Equilibrium.

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Chemical Equilibrium

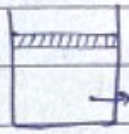
A State in which the rate of the forward reaction equal to the rate of backward reaction.

Ionic Equilibrium.

The equilibrium establish b/w the unionised molecules and the ions in the Solution of weak electrolytes.

(~face change
(Solid ⇌ Liqui

Chemical Equilibrium.

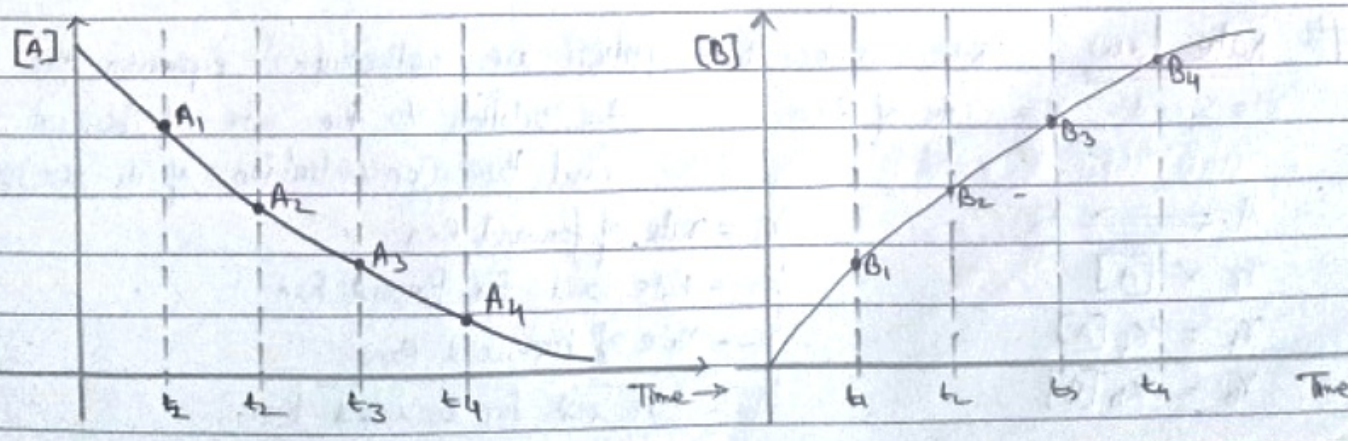


$r_f = r_b$

[ ] ← Concentration
↑
Molarity
↑
SI unit = mol/lit

$$CaCO_{3(s)} \xrightarrow{\Delta} CaO_{(s)} + CO_{2(g)} \uparrow$$

	A ⟶	B
At $t=0$	$A_0 = 100\%$	$B = 0\%$
$t = t_1$	A_1	B_1
$t = t_2$	A_2	B_2
$t = t_3$	A_3	B_3
$t = t_4$	A_4	B_4



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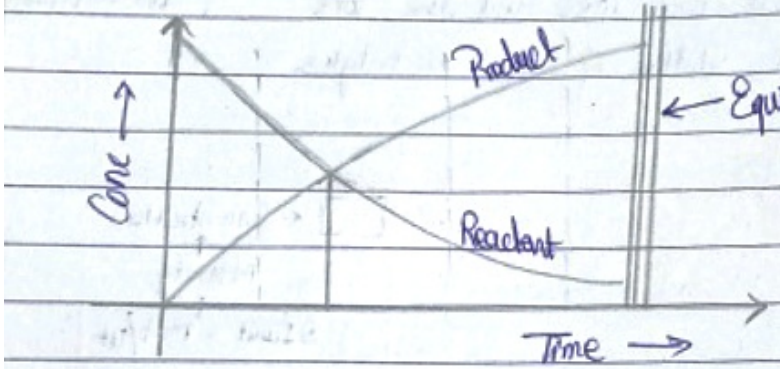
✯ Active Mass

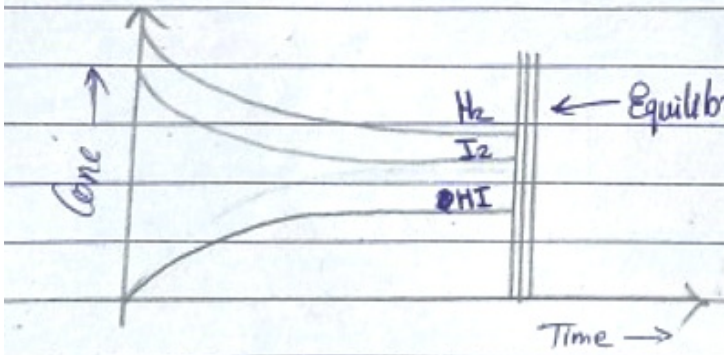
Active mass term means the concentration of the reactant which is represent in mol per lit.



Active mass of Pure liquid + Solid = 1

Solid or liquid = moles / Vol. = mass / (molar mass × Vol.) = density / molar mass = Constant

Graph of ⇒ $H_2 + I_2 \longrightarrow 2HI$



≠ Rate Law

Rate law or Rate equation are mathematical expression the discribe the relation b/w the rate of chemical reaction and the Concentration of its reaction.

r = Speed, r = rate of Rxn

$A \rightleftharpoons B$

$r_f \propto [A]$

$r_f = k_f [A]$

$r_b = k_b [B]$

r_f = rate of forward Rxn

k_f = Rate Const. for forward Rxn

r_b = rate of backward Rxn

k_b = Rate Const. for backward Rxn

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Q. K_f is higher than K_b Why?

⇒ The forward reaction is more likely to happen due to the lower energy of product.

① $2A \rightleftharpoons 3B$

$r_f = k_f [A]^2$; $r_b = k_b [B]^3$

② $H_2 + I_2 \rightleftharpoons 2HI$

$r_f = k_f [H_2][I_2]$; $r_b = k_b [HI]^2$

③ $N_2 + 3H_2 \rightleftharpoons 2NH_3$

$r_f = k_f [N_2][H_2]^3$; $r_f = k_f [NH_3]^2$

Equilibrium.

$r_f = r_b$

$aA + bB \rightleftharpoons cC + dD$

$r_f = k_f [A]^a [B]^b$ ——— (I)

$r_b = k_b [C]^c [D]^d$ ——— (II)

$k_f [A]^a [B]^b = k_b [C]^c [D]^d$

$$k_c = \frac{k_f}{k_b} = \frac{[C]^c [D]^d}{[A]^a [B]^b} = \frac{[\text{Product}]^{\text{power}}}{[\text{Reactant}]^{\text{power}}}$$

k_c = Equilibrium Constant for Concentration.

Eg 1 :- $H_2 + I_2 \rightleftharpoons 2HI$

$$k_c = \frac{[HI]^2}{[H_2][I_2]}$$

Eg 2 :- $2NH_3 \rightleftharpoons N_2 + 3H_2$

$$k_c = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

Calculation of K_p ← Equilibrium Const. for pressure.

Only. gas. Reaction.

Product ⟶ Gas.

① Homogeneous Rxn

→ All reactant and Product are in same phase.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

② Hetrogenous Rxn

→ All reactant and product are in diff. phase

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

Dalton gas theory. (Partial Pressure law)

$P_A = x_A \times P_{Total}$

$P_B = x_B \times P_{Total}$

$P_C = x_C \times P_{Total}$

$P_D = x_D \times P_{Total}$

$$aA_{(g)} + bB_{(g)} \rightleftharpoons cC_{(g)} + dD_{(g)}$$

$P_A \quad P_B \quad P_C \quad P_C$

→ The total pressure of a mixture of gases is equal to the sum of the partial pressure of the individual component gases.

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

↑ Equilibrium Constant for pressure

Relation b/w K_c & K_p

$$aA + bB \rightleftharpoons cC + dD$$

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad ; \quad K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

$$PV = nRT$$

$$P = \frac{n}{V} RT$$

$$P = [\,]RT$$

$$P_A = [A]RT$$

$$P_B = [B]RT$$

$$P_C = [C]RT$$

$$P_D = [D]RT$$

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

$$K_p = \frac{([C]RT)^c ([D]RT)^d}{([A]RT)^a ([B]RT)^b}$$

$$K_p = \frac{[C]^c [D]^d (RT)^{c+d}}{[A]^a [B]^b (RT)^{a+b}}$$

$$K_p = K_c (RT)^{(c+d)-(a+b)}$$

$(c+d)$ → Value of Product's mol

$(a+b)$ → Value of Reactant's mol

$\Delta n_g = P - R$

$\boxed{K_p = K_c (RT)^{\Delta n_g}}$

$K_p < K_c$ when Δn_g = -ve

$K_p > K_c$ when Δn_g = +ve

$K_p = K_c$ when $\Delta n_g = 0$

Rxn 1:- $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

$\Delta n_g = 2 - 4$

$= -2$ (-ve)

$\therefore K_p < K_c$

Rxn 2:- $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$K_c = \frac{[HI]^2}{[H_2][I_2]}$

$K_p = \frac{(P_{HI})^2}{(P_{H_2})(P_{I_2})}$

$\Delta n_g = 2 - 2$

$= 0$

$\boxed{K_p = K_c}$

Rxn 3:- $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

$K_c = \frac{[CaO][CO_2]}{[CaCO_3]}$

$K_p = (P_{CO_2})$

$\Delta n_g = 2 - 1$

$= 1$ (+ve)

$\boxed{K_p > K_c}$

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Rxn 4:- $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$

$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

$K_p = \frac{(P_{SO_3})^2}{(P_{SO_2})^2(P_{O_2})}$

$\Delta ng = 2-3$
$= -1$ (-ve)

$\therefore$ $K_p < K_c$

Rxn 5:- $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}$

$K_p = \frac{(P_{PCl_3})(P_{Cl_2})}{(P_{PCl_5})}$

$\Delta ng = 2-1$
$= 1$ (+ve)

$K_p > K_c$

Rxn 6:- $2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$

$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}$

$K_p = \frac{(P_{NO_2})^2}{(P_{NO})^2(P_{O_2})}$

$\Delta ng = 2-3$
$= -1$ (-ve)

$K_p < K_c$

K_x (Equilibrium Const. for mole fraction)

$$aA + bB \rightleftharpoons cC + dD$$
$$x_A \quad x_B \qquad x_C \quad x_D$$

$$K_x = \frac{(x_C)^c\,(x_D)^d}{(x_A)^a\,(x_B)^b}$$

Questions of Comparision

$$\frac{1}{2}H_2 + \frac{1}{2}I_2 \rightleftharpoons HI \qquad K_c = K$$

$$H_2 + I_2 \rightleftharpoons 2HI \qquad K_c = 22$$

$$K_c = (K)^2 = \left(\frac{[HI]}{[H_2]^{1/2}\,[I_2]^{1/2}}\right)^2$$

$$K_c = K_c = \frac{[HI]^2}{[H_2]\,[I_2]}$$

$$\frac{(K)^2}{K_c} = 1$$

$$K^2 = K_c$$

$$K_c = K^2$$

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Q. $2NO + O_2 \rightleftharpoons 2NO_2$ $\quad K_c = 8$

$2NO_2 \rightleftharpoons O_2 + 2NO$ $\quad K_c = ??$

Solve.

$$8 = \frac{[NO_2]^2}{[NO]^2[O_2]} \quad \text{——— (I)}$$

$$K_c = \frac{[NO]^2[O_2]}{[NO_2]^2} \quad \text{——— (II)}$$

Eq^n (I) × Eq^n (II)

$8 \times K_c = 1$

$K_c = \frac{1}{8}$ ✓

Q. $N_2 + 3H_2 \rightleftharpoons 2NH_3$ $\quad K_c = 64$

$NH_3 \rightleftharpoons \frac{1}{2}N_2 + \frac{3}{2}H_2$ $\quad K_c = ??$

Solve

$$64 = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$(K_c)^2 = \left(\frac{[N_2]^{1/2}[H_2]^{3/2}}{[NH_3]}\right)^2$$

$64 \times (K_c)^2 = 1$

$(K_c)^2 = \frac{1}{64}$

$K_c = \frac{1}{8}$ ✓

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Law of mass action — the rate of a reaction is proportional to the Product of the Concentration of each reactant

$$aA + bB \rightleftharpoons cC + dD$$

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad ; \quad K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

Eg 1:- given, A = 1 mol
B = 1 mol

$$2A + 1B \rightleftharpoons 3C$$

		2A	+ 1B	⇌ 3C
At	t=0	1	1	0
At	t=t	1-2x	1-x	3x

Eg 2:- given A = 2 mol
B = 3 mol

$$3A + 4B \rightleftharpoons 6C$$

		3A	+ 4B	⇌ 6C
At	t=0	2	2	0
	t=t	2-3x	3-4x	6x

Eg 3:- If are in Stoichiometry $3A + 6B \rightleftharpoons 7C$

	3A	+ 6B	⇌ 7C
t=0	3	6	0
t=t	3-3x	6-6x	7x

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Eg 4:- given, $N_2 = 0.5\,M$ $\quad N_2 + 3H_2 \rightleftharpoons 2NH_3$
$H_2 = 0.5\,M$

	N_2	$+\ 3H_2 \rightleftharpoons$	$2NH_3$
At $t=0$	0.5	0.5	0
$t=t$	$0.5-\alpha$	$0.5-\alpha$	2α

Eg 5:- Stociometry

$V=V$
$P=P$
$\alpha=\alpha$

	A	$+\ B \rightleftharpoons$	$2C$
$t=0$	1	1	0
$t=t$	$1-\alpha$	$1-\alpha$	2α
Conc	$\frac{1-\alpha}{V}$	$\frac{1-\alpha}{V}$	$\frac{2\alpha}{V}$
P	$\frac{1-\alpha}{2}P$	$\frac{1-\alpha}{2}P$	$\frac{2\alpha}{2}P$

$= \frac{1-\alpha+1-\alpha+2\alpha}{=2}$

$$K_C = \frac{[C]^2}{[A][B]} = \frac{\left(\frac{2\alpha}{V}\right)^2}{\left(\frac{1-\alpha}{V}\right)\left(\frac{1-\alpha}{V}\right)} = \frac{4\alpha^2}{(1-\alpha)^2}$$

$$K_p = \frac{(P_C)^2}{(P_A)(P_B)} = \frac{\left(\frac{2\alpha}{2}\cdot P\right)^2}{\left(\frac{1-\alpha}{2}\cdot P\right)\left(\frac{1-\alpha}{2}\cdot P\right)} = \frac{4\alpha^2}{(1-\alpha)^2}$$

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Eg6:- $A + B \rightleftharpoons C$

Eg: $PCl_3 + Cl_2 \rightleftharpoons PCl_5$

$t=0$	1	1	0	
Conc.	$\frac{1-\alpha}{V}$	$\frac{1-\alpha}{V}$	$\frac{\alpha}{V}$	$= 1-\alpha+1-\alpha+\alpha$ $= 2-\alpha$
P	$\frac{1-\alpha}{2-\alpha}P$	$\frac{1-\alpha}{2-\alpha}P$	$\frac{\alpha}{2-\alpha}P$	

$$K_c = \frac{[C]}{[A][B]} = \frac{\frac{\alpha}{V}}{\left(\frac{1-\alpha}{V}\right)\left(\frac{1-\alpha}{V}\right)} = \frac{\alpha V}{(1-\alpha)^2}$$

$$K_p = \frac{(P_C)}{(P_A)(P_B)} = \frac{\left(\frac{\alpha}{2-\alpha}\cdot P\right)}{\left(\frac{1-\alpha}{2-\alpha}\cdot P\right)\left(\frac{1-\alpha}{2-\alpha}\cdot P\right)} = \frac{\alpha(2-\alpha)}{(1-\alpha)^2 P}$$

Eg 7:-

let
V = 5 lit
P = 2 lit
$\alpha = 40\%$
A = 1 mol
B = 1 mol

$A + B \rightleftharpoons 2C$

$t=0$	1	1	0	
$t=t$	$1-\alpha$	$1-\alpha$	2α	
Conc.	$\frac{1-\alpha}{V}$	$\frac{1-\alpha}{V}$	$\frac{2\alpha}{V}$	
P.	$\frac{1-\alpha}{2}P$	$\frac{1-\alpha}{2}P$	$\frac{2\alpha}{2}P$	

$$K_c = \frac{[C]^2}{[A][B]} = \frac{\left(\frac{2\alpha}{V}\right)^2}{\left(\frac{1-\alpha}{V}\right)\left(\frac{1-\alpha}{V}\right)} = \frac{4\alpha^2}{(1-\alpha)^2}$$

$$= \frac{4 \times (0.4)^2}{(1-0.4)^2} = \frac{16}{9} = 1.77$$

1.77 = $K_c = K_p$ bcoz $\Delta ng = 0$

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Q. $N_2 + 3H_2 \rightleftharpoons 2NH_3$ $\quad \alpha = 50\% = 0.5$

$N = 1$ mol
$H = 3$ mol
$V = 2$ lit
$P = 5$ atm

	N_2	$3H_2$	$2NH_3$
$t=0$	1	3	0
$t=t$	$1-\alpha$	$3-3\alpha$	2α
Conc.	$\frac{1-\alpha}{V} = 1/4$	$\frac{3-3\alpha}{V} = 3/4$	$\frac{2\alpha}{V} = 1/2$
P	$\frac{1-\alpha}{4-2\alpha} \cdot 5$	$\frac{3-3\alpha}{4-2\alpha} \cdot 5$	$\frac{2\alpha}{4-2\alpha} \cdot 5$
	$\Downarrow 5/6$	$\Downarrow 15/6$	$5/3$

$\Delta n_g = 2 - 4 = -2$

$$K_c = \frac{(1/4)^2}{(1/4)(3/4)^3} = \frac{1}{16} \times \frac{64}{3} \times \frac{16}{9} = \frac{1}{3} = 0.33$$

$$K_p = \frac{(5/6)^2}{(15/6)(5/3)^3} = \frac{25}{36} \times \frac{78}{75} \times \frac{1}{3} = \frac{1}{6}$$

$$K_c = \frac{1}{16} \times \frac{4 \times 4 \times 4 \times 4}{3 \times 3 \times 3} = \frac{16}{27} = 0.59$$

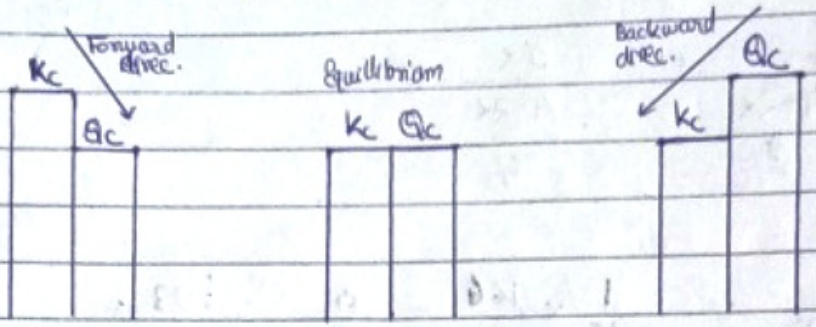
$$K_p = \frac{25}{36} \times \frac{16}{15} \times \frac{3 \times 3 \times 3}{5 \times 5 \times 5} = \frac{9}{150} = 0.06$$

$\Delta n_g = 2 - 4 = -2$ (-ve)

$K_p < K_c$

Lechatelieors principle.

If a dynamic equilibrium is disturbed by changing the condition, the position of equilibrium shifts to counteract the change to reestablish an equilibrium.



$$aA + bB \rightleftharpoons cC + dD$$

① $Q_c \downarrow \Rightarrow Q_c < K_c$ → Equilibrium shift in the forward direction

② $Q_c \uparrow \Rightarrow Q_c > K_c$ → Equilibrium shift in the backward direction

③ $Q_c = K_c$ → Equilibrium Condition.

Eg:- $A + B \longrightarrow 2C$

At $t=t$: 1 mol, 1 mol, 2 mol $\quad K_c = 3$

$$Q_c = \frac{[C]^2}{[A][B]} = \frac{(2)^2}{1 \times 1} = 4$$

$Q_c > K_c$

Eg 2:- $H_2 + I_2 \longrightarrow 2HI$

at $t=t$: 0.5, 0.5, 4 $\quad Q_c = 8$

$$K_c = \frac{(4)^2}{0.5 \times 0.5} = 64 \quad (K_c < Q_c)$$

Factor affecting Equilibrium.

→ Change in the Concentration of Reactant and Product

→ Change in temp.

→ Change in pressure, Vol.

→ Effect of addition of catalyst

→ Effect of addition of inert gas

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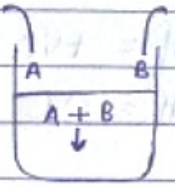
According to Lechatelieers:- If a system are in equilibrium and any change any component Concentration, pressure, Vol., Catalyst, inert gas add. Equilibrium is shifted that way where change are negligible.

① Change in Consentration.

$A + B \rightleftharpoons 2C$



Add reactant ⟶ reaction goes to forward
remove product ⟶ reaction goes to forward.

Q. $2SO_2 + O_2 \rightleftharpoons 2SO_3$ for more ($\uparrow$) forward SO_3 what can we do $P\downarrow$ $R\uparrow$

a) Add SO_3
b) Add SO_2 only & remove O_2
c) ~~Remove SO_3~~
d) ~~Add O_2 & SO_2~~

② Effect of pressure.

$A + 2B \rightleftharpoons 2C$

Δn_g = -ve ⟶ Reactant > Product
High pressure is good for forward

Δn_g = +ve ⟶ low pressure is good for forward

Q. $N_2 + 3H_2 \rightleftharpoons 2NH_3$ for making less ($\downarrow$) backward NH_3 which condition are good

a) High Pressure
b) ~~Low Pressure~~

Δn_g = (-ve) — backward
↓
Low pressure.

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③ Temprature

ΔH → +ve (Endothermic)
ΔH → -ve (Exothermic)

ΔH = +ve ⟶ High temp. for forward
ΔH = -ve ⟶ Low temp for forward.

Eg:- $N_2 + 3H_2 \rightleftharpoons 2NH_3$; $\Delta H = -46.8$ KJ/mol
for making more NH_3 which condition are good.

~~a) HP + LT~~
b) HP + HT
c) LP + HT
d) LP + LT

Δng = (-ve) High pressure
ΔH = (-ve) Low temp.

④ Effect of Catalyst.

Catalyst :- Only effect activation energy so equilibrium can't change.

⑤ Effect of Mixing inert gas.

Inert gas only decrease vol. of Container So equilibrium Const. not change.

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Questions.

1. For the following reaction, Equilibrium Constant 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 [JEE 2019]

$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$ is:

a) 10^{181} b) 10^{154}

c) 10^{25} d) 10^{77}

Solve.

$K_1 = \frac{[SO_2]}{[S][O_2]}$

$K_2 = \frac{[SO_3]^2}{[S]^2[O_2]^3}$

$K_{new} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

$= K_2 \times \frac{1}{(K_1)^2}$

$= \frac{[SO_3]^2}{[S]^2[O_2]^3} \times \frac{[S]^2[O_2]^2}{[SO_2]^2}$

$= 10^{129} \times \frac{1}{(10^{52})^2}$

$= \frac{10^{129}}{10^{104}}$

$K_{new} = 10^{25}$

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Que. In a chemical reaction, $A + 2B \longrightarrow 2C + D$, the initial concentration of B was 1.5 times of the concentration of A, but the equilibrium concentration of A and B were found to be equal. The equilibrium constant (K) for the aforesaid chemical reaction is:

[JEE 2021]

a) 16 ~~b) 4~~

c) 1 d) 1/4

Sol^n

Let $A_{initial} = A_0$

So, $B_{initial} = 1.5\,A_0$

A at t = B at t

$a_0 - x = 1.5a_0 - 2x$

$x = 0.5a_0$

	A +	2B ⟶	2C +	D
at t=0	a_0	$1.5a_0$	0	0
at t=t	$a_0 - x$	$1.5a_0 - 2x$	$2x$	x
	$a_0 - 0.5a_0$	$1.5a_0 - 2\times0.5a_0$	$2\times0.5a_0$	$0.5a_0$
	$0.5a_0$	$0.5a_0$	a_0	$0.5a_0$

$$K_C = \frac{[C]^2[D]}{[A][B]^2} = \frac{[a_0]^2[0.5a_0]}{0.5a_0\,(0.5a_0)^2} = \frac{1a_0}{0.5\times0.5_1} = 4$$

Que If the equilibrium constant for $A \rightleftharpoons B+C$ is $K_{eq}^{①}$ and that of $B+C \rightleftharpoons P$ is $K_{eq}^{②}$, the equilibrium constant for $A \rightleftharpoons P$ is [JEE 2020]

a) K_{eq}^{1} / K_{eq}^{2}
b) $K_{eq}^{1} + K_{eq}^{2}$
c) $K_{eq}^{2} - K_{eq}^{1}$
d) $K_{eq}^{1} K_{eq}^{2}$

$$K_{eq}^{1} = \frac{[B][C]}{[A]}$$

$$K_{eq}^{2} = \frac{[P]}{[B][C]}$$

$$K_{new} = \frac{[P]}{[A]}$$

$\Rightarrow K_{eq}^{1} \times K_{eq}^{2}$

$\Rightarrow \frac{[B][C]}{[A]} \times \frac{[P]}{[B][C]} = \frac{[P]}{[A]}$

Que for the reaction

$$Fe_2N_{(s)} + \frac{3}{2} H_{2(g)} \rightleftharpoons 2Fe + NH_{3(g)}$$

[JEE 2020]

a) $K_c = K_p (RT)$
b) $K_c = K_p (RT)^{3/2}$
c) $K_c = K_p (RT)^{1/2}$
d) $K_c = K_p (RT)^{-1/2}$

$\Delta n_g = 1 - \frac{3}{2}$
$= -1/2$

$K_p = K_c (RT)^{\Delta n_g}$

$K_p = K_c (RT)^{-1/2}$

$K_p = \frac{K_c}{(RT)^{1/2}}$

$K_c = K_p (RT)^{1/2}$

Que for a reaction at equilibrium $A_{(g)} \rightleftharpoons B_{(g)} + \frac{1}{2} C_{(g)}$ the relation b/w dissociation Constant (k), degree of dissociation (α) and equilibrium pressure (P) is given by. [JEE 2022]

a) $k = \dfrac{\alpha^{1/2} \; P^{3/2}}{\left(1 + \frac{3}{2}\alpha\right)^{1/2} \; (1-\alpha)}$

b) $k = \dfrac{\alpha^{3/2} \; P^{1/2}}{(2+\alpha)^{1/2} \; (1-\alpha)}$

c) $k = \dfrac{(\alpha P)^{3/2}}{\left(1 + \frac{3}{2}\alpha\right)^{1/2} \; (1-\alpha)}$

d) $k = \dfrac{(\alpha P)^{3/2}}{(1+\alpha) \; (1-\alpha)^{1/2}}$

Sol

$$A \rightleftharpoons B + \frac{1}{2} C$$

at $t=0$	1	-	-
$t=t$	$1-\alpha$	α	$\alpha/2$
P	$\frac{1-\alpha}{2+\alpha/2} \cdot P$	$\frac{\alpha}{2+\alpha/2} P$	$\frac{\alpha/2}{2+\alpha/2} \cdot P$

$1 - \alpha + \alpha + \alpha/2 = 2 + \alpha/2$

$$K_p = \frac{(P_B)\,(P_C)^{1/2}}{(P_A)} = \frac{\left(\frac{\alpha}{2+\alpha/2} P\right)\left(\frac{\alpha/2}{2+\alpha/2} \cdot P\right)^{1/2}}{\left(\frac{1-\alpha}{2+\alpha/2} P\right)}$$

$$= \frac{\alpha^{3/2} \; P^{1/2}}{(2+\alpha)^{1/2} \; (1-\alpha)}$$

Q The reaction quotient (Q) for the reaction,
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is given by

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

The reaction will proceed towards right if [JEE 2003]

a) $Q > K_c$ b) $Q = Q$
c) $Q = K_c$ ~~d)~~ $Q < K_c$

Q The value of ΔH for the reaction, $X_2(g) + 2Y_2(g) \rightleftharpoons 2XY_4(g)$ is less than zero. Formation of $XY_4(g)$ will be favoured at [JEE 2011]

a) LP & LT
b) HT & LP
~~c)~~ HP & LT
d) HT & HP

ΔH = (-ve)
Low Temp
Δn_g = 2-5
= (-ve)
High P
LT & HP

Q For the reversible reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + H$ the equilibrium shift in forward direction [JEE 2014]

a) by increasing the concentration of $NH_3(g)$
b) by decreasing the Pressure
c) by decreasing the concentrations of $N_2(g)$ and $H_2(g)$
~~d)~~ by increasing pressure and decreasing temperature

ΔH = -ve LT
Δn_g = 2-4 = -2
HP

22 Which one of the following conditions will favour maximum formation of the product in the reaction,

$A_2(g) + B_2(s) \rightleftharpoons X_2(g)$; $\Delta_r H = -X$ kJ ?

[NCERT 2018]

a) HT & HP
b) LT & LP
c) LT & HP
d) HT & LP

ΔH = -ve → LT

$\Delta n_g = 1 - 2 = -1$ → HP

14 In the two gaseous reaction (i) and (ii) at 250°C

(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 constant K_1 and K_2 are related as

[2005, 1.79 4]

a) $K_2 = \frac{1}{K_1}$

b) $K_2 = K^{1/2}$

c) $K_2 = \frac{1}{K_1^2}$

d) $K_2 = K_1^2$

$K_1 = \frac{[NO_2]}{[NO][O_2]^{1/2}}$

$K_2 = \frac{[NO]^2[O_2]}{[NO_2]^2}$

$K_2 \times (K_1)^2 = 1$

$K_2 = \frac{1}{(K_1)^2}$

Q8 In which of the following equilibrium K_c and K_p are not equal?

a) $2NO_{(g)} \rightleftharpoons N_{2(g)} + O_{2(g)}$

b) $SO_{2(g)} + NO_{2(g)} \rightleftharpoons SO_{3(g)} + NO_{(g)}$

c) $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

d) $2C_{(g)} + O_{2(g)} \rightleftharpoons 2CO_{2(g)}$

Q For the reaction,

$N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$, the equilibrium constant is K_1.

The equilibrium constant is K_2 for the reaction,

$2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$. What is K for the reaction

$NO_{2(g)} \rightleftharpoons \frac{1}{2} N_{2(g)} + O_{2(g)}$?

a) $1/4\,(K_1K_2)$

b) $[1/K_1K_2]^{1/2}$

c) $1/(K_1K_2)$

d) $1/(2K_1K_2)$

Solve

$$K_1 = \frac{[NO]^2}{[N_2][O_2]}$$

$$K_2 = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_{new} = \frac{[N_2]^{1/2}[O_2]}{[NO_2]}$$

$$\frac{1}{\sqrt{K_1 \times K_2}} = \frac{[N_2]^{1/2}[O_2]^{1/2}}{[NO]} \times \frac{[NO][O_2]^{1/2}}{[NO_2]}$$

$$\frac{1}{(K_1K_2)^{1/2}} = \frac{[N_2]^{1/2}[O_2]}{[NO_2]}$$

Q 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, [NEET 2015]

a) $K^{1/2}$ b) $\frac{1}{2}K$ c) K d) K^2

Solve $K = \frac{[NO]^2}{[N_2][O_2]}$ $K_{new} = \frac{[NO]}{[N_2]^{1/2}[O_2]^{1/2}}$

$\frac{K_{new}}{\sqrt{K}} = 1 \Rightarrow K_{new} = \sqrt{K} = K^{1/2}$

Q: A 20 lit. Container at 400 K Contains CO_2 (g) at pressure atm and an excess of SrO (neglect the vol of the vol of solid) The volume of the contained is now decreased by moving movable piston fitted in the Container. The maximum volume the Container, when pressure of CO_2 attains its maximum value, will be [NEET 2015]

(Given that: $SrCO_3(s) \rightleftharpoons SrO(s) + CO_2(g)$, $K_p = 1.6$ atm)

a) 5 L b) 10 L c) 4L d) 2L

Solve $P_1 = 0.4$ atm
$P_2 = 1.6$ atm
$V_1 = 20$ lit
$V_2 = ?$
$T_1 = 400$ K
$T_2 = 400$ K

$P_1V_1 = P_2V_2$
$0.4 \times 20 = 1.6\ V_2$
$V_2 = 5L$

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Q. $3O_{2(g)} \rightleftharpoons 2O_{3(g)}$

For the above reaction at 298 K, K_c is found to be 3.0×10^{-59}. If the concentration of O_2 at equilibrium is 0.040 M then Concentration of O_3 in M is.

[NEET 2022]

a) 1.9×10^{-63} b) 2.4×10^{31}

c) 1.2×10^{21} ✓ d) 4.38×10^{-32}

Sol.

$$K_c = \frac{[O_3]^2}{[O_2]^3}$$

$$3 \times 10^{-59} = \frac{[O_3]^2}{(0.040)^3}$$

$$[O_3]^2 = 3 \times 10^{-59} \times 4 \times 4 \times 10 \times 10^{-6}$$

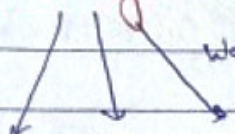
$$= 64 \times 3 \times 10^{-65}$$

$$O_3 = \sqrt{64 \times 3 \times 10^{-65}}$$

$$O_3 = 4.38 \times 10^{-32}$$

Ionic Equilibrium.

Electrolytes ⟶ ions breaks



An electrolytes is a substance that dissociates in water into charged particles called ions.

$HA + H_2O \longrightarrow HA_{(aq)}$

$BOH + H_2O \longrightarrow BOH_{(aq)}$

$AB + H_2O \longrightarrow AB_{(aq)}$

Eg:- $HCl_{(g)} + H_2O_{(aq)} \longrightarrow HCl_{(aq)}$

$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Eg:- $NaOH + H_2O \longrightarrow NaOH_{(aq)}$

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

Eg:- $NaCl + H_2O \longrightarrow NaCl_{(aq)}$

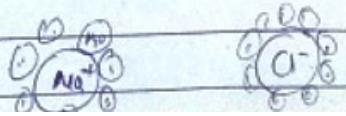
$NaCl_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

Hydration Energy (Hydration enthalpy)

The amount of energy released when one mole of ions undergo hydration.

$NaCl + H_2O \longrightarrow NaCl_{(aq)}$

$NaCl \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$



Acid/Base

① Taste — Acid ⟶ Sour in test
Base ⟶ Bitter in test

② Arrhenius — H^+ gives ⟶ Acid
Marden Arrhenius concept — OH^- gives ⟶ Base

According to Arrhenius a species which can gives H^+ called Acid. And marden Arrhenius concept a species which can gives OH^- Called Base.

Eg:- HCl, HBr, H_2SO_4 etc. are acid
NaOH, $Ca(OH)_2$, $In(OH)_2$ etc. are Base.

③ Lewis — l.p. gives ⟶ Base — NH_3, H_2O
l.p gives ⟶ Acid — BF_3, $AlCl_3$

④ Bronsted lowry Concept

→ H^+ give ⟶ Acid $\xrightarrow{\text{after giving } H^+}$ Conjugate base

→ H^+ take ⟶ Base $\xrightarrow{\text{after giving } OH^-}$ Conjugate Acid

H^+ gives

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

Acid — Base — Conj Acid — Conj Base

H^+ take.

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NOTE

Same { H^+ ⟶ Protenium
H_3O^+ ⟶ Hydronium

$H^+ + H_2O \longrightarrow H_3O^+$

Q1. $HX + B \rightleftharpoons X^- + HB^+$

Acid — Base — Conj. Base — Conj. Acid

H^+ donate (HX → X^-)

H^+ take (HB^+ → B)

Q2. $H_2SO_4^- + NH_3 \rightleftharpoons NH_4^+ + SO_4^{-2}$

Acid — Base — Conj. Acid — Conj. Base

H^+ take (NH_3 → NH_4^+)

H^+ give ($H_2SO_4^-$ → SO_4^{-2})

Q3. $[Fe(H_2O)_6]^{+3} + H_2O \rightleftharpoons H_3O^+ + [Fe(H_2O)_5OH]^{+2}$

Acid — Base — Conj Acid — Conj. Base

H^+ take (H_2O → H_3O^+)

H^+ give ($[Fe(H_2O)_6]^{+3}$ → $[Fe(H_2O)_5OH]^{+2}$)

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Q4. $HCl + NH_3 \rightleftharpoons NH_4^+ + Cl^-$

H+ Gives.

Acid — Base — Conj. Acid — Conj. Base

H+ take

Q5. $H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-$

H+ take

Acid — Base — Conj. Acid — Conj. Base

H+ gives

Q6. $HSO_4^- + R\text{-}NH_2 \rightleftharpoons R\text{-}NH_3^+ + SO_4^-$

H+ takes

Acid — Base — Conj. Acid — Conj. Base

H+ donte

Q7. $CH_3COOH + NH_3 \rightleftharpoons NH_2^+ + CH_3COO^-$

H+ donte

Acid — Base — Conj. Acid — Conj. Base

H+ takes

Q8. $NH_4^+ + S^{-2} \rightleftharpoons SH^- + NH_3$

H+ takes

Acid — Base — Conj. Acid — Conj. Base

H+ donte

Strength of Acid & Base → The extent to which it ionizes when dissolved water.

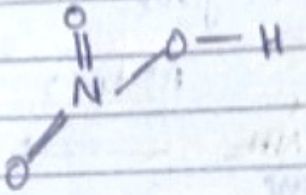
① $1H^+$ give ⟶ Monoacidic

oxy acid

Oxygen H ⟶ OH ⟶ $-O^-$

Mono basic acid

Eg:- HNO_3 ⟶

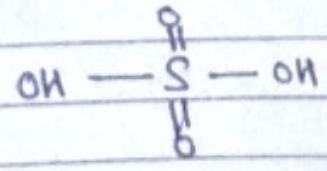


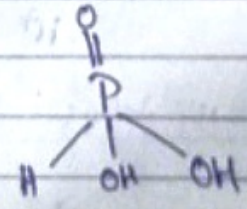
mono basic Acid — H, — OH

② Di-basic Acid

Eg:- H_2SO_4

H_3PO_3

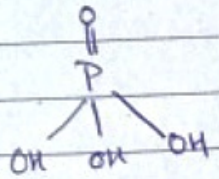




4 ⟶ 2H, 2—OH

③ Tri - basic acid

H_3PO_4



3 — O — H

3 — H give

① Mono basic ⟶ NaOH
② Di - basic ⟶ $Ca(OH)_2$
③ Tri - basic ⟶ $Ti(OH)_2$

$H_3PO_3 \longrightarrow 2H^+$

$H_3PO_4 \longrightarrow 3H^+$

Acid
- Strong. H_2SO_4, HCl, HNO_3, $HClO_4$, HBr, HI
- Week HCN, CH_3COOH, HCOOH, H_3PO_3, H_3PO_2, H_2Cl_3, $B(OH)_3$

Ionization
above 90%.
less than 5%.

Base
- Strong KOH, NaOH, $Ba(OH)_2$, CsOH, RbOH
- Week NH_4OH, $Zn(OH)_2$, $Ca(OH)_2$, $Fe(OH)_3$, $Al(OH)_3$

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New factor.

- log of Concentration

$[H^+] = 2 \times 10^{-2}$

$= 2 \times 10^{-1}$ $\quad 10^{-4}$

10^{-2} $\quad 10^{-5}$

10^{-3} $\quad 10^{-6}$

$-\log = P$

$= -\log_{10}^{-2}$

$= -(-2) \log_{10} 10 \rightarrow 1$

$= 2$

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

$pOH = -\log [OH^-]$

$pK_w = -\log [K_w]$

$pKa = -\log Ka$

$pKb = -\log Kb$

$HCl_g + H_2O_{(aq)} \rightleftharpoons H^+ + Cl^-$

$Ka = \frac{[H^+][Cl^-]}{[HCl]}$

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

$K_b = \frac{[Na^+][OH^-]}{[NaOH]}$

Water.

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

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

taking -log both side

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

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

$\therefore \log x \cdot y = \log x + \log y$

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

at 293K/25°C/NTP

$[H^+] = [OH^-] = 10^{-7}$

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

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

$\boxed{K_w = 10^{-14}}$

$pK_w = -\log 10^{-14}$

$\boxed{pK_w = 14}$

at 90°C/363K/

$[H^+] = [OH^-] = 10^{-6}$

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

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

$K_w = 10^{-12}$

$\boxed{pK_w = 12}$

pH of Water at 293 K

$[H^+] = [OH^-] = 10^{-7}$

$pH = -\log [H^+] = 10^{-7}$

$= -\log 10^{-7}$

$\boxed{pH = 7}$

pH at 373K & 90°C

$pH = -\log [H^+] = -\log 10^{-6}$

$\boxed{pH = 6}$

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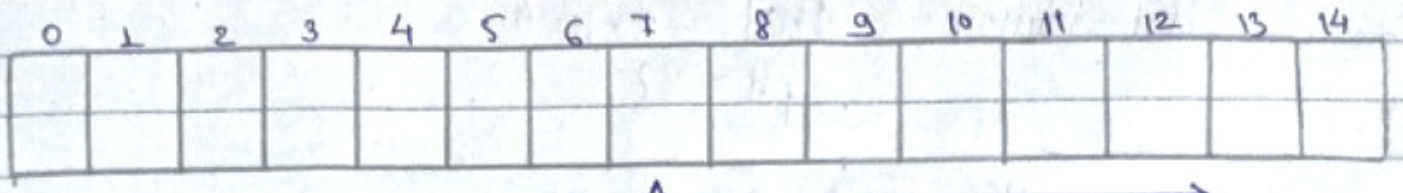
pH- meter



Increasing Acidic
OR
Increasing H^+ ions Concentration
$[H^+] > [OH^-]$

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

Increasing Base
OR
Increasing OH^- ions Concentration
$[H^+] < [OH^-]$

NOTE:
more pH → more alkline Soln.
less pH → more acidic Soln.

* Strong acid & Base are highly Corrosive

$$pH + pOH = 14$$

Neutral:

$pH = 7 \longrightarrow [H^+] = 10^{-7}$

$pH + pOH = 14$

$pOH = 7 \quad [OH^-] = 10^{-7}$

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

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Q. HCl Solution $[H^+] = 2 \times 10^{-4}$

$pH = ??$

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

	HCl	H^+	Cl^-
at $t=0$	C	0	0
$t=t$	$C(1-\alpha)$	$C\alpha$	$C\alpha$

$pH = -\log [H^+] \longrightarrow = -\log 2 \times 10^{-4}$

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

$= -\log 2 + (-\log 10^{4})$

$= -.3010 + 4$

$= 4 - 0.3010$

$\boxed{pH = 3.7}$

$NaOH = 10^{-4}$

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

$pOH = -\log 10^{4}$

$= 4$

$pH = 14 - pOH$

$= 14 - 4$

$pH = 10$

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Q1. Find the pH of 1×10^{-4} M HCl Solution?

Solve

$$pH = -\log [H^+]$$
$$= -\log [N]$$
$$= -\log (1 \times 1 \times 10^{-4})$$
$$= -\log 10^{-4}$$
$$pH = +4$$

Q What is pH of 5×10^{-3} M NaOH Solution?

Solve

$$N = n \times M$$
$$= 1 \times 5 \times 10^{-3}$$
$$N = 5 \times 10^{-3}$$

$$pOH = -\log 5 \times 10^{-3}$$
$$= -0.69 + 3$$
$$pOH = 2.31$$
$$pH = 14 - pOH$$
$$pH = 14 - 2.31$$
$$pH = 11.69$$

Normality

M x n factor.

HCl	→ 1
H_2SO_4	→ 2
H_3PO_2	→ 1
H_3PO_3	→ 2
H_3PO_4	→ 3
HNO_3	→ 1
NaOH	→ 1
$Ca(OH)_2$	→ 2
$Al(OH)_2$	→ 3

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Q What is the pH of 3×10^{-4} M H_2SO_4?

Solution

$$N = n \times M$$
$$= 2 \times 3 \times 10^{-4}$$
$$N = 6 \times 10^{-4}$$

$$pH = -\log (6 \times 10^{-4})$$
$$= -0.777 + 4$$
$$pH = 3.22$$

Q What is the pH of 5×10^{-6} M $Ca(OH)_2$ soln?

$$N = n \times M$$
$$N = 2 \times 5 \times 10^{-6}$$
$$N = 10 \times 10^{-6}$$
$$N = 10^{-5}$$

$$pH = 14 - pOH$$
$$pH = 14 - 5$$
$$pH = 9$$

Q. What is the pH of 3×10^{-3} M of H_3PO_3 soln?

$$N = n \times M$$
$$= 2 \times 3 \times 10^{-3}$$
$$= 6 \times 10^{-3}$$

$$pH = -\log (6 \times 10^{-3})$$
$$= 0.777 + 3$$
$$pH = 2.22$$

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Q1. 0.1M HCl of 100 ml $\longrightarrow N_1V_1 = 1 \times 0.1 \times 100 = 10$

0.1M HNO_3 of 100 ml $\longrightarrow N_2V_2 = 1 \times 0.1 \times 100 = 10$

New soln pH = ??

$$N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2} = \frac{10 + 10}{200} = \frac{20}{200} = 0.1$$

$$pH = -\log(0.1) = -\log 10^{-1} = 1$$

$\boxed{pH = 1}$

Q2. 0.2M H_2SO_4 of 200 ml $\longrightarrow N_1V_1 = 2 \times 0.2 \times 200 = 80$

0.2M HCl of 200 ml $\longrightarrow N_2V_2 = 1 \times 0.2 \times 200 = 40$

100 ml water

$$N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2 + V_3} = \frac{80 + 40}{500} = \frac{120}{500} = 0.24\,N$$

$$pH = -\log(0.24)$$
$$= -\log(2^3 \times 3 \times 10^{-2})$$
$$= -\log 2^3 \times 2 + 2$$
$$= -(3 \times 0.3010 + 0.4771) + 2$$

$\boxed{pH = 0.619}$

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Q. What is the pH 0.0005 N $Ca(OH)_2$ Soln?

$$pH = -\log (0.0005)$$
$$= -\log (5 \times 10^{-4})$$
$$= -(-4) 69 + 4$$
$$= 3.31$$

$$pH = 14 - pOH$$
$$= 14 - 3.31$$
$$\boxed{pH = 10.69}$$

Mixing of Acid & Base.

Acid-1 + Acid-2

$$N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$$

↓

pH

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Q. What is the pH of 0.1 M, 100 ml HNO_3 & 0.2 M, 100 ml H_2SO_4 & 300 ml of water?

$N_1V_1 = 0.1 \times 1 \times 100 = 10$

$N_2V_2 = 0.2 \times 2 \times 100 = 40$

$$N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2 + V_3} = \frac{10+40}{500} = \frac{50}{500} = 0.1$$

$\boxed{pH = 1}$

Q. What is the pH of acidic Solution Containt 0.5 M 50 ml HNO_3 & 0.5 M, 50 ml, H_3PO_4, and 100 ml of H_2O

$N_1V_1 = 0.5 \times 1 \times 50 = 25$

$N_2V_2 = 0.5 \times 3 \times 50 = 75$

$$\frac{25 \times 75}{50+50+100} = \frac{100}{200} = 0.5$$

$pH = -\log(0.5)$

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

$= -0.69 + 1$

$pH = 0.31$

Q. 0.2M 100ml NaOH → $N_1V_1 = 1 \times 0.2 \times 100 = 20$
0.2M 100ml $Ca(OH)_2$ → $N_2V_2 = 2 \times 0.2 \times 100 = 40$
300ml of H_2O

$$\frac{20+40}{500} = \frac{3}{25} \rightarrow 0.12$$

$$pH = -\log(0.2) = -\log(12 \times 10^{-2})$$
$$= -.3010 + 2$$
$$= 0.70$$

$$pOH = 14 - pH$$
$$pOH = 14 - 1.93$$
$$pOH = 13.07$$

Q. 0.2M 100 HCl → $N_1V_1 = 1 \times 0.2 \times 100 = 20$
0.1M 100 NaOH → $N_2V_2 = 0.1 \times 1 \times 100 = 10$

$$\frac{20+10}{200} = \frac{30}{200} = 0.15$$

$$-\log(0.15)$$
$$-\log(15 \times 10^{-2})$$
$$-\log(3 \times 5 \times 10^{-2})$$
$$-\log 3 + \log 5 \times 10^{-2}$$
$$-0.477 + 0.69 + 2$$
$$2.21$$

Ostwald Dilution law — The application of the law of mass action to weak electrolytes in solution.

Only for week acid or week base

- Strong acid ⟶ ionization above 90% Eg:- HCl, H_2SO_4
- Weak acid. ⟶ ionization less the 5% Eg:- CH_3COOH.

Week acid.

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

	HA	H^+	A^-
at t=0	C	0	0
t=t	$C(1-\alpha)$	$C\alpha$	$C\alpha$

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

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

Now, we take Ostwald dilution law.

$1 \gg \alpha$

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

$$K_a = C\alpha^2$$

$$\frac{K_a}{C} = \alpha^2$$

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

$$[H^+] = C\cdot\alpha$$

$$= C\sqrt{\frac{K_a}{C}}$$

$$[H^+] = \sqrt{\frac{K_a C^2}{\alpha}}$$

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

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

$$pH = -\log[H^+] = -\log (K_a \cdot C)^{1/2}$$

$$= -\frac{1}{2}\log (K_a \cdot C)$$

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

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

$$\boxed{pH = -\frac{1}{2}(\log K_a + \log C)}$$

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For Weak base

$$BOH \rightleftharpoons B^+ + OH^-$$

	BOH	B^+	OH^-
at $t=0$	C	0	0
$t=t$	$C(1-\alpha)$	$C\alpha$	$C\alpha$

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

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

neglect

Ostwald dilution law

$1 >> \alpha$

$$K_b = C\alpha^2$$

$$\alpha^2 = \frac{K_b}{C}$$

$$\alpha = \sqrt{\frac{K_b}{C}}$$

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

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

$$POH = -\log \sqrt{K_b C}$$

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

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

$$pH = 14 - pOH$$

$$\boxed{pH = 14 + \frac{1}{2}(\log K_b + \log C)}$$

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Q. Calculate pH of 0.02 M CH_3COOH if $k_a = 5 \times 10^{-6}$

$[H^+] = C\alpha$

$pH = -\frac{1}{2}(\log k_a + \log C)$

$= -\frac{1}{2}(\log 5 \times 10^{-6} + \log 2 \times 10^{-2})$

$= -\frac{1}{2}(0.69 - 6 + 0.3010 \bar{2})$

$= -\frac{1}{2}(-5.31 - 170)$

$= 7 \times \frac{1}{2}$

$pH = 3.5$

Q. What is the pH of 0.05 M CH_3COOH Solution if $k_a = 2 \times 10^{-6}$ & also Calculate degree of ionization (α)

$\alpha = \sqrt{\frac{k_a}{C}} = \sqrt{\frac{2 \times 10^{-7}}{5 \times 10^{-2}}}$

$= \sqrt{4 \times 10^{-5}} = \sqrt{140 \times 10^{-6}}$

$= \sqrt{6.3 \times 10^{-3}}$

Salt

Salt are of 4 Types

① Strong Acid & Weak base type Salt
SA + WA

Eg:- $HCl + NH_4OH$ This type of Salt formed by a ~~weak~~ Strong acid & weak base, this type of Salt are acidic in nature.

Eg:- NH_4Cl, $NaHSO_4$ etc.

② Hydrolysis Constant :- This type of Constant only used in hydrolysis of Salt.
Denoted by (K_h)

② Strong Base & Weak acid of Salt.

Example :- $Zn(OH)Cl$, $Mg(OH)Cl$, $Be(OH)_2Cl$

- This type of Salt are basic in nature.
- The pH of that type of Salt are more than 7

③ Normal Salt :-

This type of Salt formed by mixing of ① weak acid & weak base ② Strong acid & Strong base.

WA + WB = Eg :- CH_3COONH_4

SA + SB = Eg :- $NaCl$

Degree of Hydrolysis. (h)

$$h = \sqrt{\frac{K_h}{C}}$$

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Relationship b/w K_a, K_b, K_h, K_w

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

① For SA + WB

$$K_h = \frac{K_w}{K_b}, \quad h = \sqrt{\frac{K_w}{K_b \cdot C}}$$

Similarly, WA + SB

$$K_h = \frac{K_w}{K_a}, \quad h = \sqrt{\frac{K_w}{K_a \cdot C}}$$

Similarly, WA + WB

$$K_h = \frac{K_w}{K_a * K_b}, \quad h = \sqrt{\frac{K_w}{K_a \cdot K_b \cdot C}}$$

SA + WB	WA + SB	WA + WB	SA + SB
pH = less than 7	pH = more than 7	equal to 7	Equal to 7
Acidic	Basic	Neutral	Neutral
$K_h = \frac{K_w}{K_b}$	$K_h = \frac{K_w}{K_a}$	$K_h = \frac{K_w}{K_a \cdot K_b}$	—
$h = \sqrt{\frac{K_w}{K_b \cdot C}}$	$h = \sqrt{\frac{K_w}{K_a \cdot C}}$	$h = \sqrt{\frac{K_w}{K_a K_b C}}$	—
$[H^+] = Ch = \sqrt{\frac{K_w \cdot C}{K_b}}$	$[H^+] = \sqrt{\frac{K_w \cdot C}{K_a}}$	$[H^+] = \sqrt{\frac{K_w \cdot C}{K_a K_b}}$	
$pH = 7 + \frac{1}{2}\log K_b - \frac{1}{2}\log C$	$pH = 7 - \frac{1}{2}\log K_a + \frac{1}{2}\log C$		$pH = 7 - \frac{1}{2}\log K_b +$ $\frac{1}{2}\log K_a$

Q. What is the pH of CH_3COONa if $K_a = 1.8 \times 10^{-5}$
Conc. = 1M

$$pH = 7 - \frac{1}{2}\log K_a + \frac{1}{2}\log C$$

$$= 7 - \frac{1}{2}\log(1.8\times10^{-5}) + \frac{1}{2}\log 1$$

$$= 7 + \frac{1}{2}\log 10^{-5} + 0$$

$$= 7 + 2.5$$

$pH \approx 9.5$

Q. What is the value of h if $K_a = 10^{-5}$, $K_b = 10^{-5}$

WA + WB:

$$h = \sqrt{\frac{K_w}{K_a \cdot K_b}} = \sqrt{\frac{10^{-14}}{10^{-5}\times10^{-5}}} = \sqrt{10^{-4}} = 10^{-2} = 1\%$$

Q. What is the pH of NH_4OH type of Salt whose $K_b = 5\times10^{-6}$ & Conc. are 0.01 M

$$pH = 7 + \frac{1}{2}\log K_b - \frac{1}{2}\log C$$

$$= 7 + \frac{1}{2}\log(5\times10^{-5}) - \frac{1}{2}\log(10^{-2})$$

$$= 7 + \frac{1}{2}\times 0.69 \times 5 + 1$$

$$= 7 + 2.15$$

$$= 8 - 2.15$$

$$= 5.85$$

8.00
0.17
.83

Q. The pKa of acetic acid and pK_b of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.

CH_3COONH_4

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

$$pH = 7 + \frac{1}{2}(4.7 - 4.75)$$

$$pH = 7 + \frac{1}{2} \times 0.01$$ 0.005

$$= 7 + 0.005$$

$$pH = 7.005$$

Buffer Solution.

→ A Solution which rajist the change in pH on dilution or adding Small amount of acid and base.

* Properties of Buffer Solution

① The pH of Buffer Solution does not change upto the Small amount of Strong acid and Strong base.

② The pH of Buffer Solution does not depend upon the Volume of Solution, hence Solution can be diluted without change in pH.

③ The pH of Buffer Solution does not change remains their life.

* Type of Buffer Solution.

① Simple Buffer Solution :-

Example :- It is formed by mixing of weak acid & weak base like CH_3COONH_4, NH_4CN

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② Mixed Buffer Solution:-

① Acidic Buffer Solution

Acidic Buffer Solution formed by mixing of weak acid + weak acid Strong base Salt

Eg:- $CH_3COOH + CH_3COONa$

↑ pH are arround 4.75

Handoresson Hassel batch equation

$$pH = pka + \log \frac{[salt]}{[Acid]}$$

⑪ Basic Buffer Solution.

↑ is formed by mixing of weak base + Strong acid weak base Salt

Eg:- $NH_4OH + NH_4Cl$

↑ pH are arround 9.25

Mechanism ⟶ Le-Chateljer's

When we mix acid like HCl

Acidic Buffer Solution

$CH_3COOH + CH_3COONa$

① $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ pH=4.75

② $CH_3COONa \longrightarrow CH_3COO^- + Na^+$

③ $HCl \longrightarrow H^+ + Cl^-$

When we mix NaOH

① $CH_3COOH \longrightarrow CH_3COO^- + H^+$

⑪ $CH_3COONa \longrightarrow CH_3COO^- + Na$

⑪ $NaOH \rightleftharpoons Na^+ + OH^-$

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Solubility Product.
(Sparingly Soluble Salt)

AB type of Salt

Eg. NaCl

	$AB \rightleftharpoons$	A^+	+ B^-
at t=0	S	0	0
t=t	0	S	S

$K_{sp} = S.S = S^2$

Eg:- Given $K_{sp} = 8 \times 10^{-4}$
Solubility NaCl

	$NaCl \longrightarrow$	Na^+	+ Cl^-
at t=0	S	0	0
t=t	0	S	S

$K_{sp} = S^2$

$S^2 = 8 \times 10^{-4}$

$S = \sqrt{8 \times 10^{-4}}$

$S = 2\sqrt{2} \times 10^{-2}$

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A_2B / AB_2 type of Salt Eg: $CaCl_2$

	AB	$\rightleftharpoons$	A^{+2}	+	$2B^-$
at $t=0$	S		0		0
at $t=t$	0		S		S

$K_{sp} = S.(2S)^2$

$K_{sp} = 4S^3$

AB_3 type of Salt Eg: $AlCl_3$

	AB_3	$\rightleftharpoons$	A^{+3}	+	$3B^-$
at $t=0$	S		0		0
$t=t$	0		S		S

$K_{sp} = S.(2S)^3$

$= S.27.S^3$

$K_{sp} = 27S^4$

A_2B_3 type of Salt Al_2O_3

	A_2B_3	$\rightleftharpoons$	$2A^{+3}$	+	$3B^{-2}$
at $t=0$	S		0		0
at $t=t$	0		$2S$		$3S$

$K_{sp} = (2S)^2 (3S)^3$

$= 4S^2 \times 27S^3$

$K_{sp} = 108\, S^5$

A_xB_y = K_{sp} = ?

$A_xB_y \longrightarrow xA^{+y} + yB^{-x}$

S	0	0
0	S	S

$K_{sp} = (xS)^x (yx)^y$

$K_{sp} = x^x y^y S^{x+y}$

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Q. AB type Solution , $S = 5 \times 10^{-2}$, $K_{sp} = ?$

$$AB \rightleftharpoons A^+ + B^-$$
$$S \quad S$$

$$K_{sp} = S^2 = (5 \times 10^{-2})^2$$

$$\boxed{K_{sp} = 25 \times 10^{-4}}$$

Common ion effect.

Salt = ? → An effect that suppresses the ionization of an Electrolyte when another electrolyte (which contains an ion which is also present in the first electrolyte, i.e. a common ion) is added.

→ It is considerd to be a consequence of le Chatlier's principal (or the Equilibrium Law)

Make Salt.

Mix → Sand
↓
$CaCl_2$
$MgCl_2$
KCl

Conc. Solution
↓
Saturated Solution

Chemical
Saturated Sol^n →
↓
Precipitate