 = \frac{46-30}{30(2-1)} \times 100 = 53.3\%$

(DH⁻) = $\frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}$

the sevalue in the above expression, we have
 $= -5705.85 \times (-6.7447) = 38484 \text{ J} = 38.484 \text{ kJ}$
 $= -5705.85 \times (-6.$ WIC EOUILIBRIUM

Solution PCI₃ decreases.
 $[OH^-] = K_b \frac{[CH_3NH_2]}{[CH_3NH_3^2]}$
 $(0+1) \times 100 = \frac{46-30}{30 (2-1)} \times 100 = 53.3\%$
 $[OH^-] = \frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}$

se value in the above expression, we have
 -23
- **(8)** (A) **(9)** (D)

IONIC EQUILIBRIUM TRY IT YOURSELF-1

(1) (C). The reaction is NH₃ + H₂O
$$
\rightleftharpoons
$$
 NH₄⁺ + OH⁻

$$
K_{b} = 1.8 \times 10^{-5} = \frac{[NH_{4}^{+}][OH^{-}]}{[NH_{3}]}
$$

$$
=\frac{(1.5\times10^{-3})(1.5\times10^{-3})}{\text{[NH}_3]} = 1.8\times10^{-5}
$$

Equilibrium conc. of ammonia $[NH_3] = 0.125 M$ Total [NH₃] = $1.5 \times 10^{-3} + 0.125 = 0.13$ M

(2) **(B).**
$$
Zn^{++} + H_2O \rightleftharpoons Zn(OH)_2 + 2H^+
$$

$$
\Delta G^{\circ} = -2.303 \times 8.314 \text{ J} \text{K}^{-1} \times 298 \text{ K} \times \log (1.8 \times 10^{-7})
$$
\n= -5705.85 × (-6.7447) = 38484 J J = 38.484 kJ
\n(A)
\n(A)
\n(A)
\n(A)
\n(B)
\n10NIC EQUILIBRIUM
\n(C) The reaction is NH₃ + H₂O ⇒ NH₄⁺ + OH⁻
\n
$$
K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}
$$
\n
$$
= \frac{(1.5 \times 10^{-3})(1.5 \times 10^{-3})}{[NH_3]} = 1.8 \times 10^{-5}
$$
\n
$$
= \frac{(1.5 \times 10^{-3})(1.5 \times 10^{-3})}{[NH_3]} = 1.8 \times 10^{-5}
$$
\n
$$
= 1.8 \times 10^{-5}
$$
\n
$$
= 1.8 \times 10^{-5}
$$
\n
$$
= 1.8 \times 10^{-5} = 1.8 \times 10^{-5}
$$
\n
$$
= 1.8 \times 10
$$

$$
K_{\rm w} = [H^+] [OH^-] \quad \therefore \quad \frac{K_{\rm w}^2}{K_{\rm b}} = K_{\rm h}
$$
\nBy us

\n
$$
rH = r
$$

(3) (B).
$$
[H^+] = \frac{0.1}{100} = 10^{-3}
$$
; pH = 3

(4) **(C).** H₂A
$$
\rightleftharpoons
$$
 H⁺ + HA⁻; $k_1 = 1 \times 10^{-5}$
\n
$$
HA^{-} \rightleftharpoons H^{+} + A^{2-} ; k_2 = 5 \times 10^{-10}
$$
\n
$$
H_2A \rightleftharpoons 2H^{+} + A^{2-} ; k = k_1 \cdot k_2 = 5 \times 10^{-15}
$$

$$
(5) \qquad (C)
$$

(6) (C)

EXAMPLE 23.233 × 8.314 JK⁻¹ × 298 K × log(1.8 × 10⁻⁷)
\n= -5705.85 × (-6.7447)=38484 J = 38.484 kJ
\n**1000** (A), Ag⁺ + NH₃ ⇒ Ag(NH₃)⁺, K₁ = 3.5 × 10⁻³;
\n
$$
= H+ + CO3-2
$$
\n**11.1** 3.23 × 10⁻³ (10) (A), Ag⁺ + HCl⁻³ (2) (B).
\n
$$
= 2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})
$$
\n
$$
= 2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})
$$
\n
$$
= 2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})
$$
\n
$$
= 2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})
$$
\n
$$
= 2.303 \times 8.314 \text{ JK}^{-1} \times 298 \text{ K} \times \log(1.8 \times 10^{-7})
$$
\n
$$
= -5705.85 \times (-6.7447) = 38484 J = 38.484 kJ
$$
\n
$$
= 2.303 \times 8.314 J K-1 × 298 K × log(1.8 × 10-7)
$$
\n
$$
= 2.303 \times 8.314 J K-1 × 298 K × log(1.8 × 10-7)
$$
\n
$$
= -5705.85 \times (-6.7447) = 38484 J = 38.484 kJ
$$

$$
base \qquad \text{conjugate acid} \qquad \qquad \oplus
$$

(8) (B).
$$
CH_3NH_2 + HCl \longrightarrow CH_3 NH_3 + Cl
$$

Initially 0.1 0.08 –
In solution 0.02 – 0.08

$$
0.08
$$

$$
[OH^{-}] = K_{b} \frac{[CH_{3}NH_{2}]}{[CH_{3}NH_{3}^{+}]}
$$

$$
[OH^-] = \frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}
$$

$$
[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14} \times 4}{5 \times 10^{-4}} = 8 \times 10^{-11}
$$

(7) HCl
\nacid
\nH₂O
\nhase
\nH₂O
\n(8) (B). CH₃NH₂ + HCl
$$
\longrightarrow
$$
 CH₃NH₃ + Cl⁻
\nInitially 0.1 0.08 -
\nIn solution 0.02 - 0.08
\n[OH⁻] = K_b $\frac{[CH_3NH_2]}{[CH_3NH_3^+]}$
\n[OH⁻] = $\frac{5 \times 10^{-4} \times 0.02}{0.08} = \frac{5}{4} \times 10^{-4}$
\n[H⁺] = $\frac{K_w}{[OH^-]} = \frac{10^{-14} \times 4}{5 \times 10^{-4}} = 8 \times 10^{-11}$
\n(9) (A). CO₂ + H₂O \rightleftharpoons H₂CO₃ \rightleftharpoons H⁺ + HO₃⁻
\n \rightleftharpoons H⁺ + CO₃⁻²

(10) (A). $Ag^+ + NH_3 \rightleftharpoons Ag(NH_3)^+$; $K_1 = 3.5 \times 10^{-3}$; $\text{Ag(NH}_3)^+ + \text{NH}_3 \rightleftharpoons \left[\text{Ag(NH}_3)_2\right]^+; \text{K}_2 = 1.7 \times 10^{-3}$ Adding :

––––––––––––––––––––––––––––––––––––

$$
Ag^{+} + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]^{+}
$$

K = K₁ × K₂ = 5.95 × 10⁻⁶

(11) (A). Rate in weak acid =
$$
\frac{1}{100}
$$
 (rate in strong acid)

$$
\therefore \quad [H^+]_{\text{weak acid}} = \frac{1}{100} [H^+]_{\text{strong acid}}
$$
\n
$$
\therefore \quad [H^+]_{\text{weak acid}} = \frac{1}{100} M = 10^{-2} M
$$
\n
$$
\therefore \quad C\alpha = 10^{-2}
$$
\n
$$
\therefore \quad K_a = 10^{-4}
$$

TRY IT YOURSELF-2

 K_w^2 By using Henderson equation **(1) (A).** 50mL of 0.02 M NaOH contains 1 milli moles of OH– . 50mL of 0.04 M CH₃COOH contains 2 millimoles of H⁺. The combination of the two solutions will give 1 milli mole of salt, and 1 milli mole of acid will be left behind. $d = \frac{1}{100}$ (rate in strong acid)
 $d = \frac{1}{100}$ [H⁺]_{strong acid}
 $d = \frac{1}{100} M = 10^{-2} M$
 YOURSELF-2

NaOH contains 1 milli moles of OH⁻.

Duang-COOH contains 2 millimoles of H⁺.

1 of the two solutions will g $d = \frac{1}{100} [H^+]_{\text{strong acid}}$
 $d = \frac{1}{100} M = 10^{-2} M$
 YOURSELF-2

NaOH contains 1 milli moles of OH⁻.

Discriming the two solutions will give 1 milli

1 milli mole of acid will be left behind.

I milli mole of acid will b

$$
\frac{1}{K_b} = K_h
$$

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

or
$$
pH = pK_a + log \frac{[0.1 \text{ milli mole per 100mL of soln.}]}{[0.1 \text{ milli mole per 100mL of soln.}]}
$$

or $pH = -\log K_a = -\log(1.8 \times 10^{-5}) = 4.74$

(2) When half of NH_4OH is neutralised.
 $[NH_4OH] = NH_4Cl$.

$$
[\text{NH}_4\text{OH}] = \text{NH}_4\text{Cl}.
$$

By using the Henderson equation,

pOH = pK_b + log
$$
\frac{[Salt]}{[Base]}
$$
 or pOH = 4.75
pH = 14-4.75 = 9.25

(3) (C)

(4) (A). The degree of hydrolysis of a salt of weak acid and strong base is given by

$$
\alpha_h = \sqrt{\frac{K_w}{K_a C}} = \sqrt{\frac{1 \times 10^{-4}}{1 \times 10^{-5} \times 0.1}} = \sqrt{\frac{1 \times 10^{-4}}{1 \times 10^{-6}}} = 1 \times 10^{-4}
$$

or $100\alpha_h = 100 \times 1 \times 10^{-4} = 10^{-2} = 0.01$ So, the degree of hydrolysis of NaX $(0.1 M)$ is 0.01%. (5)

(5) (B). HCl is strong acid. NH_4Cl is a salt of weak base and strong acid - NH_4^+ hydrolysis to give acid solution. NaCl is a salt of strong acid and strong base - No hydrolysis neutral solution. NaCN is a salt of strong base and weak acid

 - CN– hydrolysis to give basic solution. **(6) (D).** Since it is a basic buffer so

 $pOH = pK_b + log [Salt]/[Base]$ $pOH = -\log K_b + \log [Salt]/[Base]$ $=-\log(1.8 \times 10^{-5}) + \log(0.25/0.15)$ $= 4.74 + 0.22 = 4.96$ or $pOH = 4.97$ $pH = 14 - pOH = 14 - 4.97 = 9.03$ **TRY IT YOURSELF-3**

(1) (B). Mg(OH)₂ \Rightarrow Mg²⁺ + 2OH⁻; K_{sp} = 1.2 × 10⁻¹¹.

[OH–] higher than that contained in a saturated solution will cause precipitation. Therefore, the solution must be at the point of attaining equilibrium and the concentration of ions in solution must be no greater than those required to satisfy the solubility product constant. In this solution, $[Mg^{2+}] = 0.10 M$ and $[Mg^{2+}][OH^-]^2 = 1.2 \times 10^{-11}$. . vact. N is a since the specified becomes the second of Sale $\frac{1}{2}$ ($\frac{1}{2$

$$
[OH^-]^2 = \frac{1.2 \times 10^{-11}}{0.10} = 1.2 \times 10^{-10}.
$$
 (7)

or
$$
[OH^-] = 1.1 \times 10^{-5}
$$
 or $pOH = 4.96$ or $pH = 9.04$

(2) Solubility product of Pb₃(PO₄)₂ = 1.5 × 10⁻³².
Pb₃(PO₄)₂
$$
\rightleftharpoons
$$
 3Pb2++2PO₄³⁻

If x is the solubility of $Pb_3(PO_4)_2$ Then K_{sp} = $(3x)^3 (2x)^2 = 108x^5$

$$
x = 5\sqrt{\frac{K_{sp}}{108}} = \left(\frac{1.5 \times 10^{-32}}{108}\right)^{1/5}
$$

 $x = 1.692 \times 10^{-7}$ moles/lit. Molecular mass of $Pb_3(PO_4)_2 = 811$ $x = 1.692 \times 10^{-7}$ m \times 811 g/lit = 1.37 \times 10–4 g/lit Solubility product is

RIUM	TRY SOLUTIONS	ESP ²⁺ J [C ₂ O ₄ ²⁻]	
POOH = pK _b + log	[Salt]	or pOH = 4.75	$K_{sp}(SrC2O4) = [Sr2+] [C2O42-]$
OH = 14-4.75 = 9.25	(3)	(b) HgSO4 $\rightleftharpoons Hg^{2+} + SO_4^{2-}$	
The degree of hydrolysis of a salt of weak acid and strong base is given by	(4)	(C). 50 lit 10 ⁻⁵ mole/Alt = 8 mole/m ³ .	
$\alpha_h = \sqrt{\frac{K_w}{K_a C}} = \sqrt{\frac{1 \times 10^{-4}}{1 \times 10^{-5} \times 0.1}} = \sqrt{\frac{1 \times 10^{-4}}{1 \times 10^{-6}}} = 1 \times 10^{-4}$	$[Ag+] = \frac{10^{-5}}{100} = 10^{-7} M + 10^{-7} M Br^{-}$		
So, the degree of hydrolysis of NaX (0.1 M) is 0.01%. (5)	9. HA + NaOH \rightarrow NaA + H ₂ O At the end point, the solution contains only NaA whose concentration is 0.1/2 = 0.05 M. Since the salt NaA is formed by strong alkali (NaOH)		

(C). 50 lit 10^{-5} mole AgCl + 50 lit of 2×10^{-7} M HBr

$$
[Ag^{+}] = \frac{10^{-5}}{100} = 10^{-7} M + 10^{-7} M Br^{-}
$$

K = 10⁻⁷ × 10⁻⁷ = 10⁻¹⁴ × K_{sp} = 5 × 10⁻³
So, no. precipitation and [Ag⁺] = 10⁻⁷ M

9. HA + NaOH \rightarrow NaA + H₂O At the end point, the solution contains only NaA whose concentration is $0.1/2 = 0.05$ M. Since the salt NaA is formed by strong alkali (NaOH) and weak acid HA (indicated by its low K_a value), its $R_{sp} = [11g \text{ J} [304 \text{ J} - 3] - 0.4 \times 10^2$
 $S = 8 \times 10^{-3} \text{ mole/lit} = 8 \text{ mole/m}^3$.

50 lit $10^{-5} \text{ mole AgCl} + 50$ lit of 2×10^{-7} M HBr
 $[Ag^+] = \frac{10^{-5}}{100} = 10^{-7} M + 10^{-7} M \text{ Br}^{-}$
 $K = 10^{-7} \times 10^{-7} = 10^{-14} < K_{sp} = 5 \times 10^{-3}$

So SrC₂O₄) = [Sr²⁺] [C₂O₄²⁻]

= (5.4 × 10⁻⁴)² = 2.92 × 10⁻⁷

HgSO₄ = Hg²⁺ + SO₄²⁻

K_{sp} = [Hg²⁺] [SO₄²⁻] = S² = 6.4 × 10⁻⁵

S = 8 × 10⁻³ mole/lit = 8 mole/m³.

50 lit 10⁻⁵ mole AgCl $\frac{1}{2}O_4$) = [Sr²⁺] [C₂O₄²⁻]

= (5.4 × 10⁻⁴)² = 2.92 × 10⁻⁷
 $O_4 \rightleftharpoons Hg^{2+} + SO_4^{2-}$

= [Hg²⁺] [SO₄²⁻] = S² = 6.4 × 10⁻⁵
 $\frac{1}{2}$ × 10⁻⁵ mole/lit = 8 mole/m³.

t 10⁻⁵ mole AgCl + 50 li \Rightarrow Hg²⁺ + SO₄²⁻
 g^{2+}] [SO₄²⁻] = S² = 6.4 × 10⁻⁵

0⁻³ mole/lit = 8 mole/m³.

5 mole AgCl + 50 lit of 2 × 10⁻⁷ M HBr
 $\frac{10^{-5}}{100}$ = 10⁻⁷ M + 10⁻⁷ M Br⁻

× 10⁻⁷ = 10⁻¹ M + 10⁻⁷ M Br = (5.4 × 10⁻⁴)²= 2.92 × 10⁻⁷

= (5.4 × 10⁻⁴)²= 2.92 × 10⁻⁷

0₄ ⇒ Hg²⁺ + SO₄²⁻

= [Hg²⁺] [SO₄²⁻] = S² = 6.4 × 10⁻⁵

× 10⁻⁵ mole/lit = 8 mole/m³.

+ 10⁻⁵ mole AgCl + 50 lit of 2 × 10⁻⁷

 $1_{(14.5,2010,(1,2010)-0}$ 2^{\sim}

(6) (D). BOH + HCl \longrightarrow BCl + H₂O C_c B^+ + H₂O \rightleftharpoons BOH + H⁺ $C(1-h)$ Ch Ch $K = 10^{-7} \times 10^{-7} = 10^{-14} < K_{\text{sp}} = 5 \times 10^{-3}$

So, no. precipitation and $[Ag^+] = 10^{-7}$ M

HA + NaOH \rightarrow NaA + H₂O

At the end point, the solution contains only NaA whose

concentration is 0.1/2 = 0.05 M.

Since the sa $K_{sp} = 3 \times 10^{-7}$
Ag⁺] = 10⁻⁷ M
O
on contains only NaA whose
05 M.
ed by strong alkali (NaOH)
ted by its low K_a value), its
e following reaction:
gC)
1.3010) = 9
1+ H₂O
(OH + H⁺
Ch Ch
 \times (2/5) = 7.5 ml.
 $\frac{2.5$ Volume of HCl used = $\frac{2.5 \times (2/5)}{2/15}$ = 7.5 ml. HA + NaOH \rightarrow NaA + H₂O

At the end point, the solution contains only NaA whose

concentration is 0.1/2 = 0.05 M.

Since the salt NaA is formed by strong alkali (NaOH)

since the salt NaA is formed by strong alkali (Na and weak acid HA (indicated by its low K_a value), its

bH can be evaluated by the following reaction:
 $pH = \frac{1}{2} (pK_w + pK_a + log C)$
 $= \frac{1}{2} (14 + 5.3010 + (-1.3010) = 9$

BOH + HCl $\longrightarrow BC1 + H_2O$

C
 B^+ + $H_2O \rightleftharpoons BDH + H^+$

Concentration of Salt,
$$
C = \frac{2.5 \times (2/5)}{10} = 0.1 M
$$

$$
\therefore \quad \frac{\text{Ch}^2}{1-\text{h}} = \frac{\text{K}_{\text{w}}}{\text{K}_{\text{b}}} \text{ Solving, h} = 0.27
$$

[H⁺] = Ch = 0.1 × 0.27 = 2.7 × 10⁻² M

$$
2.5 \times (2/5)
$$

\n
$$
3.6 \times (2/5) = 1.6 \times 10^{-10}
$$

\n
$$
3.7 \times (2/5) = 100 \times 10^{-10}
$$

\n
$$
3.8 \times (2/5) = 100 \times 10^{-10}
$$

\n
$$
3.8 \times (2/5) = 1.5 \times 10^{-11}
$$

\n
$$
3.8 \times (2/5) = 7.5 \text{ ml.}
$$

\n
$$
3.8 \times (2/5) = 7.5 \text{ ml.}
$$

\n
$$
3.8 \times (2/5) = 7.5 \text{ ml.}
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3.8 \times (2/5) = 7.5 \text{ ml.}
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3.8 \times (2/5) = 7.5 \text{ ml.}
$$

\n
$$
3.8 \times (2/5) = 7.5 \text{ ml.}
$$

\n
$$
3.8 \times (2/5) = 7.5 \times (2/5) = 7
$$

.

CHAPTER-7 :EQUILIBRIUM EXERCISE-1

- **(1) (D).** CO_2 (gas) $\rightleftharpoons CO_2$ (in solution) This equilibrium is governed by Henry's law which states that the mass of a gas dissolved in a given mass of a solvent at any temperature is proportional (13) to the pressure of the gas above the solvent.
- **(2) (C).** Ice and water kept in a perfectly insulated thermos flask (no exchange of heat between its contents and the surroundings) at 273K and the atmospheric pressure are in equilibrium state and the system (14) shows interesting characteristic features. The system (15) shows solid-liquid equilibrium.
- **(3) (D).** Boiling point of the liquid depends on the atmospheric pressure. It depends on the altitude of the place; at high altitude the boiling point decreases.
- **(4) (C).** H₂O (ℓ) \rightleftharpoons H₂O (vap.) Rate of evaporation = Rate of condensation
- **(5) (C).** Chemical equilibrium in the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ can be attained from either direction.
- **(6) (B).** The concentration of reactants decreases and that of products increases with time. Rate of reaction increases with time.

At equilibrium, $R_f = R_b$

(7) (D). Let us write equilibrium constant for the reaction, $H_2(g) + I_2(g) \rightleftharpoons 2HI(g);$

as
$$
K_c = \frac{[H I]^2}{[H_2][I_2]} = x
$$
 $= \frac{[C]^c}{[A]^a}$

The equilibrium constant for the reverse reaction 2HI $(g) \rightleftharpoons H_2(g) + I_2(g)$, at the same temperature is

$$
K'_{c} = \frac{[H_{2}][I_{2}]}{[HI]^{2}} = \frac{1}{x} = \frac{1}{K_{c}}.
$$
 Thus, $K'_{c} = \frac{1}{K_{c}}$

Equilibrium constant for the reverse reaction is the inverse the equilibrium constant for the reaction in (20) the forward direction.

(8) **(B).**
$$
2XY \rightleftharpoons X_2 + Y_2
$$
; $K_c = 81$

$$
XY \rightleftharpoons \frac{1}{2} X_2 + \frac{1}{2} Y_2
$$
; $K_c' = ?$; $K_c' = \sqrt{K_c} = \sqrt{81} = 9$

(9) **(A).**
$$
K_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]} = \frac{(1.9)^2}{(0.6)^2 (0.82)} = 12.229 \text{L} \text{ mol}^{-1}
$$
 (21) **(B).** N

(10) (A). The equilibrium constant for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ can be written as,

Area.
$$
\frac{1}{2} \text{Area} + \frac{1}{
$$

$$
K = \frac{(a/2) (a/2)}{(a/2)^2} = 1
$$

- **STUDY MATERIAL: CHEMISTRY**
 $K = \frac{(a/2) (a/2)}{(a/2)^2} = 1$

action is multiplied by 2, the equilibrium constant

mes square of the previous value.
 $K=7^2=49$

the reaction equilibrium constant, K_c can be

en as,
 10^{2} STUDY MATERIAL: CHEMISTRY
 $=\frac{(a/2) (a/2)}{(a/2)^2} = 1$

ion is multiplied by 2, the equilibrium constant

es square of the previous value.
 $= 7^2 = 49$

reaction equilibrium constant, K_c can be

as,
 $= \frac{1^2}{(a^2 + 10^{-3})^2$ **(12) (A).** If reaction is multiplied by 2, the equilibrium constant becomes square of the previous value. $K = 7^2 = 49$ **STUDYMATERIAL: CHEMISTRY**
 $K = \frac{(a/2)(a/2)}{(a/2)^2} = 1$
 c). If reaction is multiplied by 2, the equilibrium constant

becomes square of the previous value.
 $K = 7^2 = 49$

For the reaction equilibrium constant, K_c can b **STUDY MATERIAL: CHEMISTRY**
 $K = \frac{(a/2)(a/2)}{(a/2)^2} = 1$

If reaction is multiplied by 2, the equilibrium constant

becomes square of the previous value.
 $K=7^2=49$

For the reaction equilibrium constant, K_c can be

writ **ERIAL: CHEMISTRY**

e equilibrium constant

us value.

onstant, K_c can be
 $\frac{3 M)^2}{(2 \times 10^{-3} M)} = 0.622$

all the reactants and

e. **MATERIAL: CHEMISTRY**

by 2, the equilibrium constant

previous value.

rium constant, K_c can be
 $\frac{8 \times 10^{-3} M)^2}{(4.2 \times 10^{-3} M)} = 0.622$

stem, all the reactants and

e phase.

(RT)^{Δn}
 $\Delta n = (2 + 1) - 2 = 1$ **DYMATERIAL: CHEMISTRY**
 $\begin{bmatrix}\n=1 \\
=1\n\end{bmatrix}$

ed by 2, the equilibrium constant

the previous value.

ilibrium constant, K_c can be
 $\frac{(2.8 \times 10^{-3} \text{M})^2}{(0^{-3} \text{M})(4.2 \times 10^{-3} \text{M})} = 0.622$
 $\begin{bmatrix}\n0^{-3} \text{M} & (4.2 \$ **STUDY MATERIAL: CHEMISTRY**

(2.0 1 (2.2)

(2.2)² = 1

multiplied by 2, the equilibrium constant

are of the previous value.

9

on equilibrium constant, K_c can be

(2.8 × 10⁻³ M)²

(3.0 × 10⁻³ M) (4.2 × 10⁻³ **YMATERIAL: CHEMISTRY**

²

²

²

²
 by 2, the equilibrium constant

² previous value.

brium constant, K_c can be

<u>^{2.8}×10⁻³M)²

²

³M)(4.2×10⁻³M)²

²

²

(822

2

</u> **STUDYMATERIAL: CHEMISTRY**
 $\frac{1}{2}$ (a/2)² = 1

s multiplied by 2, the equilibrium constant

quare of the previous value.

= 49

etion equilibrium constant, K_c can be

= $\frac{(2.8 \times 10^{-3} \text{M})^2}{(3.0 \times 10^{-3} \text{M}) (4.2 \$ TUDY MATERIAL: CHEMISTRY
 $\frac{2}{2} = 1$

plied by 2, the equilibrium constant

of the previous value.

quilibrium constant, K_c can be
 $\frac{(2.8 \times 10^{-3} \text{M})^2}{\times 10^{-3} \text{M}} = 0.622$
 $\times 10^{-3} \text{M} (4.2 \times 10^{-3} \text{M}) = 0.622$ **STUDYMATERIAL: CHEMISTRY**
 $K = \frac{(a/2)(a/2)}{(a/2)^2} = 1$

ffreaction is multiplied by 2, the equilibrium constant

becomes square of the previous value.
 $K = 7^2 = 49$

For the reaction equilibrium constant, K_c can be

writ **STUDY MATERIAL: CHEMISTRY**
 $K = \frac{(a/2)(a/2)}{(a/2)^2} = 1$

aation is multiplied by 2, the equilibrium constant

mess square of the previous value.
 $K = 7^2 = 49$

the reaction equilibrium constant, K_c can be

ten as,
 $NO_1^$ **DYMATERIAL: CHEMISTRY**

=1

d by 2, the equilibrium constant

e previous value.

ibrium constant, K_c can be
 $\frac{2.8 \times 10^{-3} M)^2}{(4.2 \times 10^{-3} M)} = 0.622$
 ζ_c

eystem, all the reactants and

me phase.
 $\zeta_c(RT)^{\Delta n}$

n
- **(13) (C).** For the reaction equilibrium constant, K_c can be written as,

$$
K_c = \frac{[NO]^2}{[N_2][O_2]} = \frac{(2.8 \times 10^{-3} M)^2}{(3.0 \times 10^{-3} M)(4.2 \times 10^{-3} M)} = 0.622
$$

- **(14) (B).** When $\Delta n_g = 0$, $K_p = K_c$
- **(15) (B).** In a homogeneous system, all the reactants and products are in the same phase.
- **(16) (B).** We know that, $K_p = K_c (RT)^{\Delta n}$ For the above reaction, $\Delta n = (2 + 1) - 2 = 1$ $K_p = 3.75 \times 10^{-6} (0.0831 \times 1069) = 0.033$

(17) **(A).**
$$
\begin{array}{ccc}\n & 1 & 1 \\
 & 1 & 1 \\
 & 0.06 & 0.12 & 0.216 \\
 & 0.21 & 0.216 & 0.216\n \end{array}
$$

$$
K_c = \frac{[2]}{[X][Y]^2} = \frac{0.210}{0.06 \times 0.12 \times 0.12} = 250
$$

(18) (A). All the reactants and products are in same physical state.

(A). For a general reaction, $aA + bB \rightleftharpoons cC + dD$

CHAPERE-RE-TEORAL OF
\n**CALPEN-TEORAL OF**
\n**EXERCISE-1**
\n
$$
C_0
$$
 (gas) =
\n C_0 (in solution)
\n C_0 (mass) =
\n C_0 (in solution)
\n C_0 (in equation)
\n C_0 (in

 $K_c' = \frac{1}{K_c}$ where Δn = (number of moles of gaseous reads)
(number of moles of gaseous reads) K_c (number of moles of gaseous reactants) in the balanced chemical equation.

(20) (A).
$$
PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)
$$

Initial conc.
$$
\frac{5}{5} = 1
$$
 0 0

At equilibrium $1 - 0.4$ 0.4 0.4

$$
K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{0.4M \times 0.4M}{0.6M} = 0.266 M
$$

(21) **(B)**. NH₄COONH₂(s)
$$
\rightleftharpoons
$$
 2NH₃(g) + CO₂(g)
2p p

 $\frac{E}{[A]^n[B]^b}$ (RT)^{(x-r}4)^{-(x-r}₄)⁻(RT)^{(x-r4})^{-(x-r4)})

entiture is
 $\frac{[C]^c[D]^d}{[A]^n[B]^b}$ (RT)⁴ⁿ = K_c (RT)^{An},

-
 $\frac{1}{[A]^n[B]^b}$ (RT)⁴ⁿ = K_c (RT)^{An},

(number of moles of gaseous products)-

(number rse reaction 2HI

temperature is
 $\epsilon = \frac{[C]^e[D]^d}{[A]^a[B]^b} (RT)^{An} = K_e (RT)^{An}$
 $\epsilon = \frac{1}{K_e}$ where $\Delta n = (\text{number of moles of gaseous products}) - (\text{number of moles of gaseous reactions})$ in the

reaction is the

the reaction in (20) (A). PCI₅(g) \Rightarrow PCI₃(g) + Cl₂(g)

Init exercencion 2HI

e temperature is
 $\frac{K_c}{K_c} = \frac{1}{k_c}$

where $\Delta n = (\text{number of moles of gaseous products}) - (\text{number of moles of gasous products})$

(and the reaction is the

se reaction is the

se reaction is (20) (A). $PCl_3(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

Initial cone. $\frac{5}{5} =$ $K_p = (2p)^2 \times p = 4p^3 = 4 \times (1)^3 = 4 \text{ atm}^3$ When volume and temperature are constant, the number of moles of a gas is proportional to its partial pressure. So, $2p + p = 3$; $3p = 3$: $p = 1$ atm $\sum_{j=1}^{S} [D]^d$ (RT)^{Δn} = K_c (RT)^{Δn},
 N^a [B]^b (RT)^{Δn} = K_c (RT)^{Δn},

tre Δn = (number of moles of gaseous reactants) in the

mored chemical equation.

PCl₃ (g) \rightleftharpoons PCl₃(g) + Cl₂ (g)

(B).
$$
P_4(s) + 5O_2(g) \rightleftharpoons P_4O_{10}(s)
$$
Since concentration of solids is taken as 1,

expression for equilibrium constant involves only oxygen.

$$
K_c = \frac{1}{\left[O_2\right]^5}
$$

 (22)

(23) (D). The equilibrium between water vapour and liquid water in a closed container is an example of heterogeneous equilibrium. $H_2O(\ell) \rightleftharpoons H_2O(g)$ In this example there is a gas phase and a liquid phase. (34) In the same way, equilibrium between two solid phases and a gas phase.

$$
\text{CaCO}_3\left(s\right) \rightleftharpoons \text{CaO}\left(s\right) + \text{CO}_2\left(g\right)
$$

(24) **(C).** $2Cu(NO_3)_2$ (s) \Rightarrow $2CuO(s) + 4NO_2(g) + O_2(g)$ Since conc. of solids is taken as 1, the expression for (36) K_c becomes $K_c = [NO_2(g)]^4 [O_2(g)]$

(25) (C). Important features of equilibrium constant as follows:

- (i) Expression for equilibrium constant is applicable only when concentrations of the reactants and products have atttained constant value at equilibrium state.
	- (ii) The value of equilibrium constant is independent of initial concentrations of the reactants and products.
	- (iii) Equilibrium constant is temperature dependent (37) having one unique value for a particular reaction represented by a balanced equation at a given temperature.
	- (iv) The equilibrium constant for the reverse reaction is equal to the inverse of the equilibrium constant for the forward reaction.
- **(26) (C).** Value of K_c predicts the extent of the reaction.
- **(27) (D).** All three statements are correct.
- **(28) (C).** When $Q_p \le K_p$, the rate of forward reaction is more than rate of backward reaction.
- **(29) (B).** The equilibrium constant helps in predicting the direction in which a given reaction will proceed at (39) any stage. For this purpose, we calcualte the reaction quotient Q. The reaction quotient, Q (Q_c with molar concentration and Q_p with partial pressures) is defined in the same way as the equilibrium constant K_c except that the concentrations in Q_c are not necessarily equilibrium values.
- **(30) (D).** Important applications of equilibrium constant to
	- predict the extent of a reaction on the basis of its magnitude.
	- predict the direction of the reaction, and
	- calculate equilibrium concentrations.
- **(31) (D).** The value of K_c for a reaction does not depend on the rate of the reaction. It is directly related to the thermodynamics of the reaction and in particular, to the change in Gibbs energy ΔG . If,
	- \sim ΔG is negative, then the reaction is spontaneous and proceeds in the forward direction.
	- ΔG is positive, then reaction is considered non– (41) spontaneous. Instead, as reverse reaction would have a negative ΔG , the products of the forward reaction shall be converted to the reactants.
	- ΔG is 0, reactions has achieved equilibrium; at this point, there is no longer any free energy left to drive the reaction.
- **(32) (D).** $-\Delta G^{\circ} = RT \ln K_c$
(33) (B). A mathematical
- **(33) (B).** A mathematical expression of the thermodynamic

view of equilibrium can be described by the following equation $\Delta G = \Delta G^{\Theta} + RT \ln Q$

- where, ΔG^{Θ} is standard Gibbs energy.
- **(34) (D).** Le–Chatelier's principle is applicable to all physical and chemical equilibria.
- **(35) (C).** Since number of moles of gaseous reactants and products are same, the reaction will not be affected by changing the pressure.
- **(36) (B).** Fe³⁺(aq) + SCN⁻(aq) \rightleftharpoons [Fe(SCN)]²⁺(aq) YellowColourless Deep red When oxalic acid is added, it reacts with $Fe³⁺$ ions to form stable complex ion $[Fe(C_2O_4)_3]^{3-}$, thus decreasing the conc. of free $Fe^{3+}(aq)$. Now, according to Le Chatelier's principle, the reaction will shift in backward direction to increase the conc, of free Fe³⁺(aq). Thus, conc. of $[Fe(SCN)]^{2+}$ decreases, so the intensity of red colour decreases.
	- **(37) (D).** In general, the temperature dependence of the equilibrium constant depends on the sign of ΔH for the reaction.
		- The equilibrium constant for an exothermic reaction (negative ΔH) decreases as the temperature increases.
		- The equilibrium constant for an endothermic reaction (positive ΔH) increases as the temperature increases. Temperature changes affect the equilibrium constant and rates of reaction.
- **(38) (A).** Equilibrium constant (K) is temperature dependent as temperature is constant thus, K will remain same.
- **(39) (B).** Endothermic reaction is favoured at high temperature.
- **(40) (D).** Addition of hydrogen at equilibrium results in value of Q_c being has less than K_c . Thus, in order to attain equilibrium again reaction moves in the forward direction.

Similarly, we can say that removal of a product also boosts the forward reaction and increases the concentration of the products and this has great commercial application in cases of reactions, where the product is a gas or a volatile substance. In case of manufacture of ammonia, ammonia is liquified and removed from the reaction mixture so that reaction keeps moving in forward direction. Similarly, in the large scale production of CaO (used as important building material) from $CaCO_3$, constant removal of $CO₂$ from the kiln drives the reaction to completion. It should be remembered that continuous removal of a product maintains Q_c at a value less than K_c and reaction continues to move in the forward direction.

- **(41) (B).** Since reaction is exothermic it is favoured at low temperature. There is a decrease in number of moles hence reaction is favoured at high pressure.
- **(42) (C).** A catalyst increases the rate of the chemical reaction by making available a new low energy pathway for the conversion of reactants to products. It increases the rate of forward and reverse reaction that pass through the same transition state and does not affect equilibrium.

Catalyst lowers the activation energy for the forward and reverse reactions by exactly the same amount.

- **(43) (A).** Addition of KSCN increases the colour intensity the solution as it shifts the equilibrium to right. Addition of reagents like oxalic acid or Hg^{2+} ions which remove (58) $Fe³⁺$ or SCN⁻ ions shift the equilibrium to the left and colour intensity decreases.
- **(44) (D). Faraday** classified electrolytes into strong and weak electrolytes. **Strong** electrolytes on dissolution in water are ionised almost completely, while the **weak** electrolytes are only partially dissociated.
- **(45) (C).** Sodium chloride (common salt) is an important component of our diet and is formed by reaction between hydrochloric acid and sodium hydroxide.

$$
(46) (A).
$$

- **(47) (C).** Electron deficient species like AlCl₃, Co^{3+} , Mg²⁺ etc. can act as Lewis acids while species like H_2O , NH_3 , OH⁻ etc. which can donate a pair of electrons, can act as Lewis bases.
- **(48) (C).** BF_3 is an electron deficient compound and hence is a Lewis acid.
- **(49) (C).** (X) and (Y) in the given figure are base and acid.
- **(50) (D).** Lewis acid is acceptor of a pair of electrons while Lewis base is donor of a pair of electrons.
- **(51) (B).** According to Bronsted–Lowry theory, acid is a substance that is capable of donating a hydrogen ion H^+ and bases are substances capable of accepting a hydrogen ion, H^+ . In short, acids are proton donors and bases are proton acceptors.
- **(52) (B).** The conjugate acid should have one extra proton in each case and therefore the corresponding conjugate (64) acids are NH_3 , NH_4^+ and HCOOH respectively.
- **(53) (C).** Strong acid dissociates completely in water, the resulting base formed would be very weak, i.e., strong acids have very weak conjugate bases.
- **(54) (C).** The pH of neutral water at 25°C is 7.0.

 \therefore [H⁺] = [OH⁻] = 10⁻⁷ $(pH = -\log[H^+])$ Now, $K_w = [H^+] [OH^-] = (1 \times 10^{-7})^2 = 1 \times 10^{-14}$ As the temperature increases, ionisation of water (67) increases, thus [H+] and [OH–] increases equally. Now $K_{\rm w} = [H^+] [OH^-] > 1 \times$ $] > 1 \times 10^{-14}$ (: [H⁺] = [OH⁻]) or $[H^+]^2 > 1 \times 10^{-14}$: $[H^+] > 1 \times 10^{-7}$ and pH < 7

- **(55) (B).** NaCN solution is basic in nature since HCN formed is a weak acid and does not hydrolyse. $NaCN + H₂O \rightleftharpoons NaOH + HCN$ $NaOH \rightleftharpoons Na^+ + OH^-$ (Basic solution)
- **(56) (A).** The values of dissociation constants for successive stages decrease.
- **(57) (D).** Consider the equation of water at 298 K $K_w = [H_3O^+]$ [OH⁻] = 10⁻¹⁴ Taking negative logarithm on both sides of equation,

we obtain

$$
-\log K_w = -\log \{[H_3O^+][OH^-]\}
$$

= $-\log [H_3O^+]-\log [OH^-] = -\log 10^{-14}$
pK_w = pH + pOH = 14

(58) (D). We can distinguish acidic, neutral and basic aqueous solutions by the relative values of the H_3O^+ and OH–
concentrations.

Acidic [H₃O⁺] > [OH⁻]
\nNeutral [H₃O⁺] = [OH⁻]
\nBasic, [H₃O⁺] < [OH⁻]
\n(59) **(B).** H₂SO₄
$$
\rightleftharpoons
$$
 2H⁺ + SO₄²⁻
\n[H⁺] = 2 × 1 × 10⁻⁴ M

$$
pH = -\log(2 \times 10^{-4}) = 3.70
$$

(60) (A). MOH (aq) \rightleftharpoons M⁺ (aq) + OH⁻(aq) The equilibrium constant for base ionization is called base ionization constant and is represented by K_b . It can be expressed in terms of concentration in molarity of various species in equilibrium by the following **STUDYMATERIAL: CHEMISTRY**
 $K_w = -\log {\{H_3O^+ \}\}$ (OH⁻] + $\log {[H_3O^+] - \log {[OH^-]}} = -\log 10^{-14}$
 $+ pOH = 14$
 $+ pOH = 14$

tinguish acidic, neutral and basic aqueous

or the relative values of the H_3O^+ and OH⁻
 O^+] > [OH **STUDY MATERIAL: CHEMISTRY**
 $= -\log {\{\left[\frac{H_3 O^+}{\right]} [OH^- \right]}}$
 $\frac{1}{3} O^+]-\log [OH^-] = -\log 10^{-14}$
 $OH = 14$

guish acidic, neutral and basic aqueous

ne relative values of the $H_3 O^+$ and OH⁻
 $\frac{1}{1} SO^+ = \left[OH^- \right]$
 $\left[H_3 O^+ \$ the relative values of the H_3O^+ and OH-

the relative values of the H_3O^+ and OH-
 H_3O^+ = [OH-]
 $[H_3O^+]$ = [OH-]
 H^+ + SO₄²-
 $\times 10^{-4}$ M
 $\times 10^{-4}$ = 3.70
 $\times 10^{-4}$ = 3.70
 $\times 10^{-4}$ = 3.70
 $\times 1$ we obtain
 $-\log K_w = -\log {\{H_3O^+ \mid [OH^- \}}$
 $-\log [H_3O^+] - \log [OH^-] = -\log [10^{-14}$
 $8K_w = pH + pOH = 14$
 $8K_w = pH + pOH = 14$

We can distinguish acidic, neutral and basic aqueous

solutions by the relative values of the H_3O^+ and OH⁻

conce log $K_w = -\log {\frac{1}{2}[H_3O^+]}\text{ [OH}^{-1}]$
 $-\log [H_3O^+] - \log [OH^-] = -\log [OH^{-1}4$
 $pH + pOH = 14$

of $pH + pOH = 14$

and istinguish acidic, neutral and basic aqueous

matrations.

https://index.com/index.com/index.com/index.com/index.com/in **STUDY MATERIAL: CHEMISTRY**
 $K_w = -\log {\{[H_3O^+]\} [OH^-]\}}$
 $g[H_3O^+]-\log {[OH^-]} = -\log 10^{-14}$
 $+ \gamma OH = 14$

stinguish acidic, neutral and basic aqueous

by the relative values of the H₃O⁺ and OH⁻
 H_3O^+ = [M₃O⁺] = [OH⁻ **STOFF MATERIAL:** CHENDISTRE

Dobtain
 $-\log K_w = -\log [H_3O^+]$ [OH⁻¹]
 $=-\log [H_3O^+] - \log [OH^-] = -\log 10^{-14}$
 $\neq pH + pOH = 14$

can distinguish acidic, neutral and basic aqueous
 $\text{det}[H_3O^+] > [OH^-]$
 trid
 $\text{det}[H_3O^+] > [OH^-]$
 $\$ $K_w = -\log {\{[H_3O^+]\} [OH^-]\}$
 $\{H_3O^+]-\log [OH^-] = -\log 10^{-14}$

+ pOH = 14

+ pOH = 14

triputals acidic, neutral and basic aqueous

triputals acidic, neutral and basic aqueous
 O^+ $>[H_3O^+]=[OH^-]$
 O^+ $[-SO^{++}^-]$
 O^+ $[-O^{++}^-$

equation
$$
K_b = \frac{[M^+] [OH^-]}{[MOH]}
$$

Alternatively, if $C =$ initial concentration of base and α = degree of ionisation of base. The equilibrium constant can be written as

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

- **(61) (B).** $pH = -\log|H^+|$ $2 = -\log[H^+]$; $[H^+] = 1 \times 10^{-2}$
- **(62) (D).** In case of a conjugate acid–base pair, $K_a \times K_b = K_w$
Knowing one, the other can be obtained. It should be noted that a strong acid will have a weak conjugate base and vice-versa.

(63) (C). H₂O(
$$
\ell
$$
)+H₂O(ℓ) \rightleftharpoons H₃O⁺(aq)+OH⁻(aq)
acid base conjugate
acid base

(64) (A). Hydronium ion concentration in molarity is more conveniently expressed on a logarithmic scale known as the pH scale. The pH of a solution is defined as the negative logarithm to base 10 of the activity can be expressed in terms of concentration in molarity
of various species in equilibrium by the following
equation $K_b = \frac{[M^a][OH^a]}{[OH^a]}$
Alternatively, if C = initial concentration of base and
 $α = degree of inisation of base$. The equilibr

 $(a_{H⁺})$ of hydrogen ion.

- **(65) (C).** $pH = -\log[H^+] = -\log(3.8 \times 10^{-3}) = 2.42$
- **(66) (A).** A mixture of acetic acid and sodium acetate acts as a buffer with pH around 4.75.
- **(C).** In 0.1 mol dm⁻³ NH₄OH and 0.05 mol dm⁻³ HCl, total amount of HCl reacts with $NH₄OH$ to form $NH₄Cl$ and some $NH₄OH$ will be left unreacted. Thus, the resultant solution contains $NH₄Cl$ and $NH₄OH$ which will produce a buffer solution.
- **(68) (D).** The solutions which resist change in pH on dilution or with the addition of small amounts of acid or alkali are called buffer solutions.
	- **(69) (D).** The solubility of salts of weak acids like phosphates increases at lower pH. This is because at lower pH the concentration of the anion decreases due to its protonation. This in turn increase the solubility of the salt so that $K_{sp} = Q_{sp}$.

146

(70) (A).
$$
MX_4 \rightleftharpoons M^+ + 4X^-
$$

\nS 4s
\n $K_{sp} = s \times (4s)^4 = 256s^5$
\n(71) (B).
\n(a) Fe(OH)₃ \rightleftharpoons Fe³⁺ + 3OH⁻; $K_{sp} = 27s^4$
\nS 30
\nS 31
\nD
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(74) (A).
$$
A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}
$$

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

(75) (A). AgCN Ag+ + CN– ^Ksp = [Ag+] [CN–] = 6 × 10–17 Ni(OH)2 Ni2+ + 2OH– ^Ksp = [Ni2+] [OH–]² = 2 ×10–15 Let [Ag+] = S¹ , then [CN–] = S¹ Let [Ni2+] = S² , then [OH–] = 2S² S¹ 2 = 6 ×10–17, S¹ = 7.8 ×10–9 (S²) (2S²)² =2 ×10–15, S² =7.9 ×10–4 Ni(OH)² is more soluble than AgCN. **(76) (A).** RaSO⁴ Ra2+ + SO⁴ 2– ^Ksp = [Ra2+] [SO⁴ 2–] Concentration of SO⁴ 2– from Na2SO⁴ = 0.10 M ² 4 10 Ra in water Ba2+(aq) + SO⁴ 2 2 [Ba][SO] ^K [BaSO]

$$
Ra^{2+} = \frac{4 \times 10^{-11}}{0.10} = 4 \times 10^{-10} \,\mathrm{M}
$$
 (4)

(77) (A). The equilibrium between the undissolved solid and the ions in a saturated solution can be represented by the equation

BaSO₄ (s)
$$
\xrightarrow{\text{solution} \atop \text{in water}} Ba^{2+}(aq) + SO_4^{2-}(aq)
$$
.

The equilibrium constant is given by the equation

$$
K = \frac{[Ba^{2+}][SO_4^{2-}]}{[BaSO_4]}
$$

For a pure solid substance the concentration remains constant and we can write

 $K_{sp} = K [BaSO_4] = [Ba^{2+}] [SO_4^{2-}]$

We call K_{sp} the solubility product constant or simply solubility product.

- **(78) (B).** For precipitation, ionic product > solubility product. In all other options, ionic product is less than solubility product.
- **(79) (B).** From Le-Chatelier's principle that if we **increase** the concentration of anyone of the ions, it should combine with the ion of its **opposite** charge and some of the salt will be precipitated till once again $K_{sp}=Q_{sp}$.
Similarly, if the concentration of one of the ions is decreased, more salt will dissolve to increase the concentration of both the ions till once again
- **K**_{sp}= Q_{sp} **. (80) (C).** Solubility of a salt in a solution containing common ion is less than the solubility of the salt in pure water. More is the concentration of the common ion, lesser would be the solubility of salt. Thus, the order of solubility of AgCl in various solutions would be $s_1 > s_3 > s_2 > s_4$ $K_{sp} = Q_{sp}$

Solubility of a salt in a solution containing common

Solubility of a salt in a solution containing common

ion is less than the solubility of the salt in pure water.

More is the concentration of the common 25 than the solubility of the salt in pure water.

the concentration of the common ion, lesser

the be solubility of salt. Thus, the order of

the be the solubility of salt. Thus, the order of

ivy of AgCl in various so

EXERCISE-2

 (1) **(1) (D).** There are an infinite number of equilibrium positions at a given temperature depending on the initial concentrations.

> The size of K has no relationship to the speed at which equilibrium is achieved.

A pure liquid or solid is never included in the equilibrium expression.

(2) (A).
$$
2NO_2(g) \rightleftharpoons N_2O_4(g)
$$
, $\Delta n = 1 - 2 = -1$

$$
K_p = K_c (RT)^{\Delta n}
$$
; $\frac{K_p}{K_c} = (RT)^{-1}$

(3) (B). In presence of 10^{-10} M HCl, the common ion effect exerted by H^+ ions contributed by HCl on dissociation of water would be negligible. Thus, the $[H^+]$ coming from water would be about 1×10^{-7} M. **EXERCISE-2**
 EXERCISE-2
 (D). There are an infinite number of equilibrium positions

at a given temperature depending on the initial

concentrations.

The size of K has no relationship to the speed at

which equilibr e size of K ans no eraturosmy to the speed at

inch equilibrium is achieved.

pure liquid or solid is never included in the

uilibrium expression.
 $\overline{SO}_2(g) \rightleftharpoons N_2O_4(g), \Delta n = 1 - 2 = -1$
 \Rightarrow $= K_c (RT)^{\Delta n}$; $\frac{K_p}{K_c} = (RT)^{-$

$$
\therefore \quad [H^+]_{\text{Total}} = [H^+]_{\text{H}_2\text{O}} + [H^+]_{\text{HCl}}
$$

$$
= 10^{-7} + 10^{-10} \approx 10^{-7} \text{M}
$$

$$
\therefore \quad pH = 7
$$

(4) (C). For Fe³⁺(aq) + 3OH⁻(aq)
$$
\rightleftharpoons
$$
 Fe(OH)₃(s)

$$
K_c = \frac{1}{[Fe^{3+}][OH^-]^3}
$$

(5) (A). $pH = 10$, $[H^+] = 1 \times 10^{-10} M$ \therefore $[OH^-] = 1 \times 10^{-4} M$

EXERCISE-2

ere are an infinite number of equilibrium positions

given temperature depending on the initial

ecentrations.

size of K has no relationship to the speed at

chequilibrium is achieved.

librium expression.
 (6) (C). Let the solubility of $Ni(OH)_2$ be equal to S. Dissolution of S mol/L of Ni(OH)₂ provides S mol/L, of Ni^{2+} and 2S mol/L of OH^{-} , but the total concentration of $OH^- = (0.10 + 2S)$ mol/L because the solution already contains 0.10 mol/L of OH– from NaOH.

$$
K_{sp} = 2.0 \times 10^{-15} = [Ni^{2+}] [OH^-]^2 = (S) (0.10 + 2S)^2
$$

As K_{sp} is small, 2S < 0.10, thus, (0.10 + 2S) ≈ 0.10
Hence, 2.0 × 10⁻¹⁵ = S (0.10)²
S = 2.0 × 10⁻¹³ M = [Ni²⁺]

(7) (C). Both (A) & (B) are correct for the equation,

$$
K = e^{-\Delta G^{\Theta}/RT}.
$$

EXERVING

Both (A) & (B) are correct for the equation,
 $K = e^{-\Delta G^{\Theta}/RT}$.
 K^{Θ} are correct for the equation,
 $K = e^{-\Delta G^{\Theta}/RT}$.
 $\Delta H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

The dissociation reaction of H_2O being endothermic,
 (8) (D). $2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ The dissociation reaction of H_2O being endothermic, (18) with the increase of temperature, equilibrium constant (K_w) increases, as reaction is favoured in the forward direction. Thus, concentration of H_3O^+ as well as OH– increases and pH and pOH decreases with increase of temperature. **(3) (A)** $K_e = \frac{[PC1_3][C1_2]}{[PC1_5]} = \frac{2.1 \times 2.1}{1.9} = 2.32$
 (3) (A) All the three statements are correct for the equation,

(8) **(A)** $K_e = \frac{[PC1_3][C1_2]}{[PC1_5]} = \frac{2.1 \times 2.1}{1.9} = 2.32$

(10) All the three statemen

(9) (A).
$$
K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{2.1 \times 2.1}{1.9} = 2.32
$$

(10) (D). All the three statements are correct.

(11) **(D).**
$$
K_p = K_c (RT)^{\Delta n}
$$
; $\Delta n = 2 - 3 = -1$
T = 350 K, R = 0.083 bar L K⁻¹ mol⁻¹

$$
K_c = \frac{K_p}{(RT)^{\Delta n}}
$$
 (19) (A). $K_p = K_p$

 (12)

$$
K_c = \frac{3 \times 10^{10} \text{ bar}^{-1}}{(0.083 \text{ L bar K}^{-1}\text{mol}^{-1} \times 350 \text{ K})^{-1}} = 8.715 \times 10^{11} \text{ L mol}^{-1}
$$

\n**(B).** K_w at 90°C = [H₃O⁺][OH⁻]
\n= 10⁻⁶ × 10⁻⁶ = 10⁻¹² (2

(5) (A) $K = \frac{C^{3/2}N!}{(R!)^{2m}}$ **(A)** $K_c = \frac{C^{3/2}N!}{(R!)^{2m}}$

(13) **(C)**, A solid salt of the general formula particular constant in the general formula particular constant $\frac{C}{R}$ intermus constant $\frac{C}{R}$ inte solubility S in equilibrium with its saturated solution may be represented by the equation $M_X X_y(s) \rightarrow xM^{p+}(aq) + yX^{q-}(aq)$ (where $x \times p^+ = y \times q^-$) And its solubility product constant is given by $K_{\rm sn}$ $=[M^{p+}]^x [X^{q-}]^y = (xS)^x (yS)^y$ $K_{sp} = x^x \cdot y^y \cdot S^{(x+y)}$ $K_e = \frac{K_p}{(RT)^{\Delta n}}$
 $K_e = \frac{3 \times 10^{10} \text{ bar}^{-1}}{(0.083 \text{ L bar K}^{-1} \text{mol}^{-1} \times 350 \text{ K})^{-1}}$
 $K_e = \frac{3 \times 10^{10} \text{ bar}^{-1}}{(0.083 \text{ L bar K}^{-1} \text{mol}^{-1} \times 350 \text{ K})^{-1}}$
 $K_e = \frac{0.0821 \text{ L}}{0.0821 \text{ L}}$
 $K_e = \frac{0.0821 \text{ L}}{0.0821 \text{ L}}$
 K_e x_p = 167 bar

3x 10¹⁰ bar⁻¹

3x 10¹⁰ bar⁻¹

3x 10¹⁰ bar⁻¹

3x 10¹⁰ bar⁻¹
 0^{-1} x 350 K)⁻¹
 0^{-1} x 350 K)⁻¹
 recetations are correct.
 $x_1y_1x_2 = 0.0821$, $x_1y_2 = 0.0821$, $x_1y_1y_2 = 0.0821$, $x_1y_2y_2 = 0.0821$, $x_1y_2y_$ $K_p = 167 \text{ bar}$
 $K_c = \frac{167 \text{ bar}}{0.0821 \text{ L bar} \cdot \text{K}^{-1} \text{mol}^{-1} \times 1073 \text{K}}$
 $= 1.896 \text{ mG/L}^{-1}$

(20) (D). Exothermic reaction, decrease in number contration of reactants.

(21) (A). Nucleophiles are Lewis bases while el correct.
 x^{-1}
 $= -1$.
 x^{-1}
 $= -1$.
 $\left(\frac{1H1_1^2}{R_1} - \frac{1}{R_2}\right) = \frac{(0.40)^2}{(0.10) \times (0.20)} = 8.0$

(19) $(\Delta_k K_p = K_q(RT)^{\Delta n}; \Delta n = 1$
 $K_p = 167 \text{ bar},$
 $K_q = \frac{0.68211. \text{ bar } K^{-1}m0^{-1} \times 1073 \text{ K}}{100 \text{ bar}}$

(20) (D). Excharg x 2.1

x 2.1 = 2.32

measured at some a tracheration to the contentrations were

are corrected

stage of the reaction is given by,
 $x = 3$

and $x = 1$
 $x = 1$

(19) (Δ) x 8.715 x 10¹¹ L mol⁻¹

y at 90°C [H j O'][OH⁻¹]

= 10⁻⁶ x 10⁻⁵ c [H j O'][OH⁻¹]

= 10⁻⁶ x 10⁻⁵ e 10⁻¹²

solid salt of the general formula M^p⁺ x^q⁺ with molar

solid salt of the general formula M^p K_e $= \frac{N_p}{(R\pi)^{4M}}$
 $K_e = \frac{3 \times 10^{10} \text{ bar}^{-1}}{(0.083 \text{ L bar K}^{-1} \text{mol}^{-1} \times 350 \text{ K})^{-1}}$
 $K_e = \frac{3 \times 10^{10} \text{ bar}^{-1}}{(0.083 \text{ L bar K}^{-1} \text{mol}^{-1} \times 350 \text{ K})^{-1}}$
 $K_e = \frac{3 \times 10^{11} \text{ km}^{-1}}{(0.083 \text{ L bar K}^{-1} \text{mol}^{-1} \times 350 \text{ K})^{-1}}$ RT)^{2na}
 $x = \frac{167 \text{ bar}}{3 \times 10^{10} \text{ bar}} = 10^{-6} \text{ at } \text{K}_{\text{c}} = \frac{167 \text{ bar}}{0.0821 \text{ L bar K}^{-1} \text{mol}^{-1} \times 1073 \text{K}}$
 $\times 10^{-4} \text{ J of } 10^{11} \text{cm}^{-1}$
 $\times 10^{-1} \text{ J of } 10^{-4} \text{cm}^{-1} \times 350 \text{ K}^{-1}$
 $\times 10^{-4} \text{ J of } 60^{-4} \text{ J of }$ 50.000 (19) (A) $K_p = K_q(RT)^{6n_1}$ (0.003 (1-2NH₂) (0.10) (19) (A) $K_p = K_q(RT)^{6n_1}$ (0.003 (1.10) = ((KT)⁴⁶

((BT)⁴⁶

((BT)⁴⁶ (0.82) L bar K⁻¹mol⁻¹ × 1073K

(19) (A) $K_p = 167$ bar,

(19) (A) $K_p = 167$ bar,
 K 2 + 10⁻⁶ × 10⁻¹ Lmo⁻¹ Lmo-1

9.9⁻¹ C=[H₃O⁺](OH⁻¹ Leo (**A**) Alto-line increase in concentration of reactants.

3.3 kg/s in equilibrium with its staturated solution

(2) (**A**) Methoply is visible are Lewis b

$$
S^{(x+y)} = \frac{K_{sp}}{x^x \cdot y^y}; \quad S = \left(\frac{K_{sp}}{x^x \cdot y^y}\right)^{1/x+y}
$$
 (26) (A).
(i) Increase in concentration of forward reaction.

(14) (C)
$$
a+b \rightleftharpoons c+d
$$

\n $x \times 2x \times 2x$
\n $2x \times 2x \times 4x^2$

$$
K_c = \frac{2x \times 2x}{x \times x} = \frac{4x^2}{x^2} \Rightarrow K_c = 4
$$

(15) (B). PCl₅ (g) \rightleftharpoons PCl₃(g) + Cl₂ (g); K_c = 8.3 × 10⁻³. For $\text{PCl}_3(g)$ + $\text{Cl}_2(g) \rightleftharpoons \text{PCl}_5(g)$; $K'_c = \frac{1}{1} = 120.48$ H_2S $\frac{1}{2}$ = $\frac{1}{2}$ = 120.48

$$
\lambda_c' = \frac{1}{8.3 \times 10^{-3}} = 120.48
$$

(16) (D). Sulphuric acid :

- It is a strong acid in the first dissociation step (is very large)
- It is a weak acid in the second step.
- **(17) (D).** Production of ammonia according to the reaction, N_2 $(g) + 3H_2(g) \rightleftharpoons 2NH_3(g); \Delta H = -92.38 \text{ kJ mol}^{-1}$ is an exothermic process. According to LeChatelier's principle, raising the temperature shifts the equilibrium to left and decreases the equilibrium concentration

of ammonia. In other words, low temperature is favourable for high yield of ammonia, but practically very low temperatures slow down the reaction and thus a catalyst is used.

EXERCISE AND CONSIDERATE CHEATERT (A) $K = C^{-\alpha G^O / RT}$

Hoth (A) & (B) are correct for the equation,
 $K = C^{-\alpha G^O / RT}$

Hoth (A) are correct for the equation,
 $K = C^{-\alpha G^O / RT}$

The dissociation reaction of H₂O being end **(O.B. SOLUTIONS**) **STUDYMATERIAL: CHEMISTRY**

(G^O/RT

(G^O/RT) are correct for the equation,
 G^Q/RT for ammonia. In other words, low temperature is
 \cos^Q/RT for a $\sin^Q/2$ for a starting the properties of the posit **EXECUTE ANTERIAL: CHEMISTI

AND (A) & (B) are correct for the equation, of ammonia. In other words, low temperature
** $= e^{-AG^0/RT}$ **.
** $O(t) \rightleftharpoons H_3O^+(ar)$ **
** $O(t) \rightleftharpoons H_3O^+(ar)$ **

(D)** $O(t) \rightleftharpoons H_3O^+(ar)$ **

(D) O(t) \rightleftharpoons H_3O^+(ar)** correct for the equation,

a causal for all piveld of ammonia, but
 $\frac{1}{2}$ for all piveld of ammonia, but
 $\frac{1}{2}$ for $\frac{1}{2}$ for = $e^{-\Delta G^{\circ}/R}$

= $e^{-\Delta G^{\circ}/R}$. However low temperatures slow down the reaction and

the increase of lemperatures slow down the reaction and

the increase of lemperature equilibrium constant H₂ (e) H₂ (e) H₂ (e) H **EXECUTE AND CONSIDENT CONSIDENT SOLUTIONS**

Both (A) & (B) are correct for the equation,
 $K = \cos^{-0.05}$ (R) $K = \cos^{-0.05}$ (R) $\cos^{-0.05}$ (B) $\sin^{-0.05}$ (C) $\cos^{-0.05}$ (Fig. Consider the guassitus commonia. but president in **EXECUTE ANTIFORM CONSIDERATE. CHEATISTRY**

(A) & (B) are correct for the equation, of ammonia. In other words, low temperature is
 $-\Delta G^0/RT$
 $(\Delta) \Delta G^0/RT$

(A) $\Delta G^0/RT$

(A) $\Delta G^0/RT$

(A) $\Delta G^0/RT$

(A) $\Delta G^0/RT$

(A) are correct for the equation,

favourable for high yield of ammonia, but practical!

faq) + OH⁻ (aq)

favourable for high yield of ammonia, but practical!

very low temperatures slow down the reaction and

action of H₂ **(18) (C).** Consider the gaseous reaction of H_2 with I_2 : $H_2(g) + I_2(g) \rightleftharpoons 2HI(g), K_c = 57.0$ at 700 K Suppose we have molar concentrations $[H_2]_t = 0.1 M, [I_2]_t = 0.20 M$ and $[HI]_t = 0.40$ M (the subscript t on the concentration symbols means that the concentrations were measured at some arbitrary time t, not necessarily at **STUDY MATERIAL: CHEMISTRY**

ammonia. In other words, low temperature is

vourable for high yield of ammonia, but practically

ty low temperatures slow down the reaction and

as a catalyst is used.

usata a state of H₂ **STUDY MATERIAL: CHEMISTRY**

of ammonia. In other words, low temperature is

favourable for high yield of ammonia, but practically

devery low temperatures slow down the reaction and

thus a catalyst is used.

Consider th **STUDY MATERIAL: CHEMISTRY**

mmonia. In other words, low temperature is

urable for high yield of ammonia, but practically

low temperatures slow down the reaction and

a catalyst is used.

sider the gaseous reaction of H wourable for high yield of ammonia, but practically
rourable for high yield of ammonia, but practically
ry low temperatures slow down the reaction and
us a catalyst is used.
 $\frac{1}{2}$ (g) + $\frac{1}{2}$ (g) $\frac{1}{2}$ (g) $k_c =$ **STUDY MATERIAL: CHEMISTRY**

mmonia. In other words, low temperature is

uurable for high yield of ammonia, but practically

low temperatures slow down the reaction and

a catalyst is used.

sider the gaseous reaction of

equilibrium). Thus, the reaction quotient, Q_c at this
stage of the reaction is given by,

$$
Q_c = \frac{[H I]_t^2}{2(0.40)^2} = \frac{(0.40)^2}{2(0.40)^2} = 8.0
$$

= 350 K, R = 0.083 bar L K⁻¹ mol⁻¹
\n
$$
c = \frac{K_p}{(RT)^{\Delta n}}
$$
\n(19) (A) K_p = K_c(RT)^{Δn}, Δn = 1
\n
$$
K_c = \frac{167 \text{ bar}}{0.0821 \text{ L bar K}^{-1} \text{mol}^{-1} \times 1073 \text{K}}
$$
\n= 1.896 mol L⁻¹
\n= 1.896 mol L⁻¹

$$
K_c = \frac{167 \text{ bar}}{0.0821 \text{ L bar K}^{-1} \text{mol}^{-1} \times 1073 \text{K}}
$$

= 1.896 mol L⁻¹

- **(20) (D).** Exothermic reaction, decrease in number of moles, increase in concentration of reactants.
- **(21) (A).** Nucleophiles are Lewis bases while electrophiles are Lewis acids.
- **(22) (B).** Decrease in number of moles in the reaction makes it favourable at high pressure.

(23) (B). Acidic strength
$$
\propto \sqrt{K_a}
$$

(24) (D). AgCl + 2NH₃ \rightarrow [Ag(NH₃)₂]Cl (Soluble)

Thus, AgCl is most soluble in aqueous ammonia.

(25) **(D).** CH₃COONa + H₂O
$$
\rightleftharpoons
$$
 CH₃COOH + NaOH
NaOH \rightleftharpoons Na⁺ + OH⁻ (Basic solution)

$$
^{+y} \t(26) \t(A).
$$

 (19)

- $S = \frac{S_P}{S_V}$ (i) Increase in concentration of reactants favours forward reaction. Interestion of reactants.

Nucleophiles are Lewis bases while electrophiles are

Lewis acids.

Decrease in number of moles in the reaction makes it

favourable at high pressure.

Acidic strength $\propto \sqrt{K_a}$

AgCl + 2NH₃ eopnies are Lewis bases while electrophiles are

s acids.

s acids

case in number of moles in the reaction makes it

lic strength $\propto \sqrt{K_a}$
 $1+2NH_3 \rightarrow [Ag(NH_3)_2]Cl$

(Soluble)

is soluble in aqueous ammonia.

COONa + H₂ Acidic strength $\propto \sqrt{K_a}$
AgCl + 2NH₃ \rightarrow [Ag(NH₃)₂]Cl
(Soluble)
Thus, AgCl is most soluble in aqueous ammonia.
CH₃COONa + H₂O \rightarrow CH₃COOH + NaOH
NaOH \rightleftharpoons Na⁺ + OH⁻ (Basic solution)
Increase in co NH₃ \rightarrow V_a

(Soluble)

Cl is most soluble in aqueous ammonia.

NNa+H₂O \rightleftharpoons CH₃COOH + NaOH

Na+H₂O \rightleftharpoons CH₃COOH + NaOH

in concentration of reactants favours

in concentration of products favours

aca
	- (ii) Increase in concentration of products favours backward reaction.
	- (iii) Removal of CO will favour backward reaction (decrease in concentration of reactants).
	- (iv) Removal of $CH₃OH$ will favour forward reaction (increase in concentration of products).
	- **(27) (A).** For the reaction:

$$
H_2S \rightleftharpoons H^+ + HS^-; K_{a_1}
$$

$$
K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}
$$
 ... (i)

For the reaction: $HS^- \rightleftharpoons H^+ + S^{2-}$; K_{a_2}

$$
\zeta_{a_2} = \frac{[H^+][S^2^-]}{[HS^-]}
$$
 ... (ii)

For the reaction: $H_2S \rightleftharpoons 2H^+ + S^{2-}$; K_{a_3}

$$
K_{a_3} = \frac{[H^+]^2 [S^{2-}]}{[H_2S]}
$$
 ... (iii)

By multiplying eqn. (i) and eqn. (ii), we get

$$
K_{a_1} \times K_{a_2} = \frac{[H^+] [HS^-]}{[H_2S]} \times \frac{[H^+] [S^{2-}]}{[HS^-]}
$$

$$
= \frac{[H^+]^2 [S^{2-}]}{[H_2S]} = K_{a_3}
$$
(3)

(28) (A). $\frac{1}{2}$ H₂ (g) + $\frac{1}{2}$ I₂ (g) \implies HI (g) \therefore K_c = 5 ...(i) $\frac{1}{2}$ H₂(g) + $\frac{1}{2}$ I₂(g) \implies HI(g) ; K_c = 5 ...(i) $1 \t\t (38)$ $\frac{1}{2}I_2(g) \rightleftharpoons H1(g)$; K_c = 5 ...(i) Multiply eqn (i) by 2,

 $H_2(g) + I_2(g) \rightleftharpoons 2HI(g); K_c = (5)^2$...(ii) Now, reverse the reaction

$$
2\text{HI} \text{ (g)} \rightleftharpoons \text{H}_2 \text{ (g)} + \text{I}_2 \text{ (g)} ; \text{ K}_c = \frac{1}{(5)^2}
$$

$$
\therefore K_c = \frac{1}{25} = 0.04
$$

(29) **(B).** N₂ (g) + 3H₂ (g) \implies 2NH₃ (g)
At eq. 1-x 3-3x 2x
4-2x=3
x= 0.5

$$
K_p
$$
 for dissociation of NH₃ = $\frac{(0.5) (1.5)^3}{1}$ 2 × 10⁻⁷1

At
$$
t = 0
$$
 1
\nAt eq. $1-x$ 3 – 3x 2x
\n $x = 0.5$
\nAt eq. $1-x$ 3 – 3x 2x
\n $x = 0.5$
\nK_p for dissociation of NH₃ = $\frac{(0.5)(1.5)^3}{1}$
\n**(30)** (D) ΔG° = -RT ln K
\nIf ln K = -ve then only
\n ΔG° +ve (or > 0). Thus, K < 1
\n ΔG° +ve (or > 0). Thus, K < 1
\n**(31)** (A), pH = $\frac{1}{2}$ [pK_w + pK_n - pK_b]
\n $pH = 7 + \frac{1}{2}$ (5.76 – 5.25) = 7.255
\n**(32)** (B), H1 = -log [H⁺]
\n 50 ₄
\n 50 ₂²
\n $pH = 14$ – 10g [H⁺]
\n 14 – 4.5229 = 4.7
\n $pH = 14$ – pH,
\n $pH = 12$ – 4.5229
\n $pH = 14$ – pH,
\n $pH = 12$ – 4.5229
\n $pH = 14$ – pH,
\n $pH = 12$ – pH,
\n $pH = 14$ – pH,
\n $pH = 14$ – pH,
\n p

IDENTIFY
\nK₃₃ =
$$
\frac{[H^+]^2 [S^2^-]}{[H_2S]}
$$
 ...(ii)
\nBy multiplying eqn. (i) and eqn. (ii), we get
\n $K_{a_1} \times K_{a_2} = \frac{[H^+][HS^-]}{[H_2S]}$ (*i* and (iii) $A(t=0, 0.5 - 0.5 - x.2x$
\n $K_{a_1} \times K_{a_2} = \frac{[H^+][HS^-]}{[HS^-]} \times \frac{[H^+][S^2^-]}{[HS^-]}$
\n $= \frac{[H^+]^2 [S^2^-]}{[HS^-]} = K_{a_3}$
\n $= \frac{[H^+]^2 [S^-]}{[HS^-]} = K_{a_3}$
\n $= \frac{[H^$

$$
= \frac{(0.5) (1.5)^3}{1}
$$
 (40) (C). 50 lit 10⁻⁵ mole AgCl + 50 lit of 2 × 10⁻⁷ M HBr

(II)
$$
K_c = \frac{0.01 \times 0.01}{[0.01]^2} = 1
$$

\n(III) $K_c = \frac{[0.01]^2}{0.01 \times [0.01]^3} = 100$
\n(C) 50 lit 10⁻⁵ mole AgCl + 50 lit of
\n 2×10^{-7} M HBr
\n $[Ag^+] = \frac{10^{-5}}{100}$
\n $K = 10^{-7} \times 10^{-7} = 10^{-14} < K_{sp} = 5 \times 10^{-3}$
\nSo, no, precipitation and $[Ag^+] = 10^{-7}$ M
\nE XERCISE-3
\nI. We can write, $2HI(g) \implies H_2(g) + I_2(g)$
\ninitial pressure : 0.2 atm 0 0
\nEquilibrium partial pressure : 0.04 atm
\n $(0.2 - 0.04)$ $\frac{(0.2 - 0.04)}{2}$
\n= 0.08 0.08
\nThen, $K_p = \frac{[H_2][I_2]}{[H_1]^2} = \frac{0.08 \times 0.08}{(0.04)^2} = 4.0$
\n3. $AB(g) \implies A(g) + B(g)$
\nInitial 1 mole 0 0
\n4t equilibrium 1-0.5 0.5 0.5

EXERCISE-3

(1) 4. We can write, $2H\text{I}(g) \rightleftharpoons H_2(g)$
Initial pressure: 0.2 atm $\rightleftharpoons H_2(g) + \qquad I_2(g)$ $I_2(g)$ Initial pressure : Equilibrium partial pressure : 0.04 atm

$$
\begin{array}{ccc}\n(0.2 - 0.04) & (0.2 - 0.04) \\
2 & 2 \\
= 0.08 & 0.08\n\end{array}
$$

Then,
$$
K_p = \frac{[H_2][I_2]}{[H_2]^2} = \frac{0.08 \times 0.08}{(0.04)^2} = 4.0
$$

EXERCISE-3

255

255

255

255

10

25 1 initial pressure : 0.2 atm

26 0

26 1 initial pressure : 0.04 atm

26 0.08

26 0.08

27 0.08

27 0.08

27 0.08

29

27 1 111²

27 0.08×0.08

29

27 0.04²

27 0.04²

29

29 $\frac{10^{-5}}{0.01 \times [0.01]^3}$ - 100
 $\frac{10^{-5}}{100}$
 $\frac{10^{-5}}{100}$
 $\frac{10^{-5}}{100}$
 $\frac{10^{-5}}{100}$
 $\frac{10^{-5}}{100}$
 $\frac{10^{-5}}{100}$
 EXERCISE-3

Trie, $\frac{2HI(g)}{2}$ \Rightarrow H₂(g) + I₂(g)
 $\frac{10^{-7}M}{2}$
 $\frac{10^{-7}M}{2$ 11³
 11^{3}
 11^{3}
 11^{3}
 11^{4} < K_{sp} = 5×10⁻³

and [Ag⁺] = 10⁻⁷ M
 CISE-3

(2) $\implies H_2(g) + I_2(g)$

on 0

on 0

on 0

on 0

<u>0

(0.04</u> atm

<u>0

(0.04</u> atm

<u>2</u>

(0.04 atm

(0.04 atm

(0.04 atm

(0.04 atm **(2) 3.** AB(g) \rightleftharpoons A(g)+B(g) Initial 1 mole 0 0 At equilibrium 1–0.5 0.5 0.5 Total number of moles = 1.5 Equilibrium pressure = P

$$
\therefore K_p = \frac{\left(\frac{0.5}{1.5}P\right)^2}{\left(\frac{0.5}{1.5}\right)P} = \frac{0.5}{1.5}P = \frac{1}{3}P \qquad \therefore P = 3K_p
$$

(3) 25. Since the number of moles in a gas under the same conditions of temperature and pressure are proportional to volumes (Avogadro's Law), the volumes in mL of gases may be used instead of concentrations in the case of reversible reactions in which there is no change in the number of molecules of reactants and products. $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{1}{2}$
 $\frac{2}{2}$
 $\frac{2}{2}$
 $\frac{1}{2}$
 $\frac{2}{2}$
 $\frac{2}{2}$
 $\frac{1}{2}$
 $\frac{2}{2}$
 $\frac{1}{2}$
 Pressure is

note the number of moles in a gas under the same

ions of temperature and pressure are proportional to

es (Avogadro's Law), the volumes in mL of gases

e used instead of concentrations in the case of
 0.112 **EXERCISE AND COREMATE SOLUTIONS**

(S) 2. Initial mole $1_2 \rightleftharpoons 21$
 $K_p = \frac{\left(\frac{0.5}{1.5}\right)^2}{\left(\frac{0.5}{1.5}\right)^p} = \frac{0.5}{1.5}P = \frac{1}{3}P$ $\therefore P = 3K_p$

A requilibrium $1 - x$

Total moles $1 - x - x = 1 + x$

Total moles $-1 - x + 2x = 1 +$ **EXECUTE ALTERAL CHE SOLUTIONS**

(S) 2. Initial mole $I_2 \rightleftarrow \frac{1}{5}$ or $I_1 \rightleftarrow \frac{3}{1.5}$
 $\frac{1}{2} \leftarrow \frac{3}{1.5}$
 $\frac{1}{2} \leftarrow \frac{1}{2}$
 Pressure is proportional to nun

the number of moles in a gas under the same

of temperature and pressure are proportional to

Avogadro's Law), the volumes in mL of gases

eed instead of concentrations in the case of

mol content and product is the same of moles in a gas under the same

to the member of moles in a gas under the same

1 stated value of press

1 stated and products of molecules of press

to denote the same in mL of gases
 2 so of emperature and pressure are proportional value of pressure are the number of moles in a gas under the same

so (Avogadro's Law), the volumes in mL of gases

e used instead of concentrations in the case of
 $0.774 = 1$

Given : $2x = 30.8$ mL So, $x = 15.4$ mL

Vol. of H_2 at equilibrium = $(25-15.4)$ = 9.6 mL. Vol. of I_2 at equilibrium = $(18 - 15.4) = 2.6$ mL.

$$
K_c = \frac{[H I]^2}{[H_2][I_2]} = \frac{(30.8)^2}{9.6 \times 2.6} = 38.0
$$

Consider the equation : $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)$ Let the degree of dissociation be x

$$
K_c = \frac{x^2}{4(1-x)^2} \qquad \left\{ K_c = \frac{1}{K_c} = \frac{1}{38} \right\}
$$

$$
\frac{1}{38} = \frac{x^2}{4(1-x)^2} \qquad \text{or} \qquad \frac{1}{6.1644} = \frac{x}{2(1-x)}
$$

or $x = 0.245 \text{ i.e.,} \qquad 24.5\% \text{ dissociated.}$

(4) 1. NH₃(g)
$$
\rightleftharpoons \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)
$$
. Total moles

Initial moles $t = 0$

1 0 0 1 Moles equilibrium

 $1-\alpha$ $\alpha/2$ $3\alpha/2$ $1+\alpha$ Partial pressure

$$
\left\{\frac{1-\alpha}{1+\alpha}\right\}p \qquad \left\{\frac{\alpha}{2(1+\alpha)}\right\}p \qquad \left\{\frac{3\alpha}{2(1+\alpha)}\right\}p \qquad (7) \quad \textbf{20. & (8) 8.} \\
\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{H}\text{S}^- \\
\text{IN} \quad \frac{1^{1/2}\text{FH} \quad \frac{1^{3/2}}{2}}{}
$$

$$
K_{p} = \frac{[N_{2}]^{1/2}[H_{2}]^{3/2}}{[NH_{3}]}
$$
 $K_{a_{1}} = \frac{[H]}{[H_{a_{1}}]}$

$$
= \frac{\left[\frac{\alpha}{2(1+\alpha)}p\right]^{1/2}\left[\frac{3\alpha}{2(1+\alpha)}p\right]^{3/2}}{\left[\frac{1-\alpha}{1+\alpha}p\right]} = \frac{p\alpha^2\sqrt{27}}{4(1-\alpha^2)}
$$

 $K_{a_2} = \frac{[H^+][S^{2-}]}{[HS^-]}$
\nMultiplying both the equations

Solving for ' α ' we get : α = K_a p Due K_a We know, $K_p = K_c (RT)^{\Delta n}$; 78.1 = $K_c (0.0821 \times 673)^1$ $K_c = 1.4\dot{1}3$ moles litre⁻¹

1. CHEARNING
 1.5 P $\frac{1}{1.5}$ **CD 1.5 EXECUTIONS**
 1.5 0.5 EXECUTIONS
 1.5 CHEARNING
 1.5 P $\frac{1}{1.5}$ **CD 1.5 EXECUTIONS**
 1.5 EXECUTIONS
 1.5 EXECUTIONS
 1.5 EXECUTIONS
 1.5 EXECU (O.B. SOLUTIONS STUDYMATERIAL: CH
 $\left(\frac{0.5}{0.5}\right)^2 = \frac{0.5}{1.5}P = \frac{1}{3}P$.: $P = 3K_p$ At equilibrium $1 - x$ 2x

Total moles $= 1 - x + 2x = 1 + x$

Pressure is proportional to number of moles,

number of moles in a gas under **EXECUTIONS** STUDYMATERIAL:

TRINKS (5) 2. $I_2 \rightleftharpoons 2I^-$
 $\left(\frac{0.5}{1.5}\right)^2 = \frac{0.5}{1.5}P = \frac{1}{3}P$ $\therefore P = 3K_p$

The and $P = 3K_p$

Total moles $I = 1 - x + 2x = 1 + x$

Total moles $I = 1 - x + 2x = 1 + x$

Pressure is proportional to **(5) 2.** $I_2 \rightleftharpoons$ \rightleftharpoons 21 Initial mole 1 At equilibrium $1-x$ 2x Total moles = $1 - x + 2x = 1 + x$ Pressure is proportional to number of moles,

Q.B-SOLUTIONS
\nSTDDYMATERIAL: CHEMISTRY
\n(5) 2.
$$
1_2 \rightleftharpoons 21^-
$$

\nInitial mole $1 - x$
\n $1 - x$
\

$$
K_{p} = \frac{P_{\rm i}^{2}}{P_{\rm i2}} = \frac{\frac{3.2 \times 10^{2}}{1000} \frac{\text{m}^{3}}{\text{s}}}{\frac{3.2 \times 10^{2} \text{m} \cdot \text{s}}{\text{s}^{2} + \text{s}^{2}} \frac{\text{s}^{2}}{\text{m}^{2}} = \frac{4 \times^{2} P_{\rm i}}{(1 - \text{s}^{2})} = \frac{4 \cdot (0.51)^{2} \cdot 0.112}{(1 - 0.51^{2})}
$$

 $= 0.1575$ atm.

 $\times 2.6$ **(6) 28.** A^{2-} will react first and product is HA⁻ m mol of product = $1.00 \times 0.300 = 0.300$ $HA^- + H_2O \rightleftharpoons H_2A + OH^-$

at pH 9.00 which is equal to
$$
\frac{pKa_1 + pKa_2}{2}
$$

in a gas under the same

sesure are proportional to

volumes in mL of gases

volumes in mL of gases

there is no change in the

and products.

Fig. and products

and products
 $K_p = \frac{P_t^2}{P_1} = \frac{\frac{2 \times 2 \times 6^2}{81 + x \cdot 6} P^$ ies in a gas under the same

pressure are proportional to $\frac{112}{1} = \frac{1+x}{1}$ or $x = 0.51$

the volumes in mL of gases

centrations in the case of $\frac{0.112}{0.074} = \frac{1+x}{1}$ or $x = 0.51$

there is no change in the

sixte $(x - x)$ Number of A^{2-} initially present in the solution (g) $+ \frac{1}{2}H_2(g)$. Total moles
and HA⁻ is equal to 1 3×10^{3} at pH 10.33 the system is a buffer in which the ratio of A^{2–} FIA⁻¹_{lmitial} + [HA⁻¹_{lmitial} + [HA⁻¹_{lmitial} = [A²-¹_{lmitial} - [HA⁻¹]_{nitial} - [HA⁻¹] 0 1 [HA⁻¹]_{lmial} = 3.00 - 0.300 - 0.300 = 2.40 mmol

mmol of Na₂A = 3.00

3a/2 1 + α mmol of NaHA = 2.40

mmol of NaHA = 2.40

(7) **20. & (8) 8.**

 $K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}$ (i)

 $K_{a_2} = \frac{[H^+][S^2^-$ All A^{2-} are protonated as HA^{-} $= 0.300 \times 10.0 = 3.00$ m mol $\text{[HA^-]}_{initial} + \text{[HA^-]}_{formed} = \text{[A}^{2-}]_{initial} - \text{[HA^-]}_{formed}$
 $\text{[HA^-]}_{initial} = 3.00 - 0.300 - 0.300 = 2.40 \text{ m mol}$ m mol of $Na₂A = 3.00$ m mol of NaHA $= 2.40$ duct is HA⁻
= 0.300
 $\frac{(a_1 + pK_{a_2})}{2}$
t in the solution
er in which the ratio of A^{2-
 $\frac{(a_1 - pK_{a_2})^2}{2}$
= 2.40 m mol
 $\frac{(2 \times 3.00 + 2.40)}{0.30}$ = 28.00 ml} $\frac{1}{x^2} = \frac{4(0.31)^8 - 0.112}{(1 - 0.51^2)}$

and is HA⁻¹

and is HA⁻¹

and the solution

in which the ratio of A²⁻¹

and in which the ratio of A²⁻¹

and and 2.40 m mol

and 2.40 m mol

b 3.00 + 2.40

a 28.00 ml + H₂O \rightleftharpoons H₂A + OH

H 9.00 which is equal to $\frac{pKa_1 + pKa_2}{2}$

A² - are protonated as HA⁻

hber of A² - initially present in the solution

300 × 10.0 = 3.00 m mol

HA - is equal to 1

HA - is equal to 1

-0 which is equal to $\frac{pKa_1 + pK_{a_2}}{2}$

are protonated as HA⁻
 $\frac{pKa_1 + pK_{a_2}}{2}$

are protonated as HA⁻
 $\frac{33}{2}$ the system is a buffer in which the ratio of A²⁻

is equal to 1

is a tyric in a buffer in whi mber of A² initially present in the solution

300 × 10.0 = 3.00 m mol

H 10.33 the system is a buffer in which the ratio of A²

HA⁻ is equal to 1
 $T_{\text{initial}}^{\text{initial}} + [H\text{A}^-]_{\text{formed}}^{\text{initial}} = [A^2^-]_{\text{initial}}^{\text{initial}} - [H\text{A}^-]_{\text$ =3.00 mmol

e system is a buffer in which the ratio of A²⁻¹

[HA⁻]_{formed} = [A²⁻]_{initial} – [HA^{-]}_{formed}

1.9.00 – 0.300 – 0.300 = 2.40 mmol

A = 3.00

A = 2.40

of HCl required = $\frac{2 \times 3.00 + 2.40}{0.30}$ = 28 mitial ⁺ [ITA Jformed - [A Jinitial - [ITA Jformed

nitial ⁼ 1.00 - 0.300 - 0.300 = 2.40 mmol

of Na₂A = 3.00

of NaHA = 2.40

olume of HCl required = $\frac{2 \times 3.00 + 2.40}{0.30}$ = 28.00 ml
 (8) 8.
 $S \rightleftharpoons H^+ + HS^-$
 \pm [HA⁻]_{formed} = [A²⁻]_{initial} - [HA⁻]_{formed}
= 3.00 - 0.300 - 0.300 = 2.40 m mol
 $h_2A = 3.00$
ie of HCl required = $\frac{2 \times 3.00 + 2.40}{0.30}$ = 28.00 ml
 \pm
A = 2.40
if H = 2.40
if H⁺ HS⁻ 0.30
 \pm H⁺

Total volume of HCl required $=$ $\frac{0.30}{0.30}$ = 28.00 ml

 $(2(1+\alpha))^F$ $H_2S \rightleftharpoons H^+ + HS^-$

$$
K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}
$$
(i)

Further $HS^{-} = H^{+} + S^{2-}$

Total volume of HCl required
$$
= \frac{2 \times 3.00 + 2.40}{0.30} = 28.00 \text{ ml}
$$

\n
$$
\frac{3\alpha}{2(1+\alpha)}\left\{p\right\}
$$
\n(7) 20. & (8) 8.
\nH₂S \rightleftharpoons H⁺ + HS⁻
\n
$$
K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}
$$
\n
$$
K_{a_2} = \frac{[H^+][S^{2-}]}{[HS^-]}
$$
\n
$$
K_{a_2} = \frac{[H^+][S^{2-}]}{[HS^-]}
$$
\n
$$
multiplying both the equations
$$
\n
$$
\frac{1}{2} \frac{p}{K_p}\Big|^{1/2}
$$
\n
$$
K_{a_1} \times K_{a_2} = \frac{[H^+]^2[S^{2-}]}{[H_2S]}
$$
\nDue to common ion, the ionisation of H₂S is suppressed
\n(0.0821 × 673)¹

$$
K_{a_1} \times K_{a_2} = \frac{[H^+]^2 [S^{2-}]}{[H_2 S]}
$$

3 3 p ¹ 3α/2 1+α mmol of Nat¹₁A = 2.40

Total volume of HCl required = $\frac{2 \times 3.00 + 2.40}{0.30}$ = 28.0
 $\left(\frac{3\alpha}{2(1+\alpha)}\right)$ P **20. & (8) 8.**
 $H_2S \rightleftharpoons H^+ + HS^-$
 $K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}$

Further $HS^- = H^+ + S^2$ -
 0

1
 $\frac{1}{3\alpha/2}$
 $1+\alpha$
 $\frac{1}{2(1+\alpha)}$
 $1+\alpha$
 $\frac{1}{2(1+\alpha)}$
 $1+\alpha$
 $\frac{1}{2(1+\alpha)}$
 $\frac{1}{2(1+\alpha)}$
 $\left(\frac{3\alpha}{2(1+\alpha)}\right)$
 $\left(\frac{7}{2}\right)$
 $\left(\frac{3\alpha}{2(1+\alpha)}\right)$
 $\left(\frac{7}{2}\right)$
 $\left(\frac{3\alpha}{2(1+\alpha)}\right)$
 $\left(\frac{7}{2}\right)$
 $\left(\frac{3\$ **8.**
 \Rightarrow H⁺ + HS⁻
 $=$ $\frac{[H^+][HS^-]}{[H_2S]}$ (i)
 $HS^- = H^+ + S^{2-}$
 $= \frac{[H^+][S^{2-}]}{[HS^-]}$ (ii)

ing both the equations
 $a_1 \times K_{a_2} = \frac{[H^+]^2[S^{2-}]}{[H_2S]}$

ommon ion, the ionisation of H₂S is suppres Due to common ion, the ionisation of H_2S is suppressed and the $[H^+]$ in solution is due to the presence of 0.3 M HCl.

QUILIBRIUM	Q.B. SOLUTIONS	
$[S^{2-}] = \frac{K_{a_1} \times K_{a_2}[H_2S]}{[H^+]^2} = \frac{1.0 \times 10^{-7} \times 1.3 \times 10^{-13} \times (0.1)}{(0.3)^2}$	initial, M	0.
$= 1.44 \times 10^{-20} M$	changing, M	
Putting the value of $[S^{2-}]$ in Eq. (ii)	Substitute into solve for y.	

Putting the value of $[S^2]$ in Eq. (ii)

$$
1.3 \times 10^{-13} = \frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}
$$

or [HS⁻] =
$$
\frac{0.3 \times 1.44 \times 10^{-20}}{1.3 \times 10^{-13}} = 3.3 \times 10^{-8} \,\mathrm{M}
$$

(9) 9. Using Handerson equation for basic buffer

$$
pOH = pK_b + \log \frac{[salt]}{[base]}
$$

: For NH₄OH,
$$
(pK_a)_{NH_4^+}
$$
 + $(pK_b)_{NH_4OH} = 14$

$$
[S^{2-}] = \frac{K_{a_1} \times K_{a_2}[H_2S]}{[H^+]^2} = \frac{1.0 \times 10^{-7} \times 1.3 \times 10^{-13} \times (0.1)}{(0.3)^2}
$$
 initial, M (0.050
\nPutting the value of $[S^2]$ in Eq. (ii)
\nPutting the value of $[S^2]$ in Eq. (ii)
\n
$$
1.3 \times 10^{-13} = \frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}
$$
\nor [HS⁻] = $\frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}$
\nor [HS⁻] = $\frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}$
\n
$$
1.3 \times 10^{-13} = \frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}
$$
\n
$$
1.3 \times 10^{-13} = 3.3 \times 10^{-8} \text{ M}
$$
\nIf $y \ll 0.050$ then 5.6 ' 10'⁻¹⁰ $\frac{y^2}{0.050}$
\n9. Using Handerson equation for basic buffer
\n
$$
1.5 \times 10^{-13} = 3.3 \times 10^{-8} \text{ M}
$$
\nIf $y \ll 0.050$ then 5.6 ' 10'⁻¹⁰ $\frac{y^2}{0.50}$
\n
$$
y = 5.3 \times 10^{-6}
$$
\nThe assumption is 5.6 ' 10'⁻¹⁰ $\approx 2.8 \times 10^{-6}$
\nThe assumption is 5.6 ' 10'⁻¹⁰ $\approx 2.8 \times 10^{-6}$
\nThe assumption is 5.90H = 4.74 + log $\frac{2.5}{5}$ = 4.44
\n
$$
\therefore 1.1 = 14 - 4.44 = 9.56
$$
\n
$$
1.1 = 14 - 4.44 = 9.56
$$
\n
$$
1.1 = 14 - 4.44 = 9.56
$$
\n
$$
1.1 = 14 -
$$

$$
\therefore \text{ pH} = 14 - 4.44 = 9.56
$$

(10) 10. Let h be the degree of hydrolysis of the salt (anilinium acetate).

..
$$
pOII = 4.74 + log \frac{1}{5} = 4.44
$$

\n.. $pH = 14 - 4.44 = 9.56$
\n**(10) 10.** Let h be the degree of hydrolysis of the salt (anilinium acetate).
\nC 6 H 5 N H 3⁺ (a q) + C H 3 C O 2⁻ (a q) +
\nH₂O \rightleftharpoons C₆H₅NH₃OH+CH₃CO₂H

$$
0.1 (1-h) \t\t 0.1 (1-h) \t\t 0.1 h \t\t 0.1 h \t\t (2)
$$

= t 2.5 - 5 5 0
\n∴ pOH = 4.74 + log
$$
\frac{2.5}{5}
$$
 = 4.44
\n∴ pH = 14-4.44 = 9.56
\n∴ R = 14-4.44 = 9.56
\n= 1.5 N
\n

$$
\sqrt{\frac{1.008 \cdot 10^{-14}}{1.75 \cdot 10^{-5} \cdot 3.83 \cdot 10^{-10}}} = \frac{h}{(1-h)}
$$

h=0.55

(11) 5. First, write the chemical equation. $NH_4^+(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$ Calculate K_a from K_b for ammonia.

$$
K_a = \frac{K_w}{K_b} = \frac{1.0 \cdot 10^{-14}}{1.8 \cdot 10^{-5}} = 5.6 \times 10^{-10}
$$

Now, write the iCe [initial concentration (i), change due to the reaction (C) and the equilibrium concentration (e)] table.

Change, M –y +y +y
\nequilibrium, M 0.050-y y y
\nSubstitute into the equilibrium constant expression and
\nsolve for y.
\n
$$
K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]} \t ; 5.6'10^{-10} = \frac{(y)(y)}{0.050-y}
$$

\nIf y << 0.050 then 5.6'10⁻¹⁰ » $\frac{y^2}{0.050-y}$
\nIf y << 0.050 then 5.6'10⁻¹⁰ » $\frac{y^2}{0.050}$
\n $y^2 = 0.050 \times 5.6 \times 10^{-10} = 2.8 \times 10^{-11}$
\n $y = 5.3 \times 10^{-6}$
\nCheck the assumption. Is 5.3 × 10⁻⁶ << 0.050? Yes, it is.
\nThe assumption is good and we can accept the value.
\n $[H_3O^+] = 5.3 \times 10^{-6}$; pH = -log [5.6 × 10⁻⁶] = 5.28
\n**EXECUTE-4**
\n**SECTION-A [CHEMICAL EOUILLBRIUM]**
\n(1) (C). We know that, K_p = K_c(RT)^{Δn} [.: Δn = n_p - n_R]
\nFrom the reaction, $CO + \frac{1}{2}O_2 \longrightarrow CO_2$
\n $n_R = 1 + \frac{1}{2} = 1\frac{1}{2}$; $n_p = 1$; Δn = 1 - 1 $\frac{1}{2} = -\frac{1}{2}$

$$
y^2 = 0.050 \times 5.6 \times 10^{-10} = 2.8 \times 10^{-11}
$$

y = 5.3 × 10⁻⁶

Check the assumption. Is $5.3 \times 10^{-6} \ll 0.050$? Yes, it is. The assumption is good and we can accept the value.

 $[H_3O^+] = 5.3 \times 10^{-6}$; pH = $-\log[5.6 \times 10^{-6}] = 5.28$

EXERCISE-4 SECTION - A [CHEMICAL EQUILIBRIUM]

(1) (C). We know that,
$$
K_p = K_c (RT)^{\Delta n}
$$
 [: $\Delta n = n_p - n_R$]

ng the value of [S²]-j in Eq. (ii)
\n
$$
3 \times 10^{-13} = \frac{0.3 \times 1.44 \times 10^{-20}}{[HS^-]}
$$

\n $1.3 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.3 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-20} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-20} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-8} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-13} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-13} M$
\n $1.4 \times 10^{-13} = 3.3 \times 10^{-13} M$
\n $1.4 \times 10^{-13} M$
\n $$

(2) (C). Volume of reactant = Volume of product so, then equilibria is not affected by change in volume of the flask.

$$
(3) \quad (A). N_2O_4(g) \rightleftharpoons 2NO_2(g)
$$

$$
K_C = \frac{[NO_2]^2}{[N_2O_4]}
$$

Given value [NO²] = 1.2 × 10–2 mol/lit [N2O⁴] = 4.8 × 10–2mol/lit

So,
$$
K_C = \frac{(1.2 \times 10^{-2})^2}{(4.8 \times 10^{-2})} = \frac{1.44 \times 10^{-4}}{4.8 \times 10^{-2}} = 0.3 \times 10^{-2}
$$

= 3 × 10⁻³ mol/lit

(4) **(A).** The reaction
\n
$$
2SO_2(g) + O_2(g) \rightleftharpoons 2 SO_3(g); \Delta H^o = -198 \text{ kJ}
$$
\nSince the reaction is exothermic so, high temp. is
\nunfavourable, similarly in forward direction, the
\nvolume is low so pressure is high.

(D).
$$
P_4(S) + 5O_2(g) \rightleftharpoons P_4O_{10}(s)
$$

 (5)

 (13)

C 5 2 ¹ ^K [O] ;[P⁴] = [P4O10] = 1, due to solid state **(6) (A).** CO + Cl² COCl² n^R = 2 n^P = 1 (RT)–1 K RT 2 2 [N][O] = 4 × 10–4

$$
\overrightarrow{K}_{P} = K_{C} (RT)^{\Delta h}, K_{P} = K_{C} (RT)^{-1}
$$

$$
\frac{K_{P}}{K_{C}} = \frac{1}{RT}
$$

(7) **(B).** N₂(g) + O₂(g)
$$
\rightleftharpoons
$$
 2NO(g)
\n
$$
K_C = \frac{[NO]^2}{[N_2][O_2]} = 4 \times 10^{-4}
$$

For the reaction, $NO(g) \rightleftharpoons \frac{1}{2}N_2 + \frac{1}{2}O_2$ (14) (B). From $\frac{1}{2}N_2 + \frac{1}{2}O_2$ (14) (B). From

EXAMPLE 2.12.12.12.13	5.23.12.12.13	6.23.13.13
\n $K_C = \frac{1}{[O_2]^5}$ \n $E_C = \frac{[SO_2][O_2]^{1/2}}{P_2}$ \n	\n $K_C = \frac{[SO_2][O_2]^{1/2}}{[SO_3]} = 4.9 \times 10^{-2}$ \n	
\n $K_C = \frac{[SO_2][O_2]^{1/2}}{K_C}$ \n	\n $K_C = \frac{[SO_2][O_2]^{1/2}}{[SO_3]} = 4.9 \times 10^{-2}$ \n	
\n $K_C = \frac{[SO_2]^2}{[N_2][O_2]} = 4 \times 10^{-4}$ \n	\n $K_C = \frac{[SO_3]^2}{[SO_2]^2}$ \n	
\n $K_C = \frac{[NO]^2}{[N_2][O_2]} = 4 \times 10^{-4}$ \n	\n $K_C = \frac{[SO_2]^2}{[SO_2]^2}$ \n	
\n $K_C = \frac{[NO]^2}{[N_2][O_2]} = 4 \times 10^{-4}$ \n	\n $K_C = \frac{1}{[K_C]^2} = \frac{1}{4.9 \times 10^{-2} \times 4.9 \times 10^{-2}}$ \n	
\n $K_C = \frac{[N_2]^{1/2} \cdot [O_2]^{1/2}}{[NO]}$ \n	\n $K_L = \frac{[CO_2][H_2]}{[CO][H_2O]}, K_2 = \frac{[CO][H_2]^3}{[CH_4][H_2O]},$ \n	
\n $K_C = \frac{1}{[K_C]^{1/2}} = \frac{1}{\sqrt{K_C}} = \frac{1}{\$		

- **(8) (B).** $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$ $K_C = 1.8 \times 10^{-6}$, T = 184 °C = 184 + 273 = 475 K $R = 0.0831$ kJ/mol k, $K_p = K_C (RT)^{\Delta n}$ $K_p = K_C(RT)$ $\frac{P}{P} - RT$ a v v v \mathcal{C} and \mathcal{C} and \mathcal{C} $\frac{K_P}{K_C}$ = RT . So, $K_P > K_C$ (15)
- **(9) (C).** Cl_{2(g)} + 3F_{2(g)} \implies 2ClF_{3(g)}; Δ_r H = –329 kJ From the above reaction, the quantity of ClF_3 increase with adding F_2 .
- **(10) (B).** With decrease in temperature (increase in 1/T), the value of K_{eq} is increasing. Thus the reaction must be exothermic.
- **(11) (C).** $NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$ x x Already present $NH_3 = NH_3(g)$ 0.50 atm So, at equilibrium total pressure = $0.5 + 2x = 0.84$

 $x = 0.17$. So, at equilibrium,

$$
P_{\text{NH}_3} = (0.50 + 0.17) = 0.67
$$
; $P_{\text{H}_2\text{S}} = 0.17$

Hence, $K_p = P_{NH_3} \times P_{H_2S} = 0.67 \times 0.17 = 0.1139$

(12) **(D)**
$$
PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)
$$

Initial a At equilibrium $(a - ax)$ ax ax Total moles at equilibrium = $a + ax = a(1 + x)$

Mole fraction of PCI₃ =
$$
\frac{ax}{a(1+x)} = \frac{x}{(1+x)}
$$

P_{PCI₃} = $\left(\frac{x}{1+x}\right)$ P

$$
(D). SO3(g) \rightleftharpoons SO2(g) + \frac{1}{2}O2(g)
$$

S
\nSTUDYMATERIAL: CHEMISTRY
\n(D). SO₃(g)
$$
\Leftrightarrow
$$
 SO₂(g) + $\frac{1}{2}$ O₂(g)
\nK_C = $\frac{[SO_2][O_2]^{1/2}}{[SO_3]} = 4.9 \times 10^{-2}$
\nFor the reaction, 2SO₂ (g) + O₂ (g) \Leftrightarrow 2SO₃ (g)
\nK_C' = $\frac{[SO_3]^2}{[SO_2]^2 [O_2]}$

For the reaction, $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$

$$
K'_{C} = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}
$$

S **STUDY MATERIAL: CHEMISTRY**
\n**(D).** SO₃(g)
$$
\rightleftharpoons
$$
 SO₂(g) + $\frac{1}{2}$ O₂(g)
\n
$$
K_C = \frac{[SO_2][O_2]^{1/2}}{[SO_3]} = 4.9 \times 10^{-2}
$$
\nFor the reaction, 2SO₂ (g) + O₂ (g) \rightleftharpoons 2SO₃ (g)
\n
$$
K_C' = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}
$$
\n
$$
K_C' = \frac{1}{[K_C]^2} = \frac{1}{4.9 \times 10^{-2} \times 4.9 \times 10^{-2}}
$$
\n
$$
= 4.16 \times 10^2 = 416
$$
\n**(B).** From the equation,
\n
$$
K_1 = \frac{[CO_2][H_2]}{[CO][H_2O]}, K_2 = \frac{[CO][H_2]^3}{[CH_4][H_2O]},
$$
\n
$$
K_3 = \frac{[CO_2][H_2]^4}{[CH_4][H_2O]^2}
$$
\nOn multiplying K_1 and K_2
\n
$$
K_1 \cdot K_2 = \frac{[CO_2][H_2]^2}{[CO_2][H_2][CO][H_2]^3} = \frac{[CO_2][H_2]^4}{[CO_2][H_2]^4} = K_3
$$

 $1 \t= 4.10 \times 10^{-7}$ $\frac{1}{2}$ O₂ (14) (B). From the equation,

$$
K_1 = \frac{[CO_2][H_2]}{[CO][H_2O]}, K_2 = \frac{[CO][H_2]^3}{[CH_4][H_2O]},
$$

$$
\frac{1}{1 \times 10^{-2}} = 50
$$
 K₃ =
$$
\frac{[CO_2][H_2]^4}{[CH_4][H_2O]^2}
$$

On multiplying K_1 and K_2

K_C =
$$
\frac{1}{|O_2|^2}
$$
 :[P₄] = [P₄O₁₀] = 1, due to solid state (13) (D). SO₃(g) ⇒ SO₂(g) + $\frac{1}{2}O_2(g)$
\n $CO + C12 ⇒ COC12$
\n $CO + C12 ⇒ COC12$
\n $CO + C12 ⇒ COC12$
\n $Ke = \frac{[SO_2][O_2]^{1/2}}{[SO_2]^2}$
\n $Ke = \frac{1}{[SO_2]^2}$
\n $Ke = \frac{[SO_2][H_2]}{[SO_2]^2}$
\n $Ke = \frac{[SO_2][H_2]^3}{[SO_2]^2}$
\n $Ke = \frac{[CO_2][H_2]^4}{[SO_2]^2}$
\n $Ke = \frac{[CO_2][H_2]^3}{[SO_2]^2}$
\n $Ke = \frac{[CO_2][H_2]^4}{[SO_2]^2}$
\n $Ke = \frac{[CO_2][H_2]^4}{[SO_2]^2}$
\n $Ke = K2(g)(g) + O_2(g)$
\n $Ch = 1.8 \times 16.9, K₂ = K_C$

EQUILIBRILM)	Q.B. SOLUTIONS	SEOLUTION-B [IONIC EQUILIBRILM]		
(18)	(D) For reaction: $SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g)$	(1)	$Mod(9)$	$Mod(9)$
$\Delta n_g = -\frac{1}{2} = x$	$K_{sp} = [Mg^2]^2$ [OH] ² = $(x)(2x)^2 = 4x^3$			
(19)	(A) $2A \rightleftharpoons B + C$; $\Delta G^\circ = 2494.2$ J	(2)	(C) . Find A_2 by ϕ = $4 \times (1.0 \times 10^{-5})^3$	
$\Rightarrow 2494.2 - 2.303 \text{ N} = 10 \text{ N}$	4×10^{-15}			
$\Rightarrow -0.434 - \log K_C$	(3)	(D). $MX_4 \rightleftharpoons M^4 + 4X$		
$\Rightarrow -0.434 - \log K_C$	(3)	(D). $MX_4 \rightleftharpoons M^4 + 4X$		
$\Rightarrow -0.434 - \log K_C$	(3)	(D). $MX_4 \rightleftharpoons M^4 + 4X$		
$\Rightarrow -0.434 - \log K_C$	(3)	(4)	(A). For the salt,	
$\text{Now } [A] = 1/2, [B] = 2$ and $[C] = 1/2$	$K_{sp} = [M^4][X^4]$			
$\text{Now } [A] = 1/2, [C] = 2$ and $[C] =$				

 $\rm Q_C$ > K_C, reaction will shift in backward direction. **(20) (B).** Initially at equilibrium

Now,
$$
Q_C = \frac{[C][B]}{[A]^2} = \frac{(1/2)(2)}{(1/2)^2} = 4
$$

\n $Q_C > K_C$, reaction will shift in backward direction.
\n**B)**. Initially α equation with $A + B \rightleftharpoons C + D$; $K_{eq} = 100$
\n $1 - 1 - 1 - 1 = 1$
\n $(1-x)(1-x)(1+x)(1+x)(1+x) = (1+x)^2$
\n $K_{eq} = \frac{[C][D]}{[A][B]} = \frac{(1+x)(1+x)(1+x)}{(1-x)(1-x)} = \frac{(1+x)^2}{(1-x)^2}$
\n $10 = \frac{1+x}{1-x}$. On solving, $x = 9/11$, $[D] = 1.818$
\n $10 = \frac{1+x}{1-x}$. On solving, $x = 9/11$, $[D] = 1.818$
\n $10 = \frac{-\Delta H}{RT} + \frac{\Delta S}{R}$; Slope is $\frac{-\Delta H}{R}$
\n $10 = \frac{-\Delta H}{RT} + \frac{\Delta S}{R}$; Slope is $\frac{-\Delta H}{R}$
\n $10 = \frac{-\Delta H}{RT} + \frac{\Delta S}{R}$; Slope is $\frac{-\Delta H}{R}$
\n $10 = \frac{-\Delta H}{RT} + \frac{\Delta S}{R}$; Slope is positive.
\n $10 = \frac{2S}{R}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_2}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_2}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 = \frac{S_2 - S_1}{S_2 + S_1}$
\n $10 =$

(21) (C). $\Delta G = \Delta H - T \Delta S$; $-RT \ln k = \Delta H - T \Delta S$

ln k =
$$
\frac{-\Delta H}{RT} + \frac{\Delta S}{R}
$$
; Slope is $\frac{-\Delta H}{R}$ (6) (C). Given value, pKa = 4.5
\nSince ΔH is - ve
\n \therefore Slope is positive.

(22) **(B).** A₂(g) + Br₂(g)
$$
\xrightarrow{K_1}
$$
 2AB(g).....(1) \Rightarrow eq. (1) \times 3

 $6AB(g) \xrightarrow{K_2} 3A_2(g) + 3B_2(g)$ \Rightarrow $|\overline{1}|\overline{2}|$ = k₂ \Rightarrow l 3 3 1^J 1 $\binom{3}{2}$ 1 1 1 $\binom{1}{2}$ $\binom{3}{2}$ -3

(23) (C). $S(s) + O_2(g) \rightleftharpoons SO_2(g)$; $K_1 = 10^{52}$...(1) $2S(s) + 3O_2(g) \rightleftharpoons 2SO_3(g); K_2 = 10^{129}$...(2) $2SO_2(g) + O_2(g) \rightleftharpoons SO_3(g)$; K₃ = x Multiplying equation (1) by 2; $2SO(s) + 2O_2(g) \rightleftharpoons 2SO_2(g)$; K₁' = 10¹⁰⁴ ...(3) Substracting (3) from (2) ; we get $2SO_2(g)+O_2(g) \rightleftharpoons 2SO_3(g);$ $K_{eq} = 10^{(129 - 104)} = 10^{25}$ **(3)** $\frac{(\sqrt{16})^3}{(\sqrt{16})^2}$ $\Rightarrow (\frac{1}{24})^3 = k_2 \Rightarrow k_2 = (k_1)^{-3}$
 (3) $(\sqrt{26})(k_1 + 10^2) = 2(2)(k_1 + 10^2)$
 (3) $(\sqrt{26})(k_1 + 10^2) = 2(2)(k_1 + 10^2)$
 (4) $\frac{2500}{(k_1 + 10^2)(k_2 + 200)(k_1)} = 2(2)(k_1 + 10^2)$
 (4) $\frac{2500}{$

(24) (A). In option (B)- Δn_g is –ve therfore increase in pressure will bring reaction in forward direction. In option (C) - as the reaction is exothermic therefore increase in temperature will decrease the equilibrium constant. In option (D)- Equillibrium constant changes only (10) with temperature.

> Hence, option (B), (C) and (D) are correct therefore option (A) is incorrect choice.

(25) **(D)**. If
$$
\Delta n_g \neq 0
$$

\n $K_p \neq K_c$
\n V_p [B] 11

(26) (A).
$$
K = \frac{|B|}{[A]} = \frac{11}{6} = 2
$$

SECTION - B [IONIC EQUILIBRIUM]

(1)
$$
\log(OH)_2 \rightleftharpoons Mg^{+2} + 2OH^-
$$

\n $\times 2x$
\n $\frac{2x}{12x^2 + 20H^-} = 2x^2 + 20H^-$

(2)
$$
K_{\text{sp}} = [\text{Mg}^{+2}] [\text{OH}]^{-2} = (x) (2x)^2 = 4x^3
$$

(2) **(C).** For AB₂ type
Solubility product = $4x^3 = 4 \times (1.0 \times 10^{-5})^3$
= 4×10^{-15}

UITIONS

\n**SECTION - B [IONIC EQUILLIBRIUM]**

\n(1)
$$
\begin{array}{rcl}\n& Mg(OH) & \xrightarrow{\sim} 2x \\
& K_{sp} = [Mg^{+2}] [OH]^{-2} = (x)(2x)^{2} = 4x^{3} \\
&(2) & (C). \text{ For AB}_{2} \text{ type} \\
& Solubility product & = 4x^{3} = 4 \times (1.0 \times 10^{-5})^{3} \\
& = 4 \times 10^{-15}\n\end{array}
$$
\n(3) (D). MX₄ $\rightleftharpoons M^{4+} + 4X^{-}$

\n
$$
\begin{array}{rcl}\n& K_{sp} & = [M^{4+}] [X^{-}]^{4} \\
& & K_{sp} & = (3) (4s)^{4} = 256s^{5} \\
& & s = \left(\frac{K_{sp}}{256}\right)^{1/5} \\
& & K_{sp} & = M^{+2} + 2X^{-} \\
& & K_{sp} & = 4s^{3} \\
& & K_{sp} & = 4s^{3} \\
& & K_{sp} & = 4s^{3}\n\end{array}
$$

$$
s = \left(\frac{K_{sp}}{256}\right)^{1/3}
$$

(1)
$$
\dots
$$
 Mg(OH)₂ $\underset{x}{\rightleftharpoons}$ Mg⁻² + 2OH
\n $\underset{x}{x}$ 2x
\n $K_{sp} = [Mg^{+2}][OH]^{-2} = (x)(2x)^{2} = 4x^{3}$
\n(2) (C). For AB₂ type
\nSolubility product = $4x^{3} = 4 \times (1.0 \times 10^{-5})^{3}$
\n $= 4 \times 10^{-15}$
\n(3) (D). MX₄ \rightleftharpoons M⁴⁺ + 4X⁻
\n $= (s)$ $(4s)^{4} = 256s^{5}$
\n $s = \left(\frac{K_{sp}}{256}\right)^{1/5}$
\n(4) (A). For the salt,
\n $MX_{2} \rightleftharpoons M^{+2} + 2X^{-}$
\n $= (s)$ (2s)
\n $K_{sp} = 4s^{3}$
\n $4 \times 10^{-12} = 4s^{3}$
\n $s = 1 \times 10^{-4} M$
\n(M⁺²) = s = 1 × 10⁻⁴ M
\n(M⁺²) = s = 1 × 10⁻⁴ M
\n(5) (C). H₂A $\frac{k_{a}}{a} \rightarrow H^{+} + HA^{-}$
\n $HA^{-} \xrightarrow{k_{b}} H^{+} + A^{2-}$
\n(6) (C). Given value, pKa = 4.5
\n $pH = pKa + log \frac{[A^{-}]}{[HA]} = 4.5 + log \frac{[A^{-}]}{[A]} = 4.5$

$$
(5) \qquad (C). \ H_2A \xrightarrow{k_a} H^+ + HA
$$

 $\frac{-\Delta H}{\Delta}$ **(6) (C).** Given value, pKa = 4.5

An_g -1/3
\nAs
$$
y = B + C
$$
; $\Delta G^2 = 2494.21$
\nAs $y = 230$ RT log R
\nSo $x = 230$ RT log R
\n $x = 230$

8) (B). pH =
$$
\frac{1}{2}
$$
(pK_a - pK_b) + 7 = $\frac{1}{2}$ (4.80-4.78) + 7
= $\frac{1}{2}$ (0.02) + 7 = 0.01 + 7 = 7.01

(9) (A). Concentration of Ba2+ needed to precipitate BaCO³

$$
= \frac{K_{sp} \text{ of } BaCO_3}{[CO_3^{2-}]} = \frac{5.1 \times 10^{-9}}{1.0 \times 10^{-4}} = 5.1 \times 10^{-5} \text{ M}
$$

(10) (C). A \rightarrow H₂CO₃ \Longrightarrow H⁺ + HCO₃⁻; K₁ = 4.2 × 10⁻⁷
B \rightarrow HCO₃⁻ \Longrightarrow H⁺ + CO₃² · K₃ = 4.8 × 10⁻¹¹

B
$$
\rightarrow
$$
 HCO₃⁻ \rightleftharpoons H⁺ + CO₃²⁻ ; K₂ = 4.8 × 10⁻¹¹
As K₂ << K₁
All major [H⁺]_{total} = [H⁺]_A
and from I equilibrium, [H⁺]_A \approx [HCO₃⁻] \approx [H⁺]_{total}
[CO₃²⁻] is negligible compared to [HCO₃⁻] or [H⁺]_{total}

EXAMPLE 10.1	10.2	10.3	10.4	10.4
(11)	(B) A $q^2 + Br = \frac{1}{\sqrt{6}} \times 10^2$	(20)	(D) HCl → H ⁺ + C ⁺	
(E) The probability product K _{pp} = [Br ⁻¹] \n $[Br^{-1} = \frac{K_{sp}}{Ag^2}] = \frac{5 \times 10^{-3}}{0.05} = 10^{-11}$ \n	10.5	10.4		
(E) The probability product K _{pp} = [Brr/3, Brr/3]				
(E) The probability matrix starts when 10 ⁻¹¹ moles of KBr	$K = \frac{[H^+]^2 S^2}{[H^2 S^2]}$; $H^+ = 12 \times 10^{-20}$			
(E) The probability matrix starts when 10 ⁻¹¹ moles of KBr	$K = \frac{[H^+]^2 S^2}{[H^2 S^2]}$; $H^+ = 12 \times 10^{-20}$			
(E) The probability matrix of the added = 10 ⁻¹¹ × 10 ⁻⁹ = 10 ⁻⁴				
(E) The probability matrix for the added = 10 ⁻¹¹ × 10 ⁻⁹ = 10 ⁻⁴				
(E) The probability matrix for the added = 10 ⁻¹¹ × 10 ⁻⁹ = 10 ⁻⁴				
(E) The probability matrix for the provided = 10 ⁻¹¹ × 10 ⁻⁹ = 10 ⁻⁴				
(E) The probability matrix for the provided = 10 ⁻¹¹ × 10 ⁻¹⁰ = 10 ⁻⁴				
(E) The probability matrix for the provided =				

$$
pH = 7 + \frac{1}{2}pK_a - \frac{1}{2}pK_b = 7 + \frac{3.2}{2} - \frac{3.4}{2} = 6.9
$$

- (17) **(D).** CH₃COOK is a salt of a weak acid and a strong base
 \therefore Most basic
- **(18) (A).** Methyl orange is used for titration of strong acid and weak base.

(19) **(A).** BaSO₄ (s)
$$
\rightleftharpoons
$$
 Ba²⁺ (aq) + SO₄²⁻ (aq); K_{sp} = 10⁻¹⁰
\nNa₂SO₄ \rightarrow 2Na⁺ + SO₄²⁻
\nCone. of SO₄²⁻ in final solution = $\frac{50 \times 1}{500} = 0.1$ M
\nFor final solution, [Ba²⁺] [SO₄²⁻] = 10⁻¹⁰
\n[Ba²⁺] = 10⁻⁹ M; M₁V₁ = M_fV_f
\nC × 450 = 10⁻⁹ × 500; C = 1.1 × 10⁻⁹M

1.2 × 10⁻²⁰ =
$$
\frac{(0.2)^2}{1} = 10^{-4}
$$

\n $\frac{K_3}{1} = 10^{-4}$
\n $\frac{K_4}{1} = 10^{-2} = 10^{-4}$
\n $\frac{K_5}{1} = 10^{-1}$
\n $\frac{K_7}{1} = 10^{-2} = 10^{-4}$
\n $\frac{K_8}{1} = 10^{-1}$
\n $\frac{K_9}{1} = 10^{-11}$
\n $\frac{K_{10}}{1} = 10^{-11} = 10^{-11}$
\n $\frac{K_9}{1} = 10^{-10}$
\n $\frac{K_9}{1} = 10^{-10}$
\n $\frac{K_9}{1} = 10^{-10}$
\n $\frac{K_9}{1} = 10^{-10}$
\n $\frac{K_9}{1} = 10^{-11}$
\n $\frac{K_9$

$$
1.2 \times 10^{-20} = \frac{(0.2)^2 [S^{2-}]}{0.1}; \quad [S^{2-}] = 3 \times 10^{-20} M
$$

$$
(C). 20 \text{ m}10.1 \text{ M H}_2\text{SO}_4 \Rightarrow {}^{11}\text{H}^+ = {}^{4}
$$

NH₄OH + H⁺
$$
\rightleftharpoons
$$
 NH₄[®] + H₂O
\n6 4 0 0
\n2 0 4 4

Solution is basic buffer $pOH = pK_b + log \frac{NH_4}{NH_4OH}$ NH_4^+ \pm and \pm and \pm

$$
K = \frac{[H^+]^2 [S^{2-}]}{[H_2S]}; [H^+] = 0.2 M, [H_2S] = 0.1
$$

\n1.2×10⁻²⁰ = $\frac{(0.2)^2 [S^2]}{0.1}$; $[S^2 = 3 \times 10^{-20} M$
\n(21) **(C)**.20 ml 0.1 M H₂SO₄ \Rightarrow $\eta_{H^+} = 4$
\n30 ml 0.2 M NH₄OH \Rightarrow $\eta_{NH_4} \theta + H_2 O$
\n6 M₄OH + H⁺ \Rightarrow NH₄⁴⁰ + H₂O
\n7 0 4 4
\nSolution is basic buffer $pOH = pK_b + \log \frac{NH_4^+}{NH_4OH}$
\n= 4.7 + log 2 = 4.7 + 0.3 = 5; pH = 14 - 5 = 9
\n(22) **(D)**. Zr₃(PO₄)₄(s) \Rightarrow 3Zr⁴⁺(aq.) + 4 PO₄³⁻(aq.)
\nK_{sp} = [Zr⁴⁺]³ [PO₄³⁻]⁴⁻ = (3S)³. (4S)⁴ = 6912 S⁷
\n $\therefore S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$
\n(23) **(B)**. For the salt of strong acid and weak base
\nH⁺ = $\sqrt{\frac{K_w \times C}{K_b}}$; [H⁺] = $\sqrt{\frac{10^{-14} \times 2 \times 10^{-2}}{10^{-5}}}$
\n $-\log [H^+] = 6 - \frac{1}{2} \log 20$ $\therefore pH = 5.35$
\n(24) 10.60
\nM<sub>H₂S0₄ = $\frac{9.8}{98 \times 100} = 10^{-3}$
\</sub>

(23) (B). For the salt of strong acid and weak base

$$
H^{+} = \sqrt{\frac{K_{w} \times C}{K_{b}}}; \quad [H^{+}] = \sqrt{\frac{10^{-14} \times 2 \times 10^{-2}}{10^{-5}}}
$$

$$
-\log[H^{+}] = 6 - \frac{1}{2} \log 20 \qquad \therefore pH = 5.35
$$

(24) 10.60

$$
A_{\text{H}_2\text{SO}_4} = \frac{9.8}{98 \times 100} = 10^{-3}
$$

$$
M_{\text{NaOH}} = \frac{4}{40 \times 100} = 10^{-3} = \frac{40 \times 10^{-3} - 10 \times 10^{-3} \times 2}{50} = \frac{20}{50} \times 10^{-3}
$$

[OH⁻] = $\frac{2}{5} \times 10^{-3}$; pOH = 3.397; pH = 10.603

Solution is basic buffer
$$
pOH = pK_b + log \frac{NH_4^+}{NH_4OH}
$$

\n $= 4.7 + log 2 = 4.7 + 0.3 = 5$; $pH = 14 - 5 = 9$
\n(22) **(D).** $Zr_3(PO_4)_{4}(s) \rightleftharpoons 3Zr^{4+}(aq,) + 4 PO_4^{3-}(aq,)$
\n $35 M$
\n $K_{sp} = [Zr^{4+1}]^3 [PO_4^{3-1}]^4 = (3S)^3.(4S)^4 = 6912 S^7$
\n $\therefore S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$
\n(23) **(B).** For the salt of strong acid and weak base
\n $\frac{6}{L} = 10^{-5}$
\n $H^+ = \sqrt{\frac{K_w \times C}{K_b}}$; $[H^+] = \sqrt{\frac{10^{-14} \times 2 \times 10^{-2}}{10^{-5}}}$
\n $-log [H^+] = 6 - \frac{1}{2} log 20$ $\therefore pH = 5.35$
\n(24) **10.60**
\nlitre.
\n $M_{H_2SO4} = \frac{9.8}{98 \times 100} = 10^{-3}$
\nand weak base is
\n $M_{H_2SO4} = \frac{9.8}{98 \times 100} = 10^{-3}$
\nand $\frac{4}{50} \times 10^{-3} = 50$
\nand a strong base
\n $10H^- = \frac{2}{5} \times 10^{-3}$; $pOH = 3.397$; $pH = 10.603$
\n $\frac{3.4}{500} = 6.9$
\nand a strong base
\n $10H^- = \frac{2}{5} \times 10^{-3}$; $pOH = 3.397$; $pH = 10.603$
\n $10H = pK_0 + log (160) = 10$
\n $10H = pK_0 + log (120) = 10$
\n $10H = pK_0 + log (120) = 10$

EQUILBRIUM	QB-SOLUTIONS	CDIMADVANGEDEARINING	
$[CI^-] = \frac{100 \times 0.4}{400} = 10^{-1} M$	$\left(\frac{1}{K_1}\right)^2 = \left\{\frac{[NO][O_2]^{1/2}}{[NO_2]}\right\}^2 = \frac{[NO]^2 [O_2]}{[NO_2]^2} = K_2$		
$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$	$Q = [Pb^{2+}] \times [CI^-]^2 = 1.005 \times 10^{-3} \times k_{sp}$	$\frac{1}{K_1^2} = K_2$	
(27)	$(A). Cr(OH)3(s) \rightleftharpoons Cr^{3+}(aq.) + 3OH^-(aq.)$	$\frac{1}{K_1^2} = K_2$	
$K_{sp} = 27(s)^4 = 6 \times 10^{-31}$	$(3s)$	(5)	$(A). CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O(\ell)$
$Pd = -170.8 kJ mol^{-1}$			
$[O_2]$	V	$Pd = -170.8 kJ mol^{-1}$	

EXERCISE-5

(1) (A). The solution formed from isomolar solutions of sodium oxide, sodium sulphide, sodium selenide and (6) sodium telluride are H_2O , H_2S , H_2Se and H_2Te respectively. As the acidic strength increases from H_2O to H_2Te thus pH decreases and hence the correct
of pHs is : $pH_1 > pH_2 > pH_3 > pH_4$ (7) of pHs is : $pH_1 > pH_2 > pH_3 > pH_4$

(2) (D). Given : K^b = 1.0 × 10–12 . [BOH] = 0.01 M, [OH] = ? BOH B+ + OH– t = 0 c 0 0 teq c (1 – x) cx cx 2 2 2 ^b c x cx ^K c (1 x) (1 x) ; 1.0 × 10–12⁼ ² 0.01 x On calculation, we get, x = 1.0 × 10–5

Now [OH⁻] = cx = 0.01×10^{-5} = 1.0×10^{-7} mol L⁻¹

(3) (A). In qualitative analysis of cations of second group H2S gas is passed in presence of HCl, therefore due to common ion effect, lower concentration of sulphide ions is obtained which is sufficient for the precipitation of second group cations in the form of their sulphides due to lower value of their solubility product (K_{sp}) . Here, fourth group cations are not precipitated because it require more sulphide ions for exceeding their ionic product to their solubility products which is not obtained here due to common ion effect. when $x_k = \frac{e^x}{(k_1 - 1)} = 10^{-1}$
 $k_1 = \frac{e^x}{(k_1 - 2)} = \frac{e^x}{(k_1 - 2)} = \frac{0.01 \text{ s}^2}{0.01 \text{ s}}$
 $k_2 = \frac{1}{(k_1 - 2)(2)^{1/2}} = \frac{1}{2}$
 $k_3 = \frac{1}{(k_1 - 2)(2)^{1/2}} = \frac{1}{2}$
 $k_4 = \frac{(k_2 - 2)}{(k_1 - 2)(2)^{1/2}} = \frac{0.01 \text{ s}^2}{0.01 \$ (OH = x = 0 = x = 0 = 0 = 0 = x = 0 = 0 = x = (x = 0 = x = 0 = x = (x = 0 = x = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = c = x = $K_1 = \frac{e^3x^3}{(1-x)} - \frac{1}{(1-x)}$; $1.0 \times 10^{-12} = \frac{0.01 \times 3}{0.01(1-x)}$ (b) (A). The required equation of the oxidation of NH;

On calculation, we get $x = 1.0 \times 10^{-5} = 1.0 \times 10^{-7}$ mol 1⁻¹

Now [OIT] – ex – 0.0 i × 10⁻⁵ $-\frac{c^2x^2}{c(1-x)} = \frac{cx}{(1-x)}$; $1.0 \times 10^{-12} = \frac{0.01 \times 1}{0.01 (1-x)}$

(a) (A). The required equation for the oxidation of b

signal to give NO 10^{-5} oxygen to give NO is:
 $\frac{1}{c(10+1)^2 - c} = 0.01 \times 10^{-5}$

(a) $\frac{1}{c(10+1$ eir sulphides due to lower value of their solubility

odue (K_{Sp}). Here, fourth group cations are not

eccipitated because it require more sulphide ions
 $K_1^2 = \frac{[NH_3]^4}{[N_2]^2[H_2]^6}$; $K_2^2 = \frac{[N0]^4}{[N_2]^2[0_2]}$

odue Even for expecting the same of scaling the set of the set of $\sqrt{8}$ = $\frac{F(1 \text{ g})^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1^2]}$ = $\frac{F(0)^2}{[1$ n qualitarive analysis of cations of second group
 $x_1 \leq \frac{1}{2} \exp \frac{(\log x) \log x}{(x + \log x)^2}$

ocomma ion effect, because of HCl, therefore due

ocomma ion effect, because of HCl, therefore solution

or consiste observation of calculation, we get, $x = 1.0 \times 10^{-5}$

(c) (OIT) $= x = 0.00 \times 10^{-5}$ and μ and μ

(4) (C). For equation (i),
$$
NO(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2(g)
$$

$$
K_1 = \frac{[NO_2]}{[NO][O_2]^{1/2}}
$$
(1)

For equation (ii), $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$

$$
K_2 = \frac{[NO]^2 [O_2]}{[NO_2]^2}
$$
(2)

Now, on reversing eq. (1), we get

$$
\frac{1}{K_1} = \frac{1}{\frac{[NO_2]}{[NO_2]^{1/2}}} = \frac{[NO][O_2]^{1/2}}{[NO_2]}
$$

 2 2 1/2 2 1 2 ¹ [NO][O] K [NO] 2 2 2 2 2 [NO] [O] ^K [NO] 2 2 1¹ ^K K 4 2 [CO] ^K [CH][O] ; CH O 4 2 p p ⁼ ⁵ ^K^a 1 10 ² 10 1% C 0.100

(5) (A). CH₄(g) + 2O₂(g)
$$
\Longleftrightarrow
$$
 CO₂ (g) + 2H₂O (ℓ)
\n $\Delta_r H = -170.8 \text{ kJ mol}^{-1}$

$$
K_c = \frac{[CO_2]}{[CH_4][O_2]^2}; K_p = \frac{PCO_2}{PCH_4 PO_2^2}
$$

The reaction is an example of exothermic reaction.

- (5) (A). CH₄(g) + 2O₂(g) ⇒ CO₂ (g) + 2H₂O(*c*)

(A). CH₄(g) + 2O₂(g) ⇒ CO₂ (g) + 2H₂O(*c*)
 Δ_r H = -170.8 kJ mol⁻¹
 $K_c = \frac{[CO_2]}{[CH_4][O_2]^2}$; $K_p = \frac{PCO_2}{PCH_4 PO_2^2}$

solutions of The reaction is a **(A).** For a solution of 10^{-8} M HCl $[H^+] = 10^{-8}$ $[H^+]$ of water = 10^{-7} Total $[H^+] = 10^{-7} + 10^{-8} = 10 \times 10^{-8} + 10^{-8}$ $10^{-8}(10+1) = 11 \times 10^{-8}$ PCH₄ PO₂

xample of exothermic reaction.

7⁸ M HCl [H⁺] = 10⁻⁸

10⁻⁸

10⁻⁸

10⁻⁸

HNO₂ is a weak acid and NaNO₂

acid and strong base (NaOH).

= 0.100 mol for a weak electrolyte.

tion (α)

= $\sqrt{\frac{$ eaction is an example of exothermic reaction.

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

of water = 10^{-7}
 $[H^+] = 10^{-7} + 10^{-8} = 10 \times 10^{-8} + 10^{-8}$

(0+1) = 11×10^{-8}

the answer as HNO₂ is a weak acid and NaNO₂
 ample of exothermic reaction.

⁸ M HCl [H⁺] = 10⁻⁸
 0^{-8} = 10 × 10⁻⁸ + 10⁻⁸
 0^{-8}
 0^{-8 $K_e = \frac{[CO_2]}{[CH_4][O_2]^2}$; $K_p = \frac{PCO_2}{PCH_4 PO_2^2}$

The reaction is an example of exothermic reaction.

For a solution of 10⁻⁸ M HCl [H⁺] = 10⁻⁸

H⁺] of water = 10⁻⁷

Thatal [H⁺] = 10⁻⁷ + 10⁻⁸ = 10 × 10⁻ + 2O₂(g) \rightleftharpoons CO₂ (g) + 2H₂O (ℓ)

A_FH = -170.8 kJ mol⁻¹

[CO₂]

[H₄][O₂]²; K_p = $\frac{PCO_2}{PCH_4 PO_2^2}$

(H₄][O₂]²; K_p = $\frac{PCO_2}{PCH_4 PO_2^2}$

(water = 10⁻⁷

(HT₁=10⁻⁷

(HT₁=10⁻ $\frac{1.62}{14 \cdot 10^{2}}$; $K_p = \frac{1.62}{p_{CH_4} p_{O_2}^2}$
tion is an example of exothermic reaction.
ution of 10⁻⁸ M HCl [H⁺] = 10⁻⁸
vater = 10⁻⁷
 \rightarrow [1= 10⁻⁷ + 10⁻⁸ = 10 × 10⁻⁸ + 10⁻⁸
 \rightarrow [1= 10⁻⁷ + 10⁻⁸ a
	- **(7) (D).** (D) is the answer as $HNO₂$ is a weak acid and $NaNO₂$ is salt of that weak acid and strong base (NaOH).
	- **(8) (A).** $K_a = 1.00 \times 10^{-5}$, $C = 0.100$ mol for a weak electrolyte. Degree of dissociation (α)

$$
7
$$
\n
$$
10^{-8} = 10 \times 10^{-8} + 10^{-8}
$$
\n
$$
10^{-8}
$$
\n
$$
HNO_2
$$
 is a weak acid and NaNO_2\nacid and strong base (NaOH).\n
$$
= 0.100 \text{ mol for a weak electrolyte.}
$$
\n
$$
\sin(\alpha)
$$
\n
$$
= \sqrt{\frac{K_a}{C}} = \sqrt{\frac{1 \times 10^{-5}}{0.100}} = 10^{-2} = 1\%
$$
\n
$$
\sinh \text{ for the oxidation of NH}_3 \text{ by}
$$
\nis :
\n
$$
\frac{Pt(\text{gauze})}{800^{\circ}\text{C}} \rightarrow 2\text{NO} + 3\text{H}_2\text{O}
$$
\n
$$
\frac{2[H_2O]^3}{[P_2]^{5/2}}
$$
\n
$$
\text{We must do}
$$
\n
$$
= \frac{[NO]^4}{[N_2]^2[O_2]^2}
$$

1

(9) (A). The required equation for the oxidation of $NH₃$ by oxygen to give NO is :

$$
2NH_3 + \frac{5}{2}O_2 \xrightarrow{\text{Pt (gauze)}} 2NO + 3H_2O
$$

For this
$$
K = \frac{[NO]^2 [H_2 O]^3}{[NH_3]^2 [O_2]^{5/2}}
$$

For getting the K we must do

$$
K_{c} = \frac{[CO_{2}]}{[CH_{4}][O_{2}]^{2}}; K_{p} = \frac{PCO_{2}}{PCH_{4} P_{O_{2}}^{2}}
$$

The reaction is an example of exothermic reaction.
\nI. For a solution of 10⁻⁸ M HCl [H⁺] = 10⁻⁸
\n[H⁺] of water = 10⁻⁷
\n10⁻⁸ [(10 + 1) = 11 × 10⁻⁸
\n10⁻⁸ [(10 + 1) = 11 × 10⁻⁸
\n10⁻⁸ (10 + 1) = 11 × 10⁻⁸
\n10⁻⁵ (C = 0.100 mol for a weak and NaNO₂
\nis salt of that weak acid and strong base (NaOH).
\nL. K_a = 1.00 × 10⁻⁵, C = 0.100 mol for a weak electrolyte.
\nDegree of dissociation (α)
\n
$$
= \sqrt{\frac{K_{a}}{C}} = \sqrt{\frac{1 \times 10^{-5}}{0.100}} = 10^{-2} = 1\%
$$

\nI. The required equation for the oxidation of NH₃ by oxygen to give NO is :
\n2NH₃ + $\frac{5}{2}$ O₂ $\frac{Pt(gauze)}{800°C} \times 2NO + 3H_{2}O$
\nFor this K = $\frac{[NO]^2 [H_{2}O]^3}{[NH_{3}]^2 [O_{2}]^{5/2}}$
\nFor getting the K we must do
\n $K_{1}^{2} = \frac{[N_{2}O]^6}{[N_{2}]^2 [H_{2}]^6}$; $K_{2}^{2} = \frac{[NO]^4}{[N_{2}]^2 [O_{2}]^2}$
\n $K_{3}^{6} = \frac{[H_{2}O]^6}{[H_{2}O]^2}$; $K^{2} = \frac{K_{2}^{2} \times K_{3}^{6}}{K_{1}^{2}}$ or $K = \frac{K_{2}K_{3}^{3}}{K_{1}^{2}}$
\nL. Given [H₃O⁺] = 1 × 10⁻¹⁰

$$
\frac{1}{2} O_2(g) \implies NO_2(g)
$$
\n(10) (A). Given $[H_3O^+] = 1 \times 10^{-10} M$
\nAt 25°C $[H_3O^+] = 10^{-14}$

:. [OH⁻] =
$$
\frac{10^{-14}}{10^{-10}} = 10^{-4}
$$

\nNow, [OH⁻] = 10^{-p} ^{OH} = $10^{-4} = 10^{-p}$ ^{OH}
\n:. p^{OH} = 4

e get (11) (C). HI
$$
\frac{1}{\sqrt{2}}H_2 + \frac{1}{2}I_2
$$
; K = 8
then, H₂ + I₂ $\frac{1}{\sqrt{2}}$ 2HI
K = $\left[\frac{1}{8}\right]^2 = \frac{1}{64}$

(12) (C). 3 3 **=** 3.7 × 10–4 M **(13) (D).** X Y + Z 1 0 0 Initial mole 1 – Mole at equilibrium Y Z p ¹ ^X 1 p 1 p **(17) (D).** (CH³)³ B – is electron deficient p p 1 1 ^K 1 2 1 p 1 2 ^p ^K 1 (1) For equation, A 2B 1 0 Initial mole 1–2 Mole at equilibrium 2 2 2 2 p 2 2 2 2 p ¹ 4 p ^K 1 1 p 1 (2) From eqs. (1) and (2), p1 1 1 p 2 2 ² ^K p p ⁹ K 4p 1 4p 1 2 p 36 p 1 **(14) (D).** Fe(OH)³ (s) Fe3+(aq) + 3OH– (aq) K^c = [Fe+3] [OH–]³ Since, K^c depends only on temperature 3 3 3 3 i F [Fe][OH] [Fe][OH] x × y³ = x × (y/4)³ 1 x x x 64x 64 **(15) (C).** 2AB² (g) 2AB (g) + B² (g) 1 1 – /2 1 2 p 2 . ² ^P ^K (1) ¹ 2 P

As $\alpha \ll 1$; $1 - \alpha \approx 1$; $2 + \alpha \approx 2$

 $p = \frac{1}{2}$ $\rightarrow \alpha = \frac{1}{2}$

EXECUTE: CHEMISTRY
\n[**H**⁺] = **M** =
$$
\frac{10^{-3} \times 1 + 10^{-4} \times 1 + 10^{-5} \times 1}{3}
$$

\n= $\frac{10^{-3}[1 + 10^{-1} + 10^{-2}]}{3} = \frac{10^{-3}[1.11]}{3} = \frac{1.11 \times 10^{-3}}{3}$
\n= $\frac{3.7 \times 10^{-4} \text{ M}}{3}$
\n= $\frac{3.7 \times 10^{-4} \text{ M}}{4}$
\n= $\frac{6 \times \text{p}_1}{\text{m}^2}$
\n= $\frac{6 \times \text{p}_1}{\text{p}_1} = \frac{\frac{\alpha \times \text{p}_1}{1 + \alpha} \cdot \frac{\alpha}{\alpha}}{\frac{1 - \alpha}{\alpha}} = \frac{\frac{\alpha \times \text{p}_1}{1 + \alpha \cdot \frac{\alpha}{\alpha}}}{\frac{\alpha}{\alpha}}$
\n= $\frac{\frac{\alpha^2 \text{p}_1}{1 + \alpha}}{\frac{\alpha \times \text{p}_1}{\alpha}} = \frac{\frac{\alpha}{\alpha} \times \text{p}_1}{\frac{\alpha}{\alpha}} = \frac{\frac{\alpha}{$

(18) (D).
$$
K_h = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.77 \times 10^{-5}} = 5.65 \times 10^{-10}
$$

(19) (D). No. of m. equivalent of HCl = $20 \times 0.05 = 1.0$ No. of m. equivalent of Br $(OH)_2 = 30 \times 0.1 \times 2 = 6.0$ After neutralization, no. of milli equivalents in 50 ml. of solution = $(6-1)$ = 5 No. of m. equivalent of OH⁻ is 5 in 50 ml N]
 $\frac{1}{1}$
 $\frac{1}{1}$ CM₃COO⁻+HCN

COO⁻¹] = $\frac{K_1}{K_2}$ = 3.33 × 10⁴

selectron deficient
 $\frac{1 \times 10^{-14}}{1.77 \times 10^{-5}}$ = 5.65 × 10⁻¹⁰

onivalent of HCl = 20 × 0.05 = 1.0

onivalent of Br (OH)₂ = 30 × 0.1 I (<u>→</u> CH₃COO⁻+HCN

COO⁻] = $\frac{K_1}{K_2}$ = 3.33 × 10⁴

electron deficient
 $\frac{1 \times 10^{-14}}{1.77 \times 10^{-5}}$ = 5.65 × 10⁻¹⁰

iivalent of HCl = 20 × 0.05 = 1.0

iivalent of Br (OH)₂ = 30 × 0.1 × 2 = 6.0

zation

$$
[OH^-] = \frac{5}{50} = 0.1 M
$$

(20) (D). Given, pH of Ba(OH)₂ = 12 \therefore [H⁺] = [1 × 10⁻¹²]

and [OH⁻] =
$$
\frac{1 \times 10^{-14}}{1 \times 10^{-12}}
$$
 = 1 × 10⁻² mol/L

 \therefore [H⁺] [OH⁻] = 1 × 10⁻¹⁴]

K =
$$
\frac{K}{[CN^-][CH_3COOH]} = \frac{K_1}{K_2} = 3.33 \times 10^4
$$

\n(D). (CH₃)₃ B – is electron deficient
\n(D). K_h = $\frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.77 \times 10^{-5}} = 5.65 \times 10^{-10}$
\n(D). No. of m. equivalent of HCl = 20 × 0.05 = 1.0
\nNo. of m. equivalent of Br (OH)₂ = 30 × 0.1 × 2 = 6.0
\nAfter neutralization, no. of milli equivalents in 50 ml
\nof solution = (6 – 1) = 5
\nNo. of m. equivalent of OH⁻ is 5 in 50 ml
\n[OH⁻] = $\frac{5}{50} = 0.1$ M
\n(D). Given, pH of Ba(OH)₂ = 12
\n \therefore [H⁺] = [1 × 10⁻¹²]
\nand [OH⁻] = $\frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2}$ mol/L
\n \therefore [H⁺] [OH⁻] = 1 × 10⁻¹⁴]
\nBa(OH)₂ → Ba²⁺ + 2OH⁻
\ns
\nX_{sp} = [Ba²⁺] [OH⁻]² = [s][2s]²
\n= $\left(\frac{1 \times 10^{-2}}{2}\right)$ (1 × 10⁻²)²
\n= 0.5 × 10⁻⁶ = 5.0 × 10⁻⁷ M³
\n(D). (i) For basic buffer
\npOH = pK_b + log $\frac{[salt]}{[base]}$
\n(ii) pH + nOH = 14

$$
pOH = pK_b + log \frac{[salt]}{[base]}
$$

(ii) pH + pOH = 14
Given, K_b = 1 × 10⁻¹⁰, [salt] = [base]

$$
pOH = -\log K_b + \log \frac{[salt]}{[base]}
$$

1;
$$
1 - \alpha \approx 1
$$
; $2 + \alpha \approx 2$
\n
$$
{}^{3}P_{\text{max}} \left(2K_{p}\right)^{1/3}
$$

 (21)

 $K_p = \frac{\alpha^3 P}{2} \Rightarrow \alpha = \left(\frac{2K_p}{R}\right)^{1/3}$

(22) (D). CH_3COOH (weak acid) and CH_3COONa (conjugated salt) form acidic buffer and for acidic buffer,

$$
pH = pK_a + log \frac{[salt]}{[acid]}
$$

and $[H^+] = - antilog pH$

$$
pH = -\log K_a + \log \frac{[salt]}{[acid]} [\because pK_a = -\log K_a]
$$

$$
= -\log(1.8 \times 10^{-5}) + \log \frac{0.20}{0.10}
$$

= 4.74 + log 2 = 4.74 + 0.3010 = 5.041 (29) (D). Bl
Now, [H⁺] = antilog (-5.045) = 9.0 × 10⁻⁶ mol/L (30) (A). (A

Alternative :

CH₃COOH
$$
\Longleftarrow
$$
 CH₃COO⁻ + H⁺
C-x x x
CH₃COONa \Longleftarrow CH₃COO⁻ + Na⁺
0.2 M 0.2 M

QUIBRICIM	Q.B- SOLUTIONS	ES-DLUTIONS	ES-DLUTIONS						
(D) . CH ₃ COOH (weak acid) and CH ₃ COONa (conjugated salt) from acidic buffer and for acidic buffer, pH = pK _a + log $\frac{[salt]}$	(27)	(A) . $\Delta n_g = -ve$ and $\Delta H = -ve$ The forward reaction is favored at high pressure and low temperature. pH = -log K _a + log $\frac{[salt]}$	(27)	(A) . $\Delta n_g = -ve$ and $\Delta H = -ve$ The forward reaction is favored at high pressure and low temperature. pH = -log K _a + log $\frac{[salt]}$	(28)	(C) . pOH = pK _b + log $\frac{[Balt]}{[Base]}$ = 4.74 + log $\frac{0.20}{0.30}$ = -log (1.8×10 ⁻⁵) + log $\frac{0.20}{0.10}$ = 4.74 + log 2 = 4.74 + 0.3010 = 5.041	(29)	(D) . BF ₃ is an electron deficient species. Now, [H ⁺] = antilog (-5.045) = 9.0 × 10 ⁻⁶ mol/L. 2Hernative: CH ₃ COOH \rightleftharpoons CH ₃ COO ⁺ + H ⁺ base & 4.56 = 9.44 of catalon catalon (29)	(A) . (ACl ₃ , LiCl & BeCl ₃) all these solutions are acidic due to catalon (214) of weakness with the correct subject (H ⁺) and base (OH ⁺), CH ₃ COONa \rightleftharpoons CH ₃ COO ⁻⁺ H ⁺ C-x x x 0.2 M QAM = 0.2 M 0.2 M QAM = 0.2 M 0.2 M QAM =

$$
\therefore \quad [H^+] = \frac{K_a [CH_3COOH]}{[CH_3COO^-]} = \frac{1.8 \times 10^{-5} \times 10^{-1}}{2 \times 10^{-1}} = 9 \times 10^{-6} \text{ M}
$$
\n(23) **(D).** $K_p = K_c (RT)^{\Delta n}$.
\nFor 2C(s) + O₂(g) \rightleftharpoons 2CO₂(g);
\n(d35) **(C).** $8H_3$ is a Lewis
\n**(35) (C).** Na₂CO₃ is a sp

 $\Delta n = 2 - 1 = 1 \neq 0$ **(24) (B).** $2A(g) + B(g) \rightleftharpoons 3C(g) + D(g)$ Initial moles : $1 \t 1 \t 0 \t 0$ Moles of eq. : $1-(2\times0.25)$ $1-0.25$ 3×0.25 0.25 $= 0.5$ $= 0.75$ $= 0.75$ $= 0.25$ Equilibrium constant,

$$
K = \frac{[C]^3 [D]}{[A]^2 [B]} = \frac{(0.75)^3 (0.25)}{(0.5)^2 (0.75)}
$$

- **(25) (C).** When $K_p > Q$, rate of forward reaction > rate of backward reaction.
	- \therefore Reaction is spontaneous. When ΔG° < RT ln Q, ΔG° is positive, reverse reaction is feasible, thus reaction is non-spontaneous.
When $K = O$ rate of forward reaction When $K_p = Q$, rate of forward reaction = rate of backward reaction
	- \therefore Reaction is in equilibrium. When $T\Delta S > \Delta H$, ΔG will be –ve only when $\Delta H = +ve.$

(26) **(D).** N₂(g) + O₂(g)
$$
\Longleftrightarrow
$$
 2NO(g); K₁
2NO(g) + O₂(g) \Longleftrightarrow 2NO₂(g); K₂

$$
N_2(g) + 2O_2(g) \xrightarrow{\longrightarrow} 2NO_2(g); K = K_1 \times K_2
$$

$$
\therefore \text{ For NO}_2(g) \xrightarrow{\longrightarrow} \frac{1}{2} N_2(g) + O_2(g); \ \ K' = \left[\frac{1}{K_1 K_2}\right]^{1/2}
$$

(27) (A). $\Delta n_g = -ve$ and $\Delta H = -ve$ The forward reaction is favoured at high pressure and low temperature. **EDEMADVANCED LEARNING**
 $+ O_2(g)$; $K' = \left[\frac{1}{K_1K_2}\right]^{1/2}$

voured at high pressure
 $4.74 + log \frac{0.20}{0.30}$
 $4-0.176 = 4.56$

species.

see solution are acidic due

re BaCl₂, is salt of strong **SPON ADVANCED LEARNING**

(b) + O₂(g); $K' = \left[\frac{1}{K_1K_2}\right]^{1/2}$

favoured at high pressure

= 4.74 + log $\frac{0.20}{0.30}$

1.74 – 0.176 = 4.56

nt species.

(28) (C). pOH = pK_b + log
$$
\frac{[Salt]}{[Base]}
$$
 = 4.74 + log $\frac{0.20}{0.30}$
= 4.74 + (0.301 – 0.477) = 4.74 – 0.176 = 4.56
: pH = 14 – 4.56 = 9.44

- **(29) (D).** BF_3 is an electron deficient species.
- **(30) (A).** (AlCl₃, LiCl & BeCl₂) all these solution are acidic due to cationic Hydrolysis, where $BaCl₂$, is salt of strong base & strong acid.
- **(31)** \ldots On adding small amount of acid (H^+) and base(OH⁻), weak acid or weak base will be formed respectively.

(32) (C).
$$
2SO_2 + O_2 \rightleftharpoons 2SO_3
$$
; K = 278

0.13 and CH₃COONa (conjugated
\nr and for acidic buffer,
\nr and for acidic buffer,
\nH
\nH
\n
$$
\begin{array}{ll}\n\text{(27)} & \text{(A) } \Delta n_{g} = -ve \text{ and } \Delta H = -ve \\
\text{The forward reaction is favoured at high pressure and lower temperature.} \\
\text{and by temperature.} \\
\text{and } \text{by temperature.} \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0.176 = 4.56 \\
\text{by } \frac{0.20}{0.10} &= 4.74 + (0.301 - 0.477) = 4.74 - 0
$$

(33) (C).
$$
A_2(g) + B_2(g) \xrightarrow{ } 2AB(g)
$$

$$
K_C = \frac{(2.8 \times 10^{-3})^2}{3 \times 10^{-3} \times 4.2 \times 10^{-3}} = \frac{(2.8)^2}{3 \times 4.2} = 0.62
$$

(D). BF_3 is a Lewis acid.

- **(35) (C).** Na_2CO_3 is a salt of strong base and weak acid will produce a basic solution with $pH > 7$.
- (27) (A). \ln = -ve and All = -ve
 \ln be very and catation is favoured at high pressure
 $\left[\because PK_n = -\log K_n\right]$

(28) (C). $pOH = pk_0 + \log \frac{[Salt]}{[Base]} = 4.74 + \log \frac{0.20}{0.30}$
 $= 4.74 + \log \frac{0.201}{0.30}$
 $= 4.74 + \log \frac{0.201}{0.30}$ CH₃COO⁻][H⁺]

SO₃ $\xleftarrow{3}$ 3 5O₂ $+ \frac{1}{2}$ O₂

(CH₃COOH]

SOH₃COOH]

SOH₃COOH] = C-x ≈ 0.1 M

SOH₃COOH] = C-x ≈ 0.1 M

COL 3 + x ≈ 0.2 M

did is a weak acid so, dissociation is minimum]

H₃COOH] = $\frac{243 \times 200 \text{ J} [H^3]}{100 \text{ H} [H_3\text{COO}]}$
 $\frac{243 \times 10^{-4} \text{ kg}}{10^{-1} \text{ kg}^2} = \sqrt{35.97 \times 10^{-4}} = 6$
 $\frac{243 \times 10^{-5} \times 10^{-1}}{2 \times 10^{-1}} = 9 \times 10^{-6} \text{ M}$
 $\frac{243 \times 10^{-5} \times 10^{-1}}{2 \times 10^{-1}} = 9 \times 10^{-6} \text{ M}$
 $\frac{243 \times 10^{-3$ Let $C_1C_2COH \rightleftharpoons CH_2COO+H^2$

C-x $K = \frac{1}{2}$ CH₂COO^{-+ AR} (31) the season in a monuto fracid (H⁺¹) and base(OH-),

C-x $K = \frac{1}{2}$ CH₂COO + Na⁺ (32) (C). 2SO₂ + O₂ weak a base will be formed respectively.
 after
 $\frac{1}{25}$ (GOOR $\frac{1}{25}$ CH₂COO⁻⁺ Hip (exception) is a calconing small amount of acaditing the set of the particular terms of the set of the s **(36) (D).** $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3$ It is an exothermic reaction so decreasing temperature is favorable because dissociation $(x) \propto 1/T$ & dissociation $x \propto (1/P)^{-2/2}$ $\therefore x \propto P$ So increasing pressure is favorable for forward reaction $\sqrt{\frac{1}{278}} = \sqrt{35.97 \times 10^{-4}} = 6 \times 10^{-2}$
 $+ B_2(g) \rightleftharpoons 2 AB (g)$
 $\frac{(2.8 \times 10^{-3})^2}{3 \times 10^{-3} \times 4.2 \times 10^{-3}} = \frac{(2.8)^2}{3 \times 4.2} = 0.62$

sa Lewis acid.

O₃ is a salt of strong base and weak acid will

O₃ is a salt of stro $\sqrt{\frac{1}{278}} = \sqrt{35.97 \times 10^{-4}} = 6 \times 10^{-2}$
 $+ B_2(g) \xrightarrow{ } 2 AB(g)$
 $\frac{(2.8 \times 10^{-3})^2}{3 \times 10^{-3} \times 4.2 \times 10^{-3}} = \frac{(2.8)^2}{3 \times 4.2} = 0.62$

sa Lewis acid.
 O_3 is a salt of strong base and weak acid will
 \log is a salt of stron On adding small amount of acid (H⁺) and base(OH⁻),

weak acid or weak base will be formed respectively.
 $2SO_2 + O_2 \rightleftharpoons 2SO_3$; $K = 278$
 $SO_2 \leftarrow O_2 \rightleftharpoons 2SO_3$; $K = 278$
 $SO_3 \rightleftharpoons SO_2 + \frac{1}{2}O_2$
 $K' = \sqrt{\frac{1}{278}} = \sqrt$ ionic Hydrolysis, where BaCl₂, is salt of strong

& trong acid.

diding small amount of acid (H⁺) and base(OH⁻),

acid or weak base will be formed respectively.
 $2+O_2 \xrightarrow{Q_2} 2SO_3$; K = 278
 \longrightarrow $SO_2 + \frac{1}{2}O_2$
 strong acid.
 $2\sqrt{10}$

ing small amount of acid (H⁺) and base(OH⁻),

id or weak base will be formed respectively.
 $-62 \xrightarrow{ } 250_3$; $K = 278$
 $\Rightarrow SO_2 + \frac{1}{2}O_2$
 $\frac{1}{278} = \sqrt{35.97 \times 10^{-4}} = 6 \times 10^{-2}$
 $+ B_2(g) \xrightarrow{$ ⇒ 2 AB (g)
 $\frac{0^{-3}y^2}{4.2 \times 10^{-3}} = \frac{(2.8)^2}{3 \times 4.2} = 0.62$

id.

of strong base and weak acid will

olution with pH > 7.

⇒ 2NH₃

reaction so decreasing temperature

use dissociation (x) ∞ 1/T

∞ (1/P)^{-2/2} ∴ x ∞ P
 $\frac{(0^{-3})^2}{4.2 \times 10^{-3}} = \frac{(2.8)^2}{3 \times 4.2} = 0.62$

id.

id.

condition with pH > 7.
 \Rightarrow 2NH₃

creaction so decreasing temperature

use dissociation (x) $\propto 1/T$
 $\propto (1/P)^{-2/2}$ \therefore x $\propto P$

ressure is favorable for f
	- **(37) (A).** According to Vant Hoff equation

$$
\log \frac{K_p'}{K_p} = \frac{\Delta H}{2.303R} \left[\frac{T_2 - T_1}{T_1 T_2} \right]
$$

For exothermic reaction on increasing temperature equilibrium constant decreases means $T_2 > T_1$ then $K'_p < K_p$

(38) (B). K =
$$
1.6 \times 10^{12}
$$
 = $\left[\frac{\text{Product}}{\text{Reactant}} \right]$

K have high value so products have very high conc. than reactant.

(39) (C). Ksp of Ag2CrO⁴ = 1.1 × 10–12 ^Ksp of AgCl = 1.8 × 10–10 ^Ksp of AgBr = 5.0 × 10–13 ^Ksp of AgI = 8.3 × 10–17

$$
[Ag^+]^2 [C] = 1.1 \times 10^{-12}
$$

$$
[Ag^{+}] = \sqrt{\frac{1.1 \times 10^{-12}}{[C]}}
$$

If we take $|C| = 1$ then the maximum requirement of $[Ag^+]$ will be in case of Ag_2CrO_4 .

(40) (C). $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$; K

$$
\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g); K'
$$

When a reaction is multiplied by 1/2 then $K' = (K)^{1/2}$. .

- **(41) (C).** HClO₄ is a strong acid. Acidic buffer is a mixture of a weak acid and its salt with a strong base.
- **(42) (D).** Aqueous solution of HCl is the best conductor of electric current because HCl is strong acid, so it dissociates completely into ions.
- **(43) (C).** One mole of NaOH is completely neutralised by one mole of HCl. Hence, 0.01 mole of NaOH will be completely neutralised by 0.01 mole of HCl.

NaOH left unneutralised = $0.1 - 0.01$ mol = 0.09 mol (4)

As equal volumes of two solutions are mixed,

$$
[OH] = \frac{0.09}{2} = 0.045 M
$$

- \Rightarrow pOH = log (0.045) = 1.35
- \therefore pH = 14 1.35 = 12.65
- **(44) (B).** For MY, $K_{SP} = S^2$ $\therefore S = (6.2 \times 10^{-13})^{1/2}$

pOH = -log (0.045) = 1.35
\npH = 14 - 1.35 = 12.65
\nFor MY,
$$
K_{SP} = S^2
$$
 $\therefore S = (6.2 \times 10^{-13})^{1/2}$
\nFor NY₃, $K_{SP} = 27 S^4$ $\therefore S = \left(\frac{6.2 \times 10^{-13}}{27}\right)^{1/4}$
\n $\left(\frac{6.2 \times 10^{-1$

(45) **(D).** Vant Hoff's equation is
$$
\frac{\Delta H}{RT^2} = \frac{d \left(\ln K \right)}{dT}
$$
 Meq of NaOH = 2

$$
\log \frac{K_2}{K_1} = \frac{\Delta H^{\circ}}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1} \right]
$$

(46) **(B).**
$$
C_5H_5N + H_2O \rightleftharpoons C_5H_5N^{\oplus}H + OH^{\oplus}
$$

t=0 C

$$
t = eq C (1 - \alpha) \qquad C\alpha \qquad C\alpha
$$

$$
K_{b} = \frac{C\alpha^{2}}{1 - \alpha} = C\alpha^{2}
$$

\n
$$
\alpha = \sqrt{\frac{K_{b}}{C}} = \sqrt{\frac{1.7 \times 10^{-9}}{0.1}} = 1.3 \times 10^{-4}
$$

\nBaS

IDENTIFY
\nAg⁺]²(C) = 1.1 × 10⁻¹²
\n(Ag⁺) =
$$
\sqrt{\frac{1.1 \times 10^{-12}}{[C]}}
$$

\n(Ag⁺) = $\sqrt{\frac{1.1 \times 10^{-12}}{[C]}}$
\n[Ag⁺] = $\sqrt{\frac{1.1 \times 10^{-12}}{[C]}}$
\n[Ag⁺] = $\sqrt{\frac{1.1 \times 10^{-12}}{[C]}}$
\n[Ag⁺] will be in case of Ag₂CrO₄.
\n[Ag⁺] will be in case of Ag₂CrO₄.
\n[Ag⁺] will be in case of Ag₂CrO₄.
\n $\frac{1}{2}N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$; K
\n \therefore S <0.1 \therefore S + 0.1 ≈ 0.1
\n $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g)$; K
\n \therefore 1.6 × 10⁻¹⁰ = S × 0.1 \therefore S = 1.6 × 10⁻⁹ M
\nWhen a reaction is multiplied by 1/2 then K' = (K)^{1/2}.
\n(HClO₄ is a strong acid.
\n4.21 × 10⁻¹⁰ = 10⁻¹⁰ J.
\nHClO₄ is a strong acid.
\n4.22 × 10⁻¹⁰ J.
\n3.23 × 10⁻⁴ J.
\n3.24 × 10⁻¹⁰ J.
\n3.25 × 10⁻¹¹ J.
\n4.26 × 10⁻¹⁰ J.
\n4.27 × 10⁻¹⁰ J.
\n4.28 × 10⁻¹¹ J.
\n4.29 × 10⁻¹¹ J.
\n5.20 × 10⁻¹¹ J.
\

$$
\therefore \quad 1.6 \times 10^{-10} = S \times 0.1 \quad \therefore \quad S = 1.6 \times 10^{-9} \, \text{M}
$$

(48) (A). N₂ + 3H₂
$$
\rightleftharpoons
$$
 2NH₃ ; K₁ \rightarrow (1)
\nN₂ + O₂ \rightleftharpoons 2NO ; K₂ \rightarrow (2)
\nH₂ + ½ O₂ \rightarrow H₂O ; K₃ \rightarrow (3)
\nFor reaction

$$
2NH_3 + (5/2)O_2 \xrightarrow{K} 2NO + 3H_2O \to (4)
$$

Equation (4) = equation (2) + 3 × eq. (3) – equation (1)

$$
K = \frac{K_2 \cdot K_3^3}{K_1}
$$

$$
\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g); K'
$$
\n
$$
\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g); K'
$$
\n
$$
2N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g); K'
$$
\n
$$
N_2 + O_2 \rightleftharpoons 2NN_3
$$
\n
$$
N_2 + O_2 \rightleftharpoons 2N_3
$$
\n
$$
N_2 + O_2 \rightleftharpoons 2N
$$

(50) (A). Equilibrium constant is not affected by presence of catalyst hence statement (A) is incorrect.

$$
= \frac{0.09}{2} = 0.045 M
$$

\n
$$
= -\log (0.045) = 1.35
$$

\n
$$
= 12.65
$$

\n
$$
= \log (0.045) = 1.35
$$

\n
$$
= 12.65
$$

\

 $\begin{bmatrix} 27 \\ 27 \end{bmatrix}$ **(52) (D).** Meq of HCl = $75 \times \frac{1}{5} \times 1 = 15$

$$
\frac{\Delta H}{RT^2} = \frac{d \, (\ln K)}{dT}
$$
\n
$$
Meq \, of NaOH = 25 \times \frac{1}{5} \times 1 = 5
$$

Meq of HCl in resulting solution $= 10$

Molarity of [H⁺] in resulting mixture =
$$
\frac{10}{100} = \frac{1}{10}
$$

$$
pH = -\log[H^{+}] = -\log\left[\frac{1}{10}\right] = 1.0
$$

(53) (C). Solubility of BaSO₄,

[OH]⁻ =
$$
\frac{0.09}{2} = 0.045
$$
 M
\n
$$
pOH = -log(0.045) = 1.35
$$
\n(50) (A) Equilibrium constant is not affected by presence of
\nthe 14 - 1.35 = 12.65
\nFor MY, $K_{SP} = S^2$ $\therefore S = (6.2 \times 10^{-13})^{1/2}$
\nFor NY₃, $K_{SP} = 27 S^4$ $\therefore S = \left(\frac{6.2 \times 10^{-13}}{27}\right)^{1/4}$
\nFor NY₃, $K_{SP} = 27 S^4$ $\therefore S = \left(\frac{6.2 \times 10^{-13}}{27}\right)^{1/4}$
\nFor NY₃, $K_{SP} = 27 S^4$ $\therefore S = \left(\frac{6.2 \times 10^{-13}}{27}\right)^{1/4}$
\nFor NY₃, $K_{SP} = 27 S^4$ $\therefore S = \left(\frac{6.2 \times 10^{-13}}{27}\right)^{1/4}$
\n(52) (D). Meq of HCl = $75 \times \frac{1}{5} \times 1 = 15$
\n
$$
log \frac{K_2}{K_1} = \frac{\Delta I I^{\circ}}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1}\right]
$$
\n
$$
log \frac{K_2}{K_1} = \frac{\Delta I I^{\circ}}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1}\right]
$$
\n
$$
log \frac{K_2}{K_2} = \frac{\Delta I I^{\circ}}{2.303R} \left[\frac{T_2 - T_1}{T_2 T_1}\right]
$$
\n
$$
C_{S}H_{S}N^{\oplus}H + OH^{\ominus}
$$
\n
$$
PH = -log [H^+] = -log \left[\frac{1}{10}\right] = 1.0
$$
\n
$$
qC(1 - \alpha) C\alpha C\alpha
$$
\n(53) (C) Solubility of BaSO₄,
\n
$$
s = \frac{2.4
$$

$$
K_{sp} = [Ba^{2+}] [SO_4^{2-}] = s^2
$$

= (1.04 × 10⁻⁵)² = 1.08 × 10⁻¹⁰ mol² L⁻²

(54) (C). $A_2(g) + B_2(g) \rightleftharpoons X_2(g) \Delta_f H = -X kJ$?

On increasing pressure equilibrium shifts in a direction where pressure decreases i.e. forward direction. On decreasing temperature, equilibrium shifts in exothermic direction i.e., forward direction.

So, high pressure and low temperature favours maximum formation of product.

(55) (A). Ca(OH)₂
$$
\rightleftharpoons
$$
 Ca²⁺ + 2OH⁻

pH = 9 Hence, pOH = 14 - 9 = 5; [OH⁻] = 10⁻⁵ M Before 10
\n
$$
\times 0
$$
\nHence [Ca²⁺] = $\frac{10^{-5}}{2}$ = 10

After
\n
$$
K_{sp} = [Ca^{2+}][OH^-]^2 = \left(\frac{10^{-5}}{2}\right)(10^{-5})^2 = 0.5 \times 10^{-15}
$$
\n
$$
Hy
$$
\n
$$
Th
$$

HF on loss of H^+ ion becomes F^- is the conjugate At base of HF

Example :

 $HF + H_2O \rightleftharpoons F^- + H_3O^+$

Acid Base Conjugate Conjugate base acid

