PARISHRAM

2026

CHEMISTRY

Bharat Mata ki Jai ♡

Lecture 01

CHEMICAL KINETICS

Introduction to Chemical Kinetics and Rate of Reaction

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Introduction to Chemical Kinetics ✓

(ii) Rate of Reaction ✓

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The word **kinetics** is derived from the Greek word '***kinesis***' meaning **movement**.

INTRODUCTION TO CHEMICAL KINETICS

INTRODUCTION TO CHEMICAL KINETICS

The different branches of Chemistry tells us different things about chemical reactions which are:

(i) **Thermodynamics** tells us **whether a chemical reaction is possible** (feasible/spontaneous) under given conditions.

(ii) **Chemical Equilibrium** tells us **how much reactant has been left or product has been formed**— when the reaction reaches equilibrium.

(iii) **Chemical Kinetics** tells us **how fast the reaction occurs** and the steps (mechanism) involved.

NEED OF CHEMICAL KINETICS

$$N_2 + 3H_2 \leftrightarrow 2NH_3$$

→ at room temp. it is a very slow rxn

Branch of Chemistry	What It Tells About This Reaction	Important Point
Thermodynamics	Whether the reaction is feasible or not	$\Delta G < 0$ → Reaction is spontaneous (can happen on its own)
Chemical Equilibrium	How much N_2 and H_2 will be left and how much NH_3 will be formed finally	
Chemical Kinetics	How fast the reaction will happen and how much time it will take	Reaction is very slow at room temperature; needs high T, P, catalyst

SAMAJ AAYA TOH
LIKH DO.

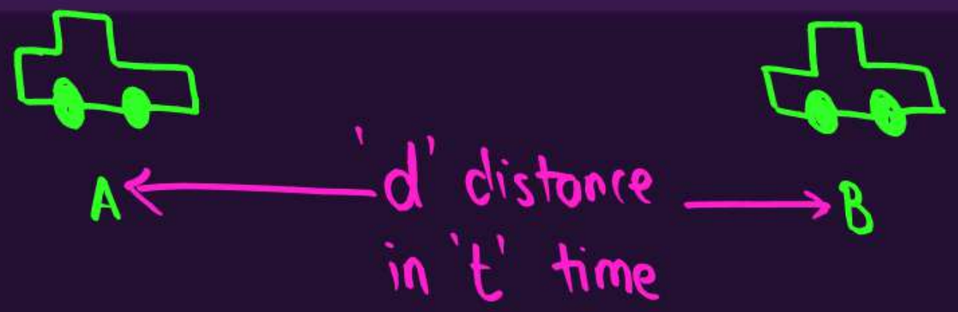
AYE BHAIYA

RATE OF REACTION

SPEED (RATE) OF REACTION

- Speed of a car is expressed in terms of 'change in the position' or 'distance covered by it in a certain period of time.'

- Similarly, speed or rate of reaction is the change in concentration of reactant(s)/product(s) in a certain period of time.



$$\boxed{\text{Speed of car} = \frac{d}{t}}$$

GIVE A THOUGHT

Can we measure the speed/rate of precipitation reactions like the one between $AgNO_3$ and NaCl?

A. YES

B. NO

Colourless Soltn

$$\underset{C}{Ag}\underset{A}{NO_3} + \underset{C}{Na}\underset{A}{Cl} \longrightarrow AgCl + NaNO_3$$

[White insoluble solid or white ppt.]

→ Double Displacement Rxn (Precipitation Rxn)

→ Also, called ionic rxn.

GIVE A THOUGHT

Can we measure the speed/rate of rusting of iron or conversion of diamond to graphite?

A. YES

B. NO

very slow rxn (takes even months or years to complete)

IMPORTANT TO REMEMBER

We can measure the rate of reactions that are neither too fast nor too slow — reactions that show noticeable change in a reasonable time (seconds to hours). → moderate rate of rxn

Example: Reaction between Mg and HCl, neutralization of weak acid with weak base etc.

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AVERAGE SPEED (AVERAGE RATE) OF REACTION

$\frac{\overset{100}{\cancel{500}}\ km}{\cancel{5}\ hr}$

$100\ \frac{km}{hr}$

- To calculate average speed, we know:

Speed OR

$$\text{Average Speed} = \frac{\text{Total Distance Travelled}}{\text{Total Time Taken}}$$

- Similarly, to calculate the average speed (average rate) of a reaction we calculate it using:

$$\text{Average Rate of Reaction} = \frac{\text{Total Change in Concentration of Reactant or Product}}{\text{Time Interval / Total time taken}}$$

OR

Rate of Rxn

AVERAGE SPEED (AVERAGE RATE) OF REACTION

$C = \frac{n}{V}$

V → constant

$C \propto n$

- Consider the below hypothetical reaction:

  R → P; means 1 mole R to 1 mole P

- Here assume that volume of the system remains constant so that [change in concentration is due to the reaction and not because of volume.]

change in concentration of R & P ∝ no. of moles of R & P

- [R]: molar concentration of reactant R (mol/L)
  [P]: molar concentration of product P (mol/L)

- At time t_1, concentrations = $[R]_1$, $[P]_1$
  At time t_2, concentrations = $[R]_2$, $[P]_2$

AVERAGE SPEED (AVERAGE RATE) OF REACTION

because conc. of reactant is (↓)

- Change in time (Δt): $= t_2 - t_1$
  Change in concentration of reactant (ΔR): $= [R]_2 - [R]_1 =$ (−ve value)
  Change in concentration of product (ΔP): $= [P]_2 - [P]_1 =$ (+ve value)

conc. of product (↑)

- Rate of disappearance of R = $-\dfrac{\text{Total Change in Concentration of Reactant}}{\text{Time Interval}}$

$$\Rightarrow -\frac{\Delta[R]}{\Delta t} \Rightarrow -\frac{[R]_2 - [R]_1}{t_2 - t_1}$$

- Rate of appearance of P = $\dfrac{\text{Total Change in Concentration of Product}}{\text{Time Interval}}$

$$\Rightarrow \frac{\Delta[P]}{\Delta t} = \frac{[P]_2 - [P]_1}{t_2 - t_1}$$

GIVE A THOUGHT

Why we put a –ve sign while writing the expression of rate of disappearance of a reactant in a reaction?

Since, ΔR is a negative quantity (as concentration of reactants is decreasing), it is multiplied with –1 to make the rate of the reaction a positive quantity.

IMPORTANT TO REMEMBER

$\frac{\text{change in conc.}}{\text{time}} = \frac{mol \cdot L^{-1}}{s}$

$= mol \cdot L^{-1} \cdot s^{-1}$

UNITS OF RATE OF REACTION

- If concentration is in **$mol \cdot L^{-1}$** and time is in seconds then the units will be **$mol \cdot L^{-1} \cdot s^{-1}$**.
- In gaseous reactions, when the concentration of gases is expressed in terms of their partial pressures, then the units of the rate equation will be **$atm \cdot s^{-1}$**.

PYQS' WALLAH

CBSE PYQ

For the hydrolysis of methyl acetate in aqueous solution, the following results were obtained:

t/s	0	10	20
$[CH_3COOCH_3]$/mol L^{-1}	0.10	0.05	0.025

(b) Calculate the average rate of reaction between the time interval 10 to 20 seconds.

$$\text{Avg. rate of rxn} = -\frac{\text{Total change in concentration } (mol \cdot L^{-1})}{\text{Total time interval } (s)}$$

$$= -\frac{(.025 - .05)}{(20 - 10)} = -\frac{(-.025)}{10} = \frac{.025}{10} = .025 \times 10^{-1}$$

$$= 2.5 \times 10^{-2} \times 10^{-1}$$

$$= 2.5 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1}$$

CBSE PYQ

For the hydrolysis of methyl acetate in aqueous solution, the following result were obtained.

t/s	0	30	60
$[CH_3COOCH_3]$/mol L^{-1}	0.60	0.30	0.15

(ii) Calculate the average rate of reaction between the time interval 30 to 60 seconds.

$$\text{avg. rate of rxn} = -\frac{\text{total change in concentration}}{\text{total time interval}}$$

$$= -\frac{(.15 - .30)}{60 - 30}$$

$$= -\frac{(-.15)}{30}$$

$$= \frac{.15}{30 \times 100} = \frac{1}{2 \times 100}$$

$$= .5 \times 10^{-2}$$

$$= 5.0 \times 10^{-1} \times 10^{-2}$$

$$= \boxed{5 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1}}$$

AVERAGE SPEED (AVERAGE RATE) OF REACTION

For the reaction:

(Stoichiometric Coefficient)

$$aA + bB \rightarrow cC + dD$$

C-I

Rate of disappearance of A $= -\frac{\Delta[A]}{\Delta t}$

Average rate is written as:

C-II

$$\text{Rate of rxn} = -\frac{1}{a}\cdot\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\cdot\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\cdot\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\cdot\frac{\Delta[D]}{\Delta t}$$

LET'S PRACTICE

For the following reaction: $2A + 3B \longrightarrow 4C$

Write the expression for the **average rate of reaction** in terms of the:

(i) Rate of disappearance of A and B

(ii) Rate of appearance of C

Also, explain why stoichiometric coefficients are used in the rate expression.

Rate of disappearance of A $= -\frac{\Delta[A]}{\Delta t}$

-"- of B $= -\frac{\Delta[B]}{\Delta t}$

-"- appearance of C $= +\frac{\Delta[C]}{\Delta t}$

Rate of rxn $= -\frac{1}{2}\frac{\Delta[A]}{\Delta t} = -\frac{1}{3}\frac{\Delta[B]}{\Delta t} = +\frac{1}{4}\cdot\frac{\Delta[C]}{\Delta t}$

LET'S PRACTICE

For the following reaction: $2A + 3B \longrightarrow 4C$

Write the expression for the **average rate of reaction** in terms of the:

(i) Rate of disappearance of A and B

(ii) Rate of appearance of C

Also, explain why stoichiometric coefficients are used in the rate expression.

We divide by stoichiometric coefficients to make the rate same no matter which reactant or product we use in the expression.

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CONCEPT POLISH - HOMEWORK

LET'S PRACTICE

For the following reaction: $N_2 + 3H_2 \rightarrow 2NH_3$
During the formation of ammonia 2 moles of N_2 disappear in 1 minute.
Find: (i) Rate of reaction (ii) Rate of disappearance of H_2 (iii) Rate of appearance of NH_3.
The volume of vessel is 1 L.

H.W.

SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai

#pwsathhai

Thank You

PARISHRAM

2026

Chemistry

Bharat Mata Ki Jai

Lecture 02

CHEMICAL KINETICS

Rate of Reaction (Contd.) and Factors Affecting Rate of Reaction

BY – PRIYA-PUTRA-SUNIL

(Sunil Sir)

TOPICS TO BE COVERED

(i) Rate of Reaction (Contd.) ✓

(ii) Factors Affecting Rate of Reaction ✓

CONCEPT POLISH – HOMEWORK DISCUSSION

LET'S PRACTICE

For the following reaction, $N_2 + 3H_2 \rightarrow 2NH_3$
During the formation of ammonia 2 moles of N_2 disappear in 1 minute.
Find: (i) Rate of reaction (ii) Rate of disappearance of H_2 (iii) Rate of appearance of NH_3.
The volume of vessel is 1 L.

(i) Rate of rxn $= -\frac{1}{1}\frac{\Delta[N_2]}{\Delta t} = -\frac{1}{3}\frac{\Delta[H_2]}{\Delta t} = +\frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$

$= -\frac{\left(\frac{-2\ mol}{1L}\right)}{60\ s} = \frac{\cancel{2}^{1}}{\cancel{60}_{30}}\ mol \cdot L^{-1} \cdot s^{-1} = \frac{1}{30}\ mol \cdot L^{-1} \cdot s^{-1}$

(ii) Rate of disappearance of $H_2 = -\frac{\Delta[H_2]}{\Delta t}$

$$= -\frac{\left(\frac{-6\ mol}{1L}\right)}{60\ s}$$

$$= \frac{\cancel{6}^{1}}{\cancel{60}_{10}}\ mol \cdot L^{-1} \cdot s^{-1}$$

$$= \frac{1}{10}\ mol \cdot L^{-1} \cdot s^{-1}$$

1 mol N_2 → 3 mol H_2

2 mol N_2 → 'x'

$\boxed{x = 6\ mol}$

1 mol N_2 → 2 mol NH_3

2 mol N_2 → 'x'

$\boxed{x = 4\ mol}$

(iii) Rate of appearance of $NH_3 = +\frac{\Delta[NH_3]}{\Delta t} = \frac{\frac{4\ mol}{1L}}{60\ s} = \frac{4}{60}\ mol \cdot L^{-1} \cdot s^{-1}$

$$= \frac{1}{15}\ mol \cdot L^{-1} \cdot s^{-1}$$

RATE OF REACTION (CONTD.)

a.b.

INSTANTENOUS RATE OF Rxn

- **Instantaneous rate** of a chemical reaction is the rate at a **particular instant of time.**

OR

- It is the change in concentration of a reactant or a product during an infinitesimally small-time interval. $\Delta t \to 0$ $\Delta t = 0$ ×

Instantenous
Rate of disappearance of reactant = $-\frac{d[R]}{dt}$

Instantenous
rate of appearance of product = $+\frac{d[P]}{dt}$ (Instantenous Rate of Rxn)

ex: (i) $t_1 = 4\,s$ $t_2 = 16\,s$

$\Delta t = 16 - 4 = 12\,s$

Average Rate of Rxn

(ii) $t = 5\,s$ $t_1 = 4.99$ $t_2 = 5.01$

$\Delta t = t_2 - t_1 = dt$

INSTANTENOUS RATE

$$aA + bB \rightarrow cC + dD$$

Instantenous Rate of rxn $= -\frac{1}{a} \cdot \frac{d[A]}{dt} = -\frac{1}{b} \cdot \frac{d[B]}{dt} = +\frac{1}{c} \cdot \frac{d[C]}{dt} = +\frac{1}{d} \cdot \frac{d[D]}{dt}$

LET'S PRACTICE

For the reaction:

$$\underbrace{2A + B}_{\text{Reactants}} \rightarrow \underbrace{3C}_{\text{Product}}$$

Which of the following correctly represents the **instantaneous rate of reaction**?

(a) $\frac{d[C]}{dt}$ ✗

(b) $-\frac{d[A]}{dt}$ ✗

(c) $-\frac{1}{2} \cdot \frac{d[A]}{dt}$ ✓

(d) $\frac{1}{3} \cdot \frac{d[C]}{dt}$ ✓

(i) Instantenous R.o.R. $= -\frac{1}{2}\frac{d[A]}{dt} = -\frac{1}{1}\frac{d[B]}{dt} = +\frac{1}{3}\frac{d[C]}{dt}$

(ii) Instantenous Rate of appearance of C $= +\frac{d[C]}{dt}$

(iii) Instantenous Rate of disappearance of A $= -\frac{d[A]}{dt}$

LET'S PRACTICE

Given the reaction:

$$2NO_2 \rightarrow 2NO + O_2$$

If $\frac{d[O_2]}{dt} = 0.01 \text{ mol L}^{-1}\text{s}^{-1}$,
what is $\frac{d[NO_2]}{dt}$?

Instantenous Rate of Rxn $= -\frac{1}{2}\frac{d[NO_2]}{dt} = +\frac{1}{2}\frac{d[NO]}{dt} = +\frac{1}{1}\frac{d[O_2]}{dt}$

$\Rightarrow -\frac{1}{2}\frac{d[NO_2]}{dt} = .01$

$\Rightarrow \frac{d[NO_2]}{dt} = .02 \text{ mol} \cdot L^{-1} \cdot s^{-1}$

LET'S PRACTICE

For the gaseous reaction:

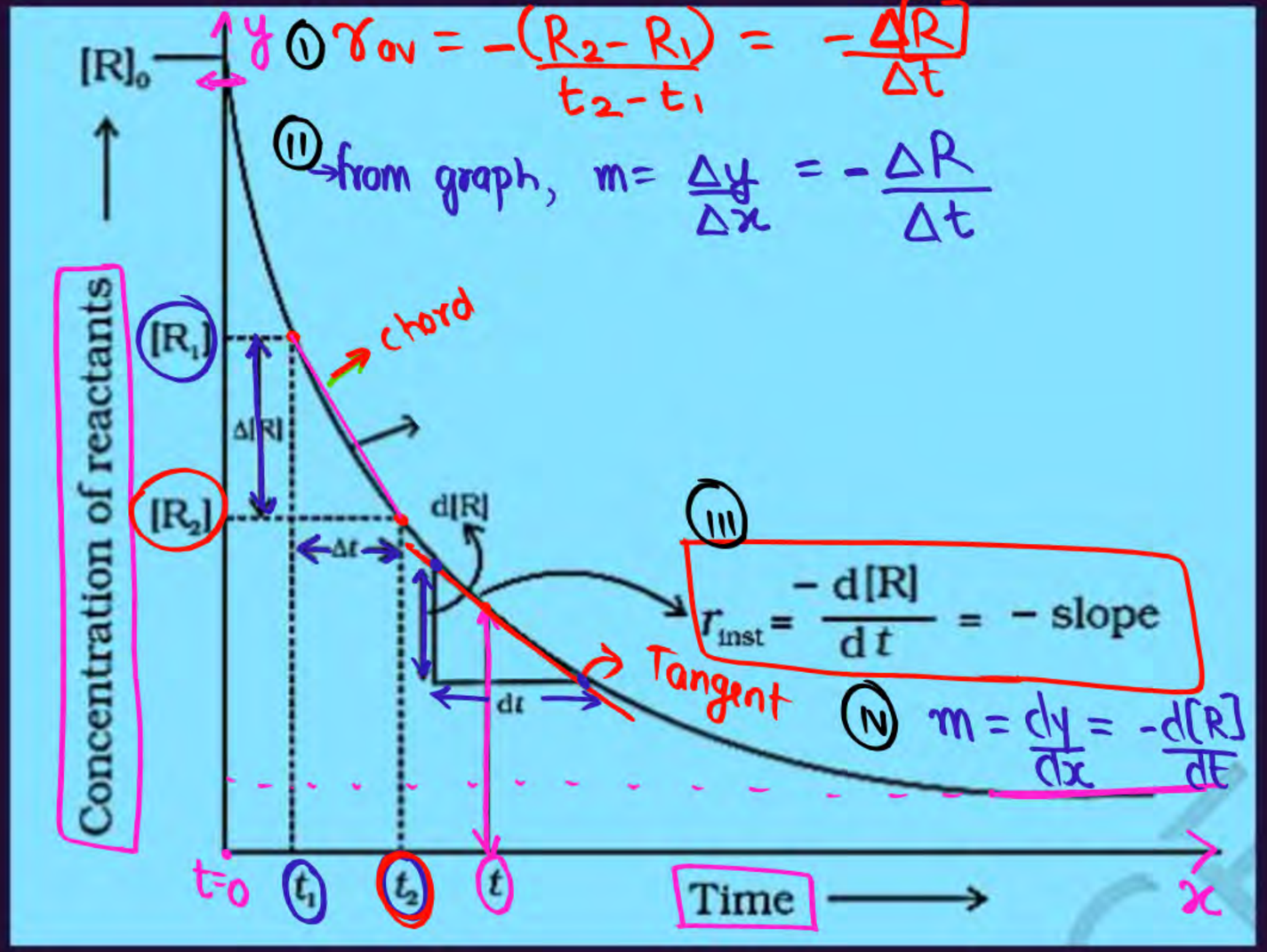
$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

If pressure of O_2 decreases from 1.2 atm to 1.0 atm in 10 seconds: Find Instantenous Rate of disappearance of O_2?

Instantaneous rate of disappearance of O_2 $= -\frac{d[O_2]}{dt} = -\left(\frac{1 - 1.2}{10}\right) = \frac{.2}{10} = .02\ atm \cdot s^{-1}$

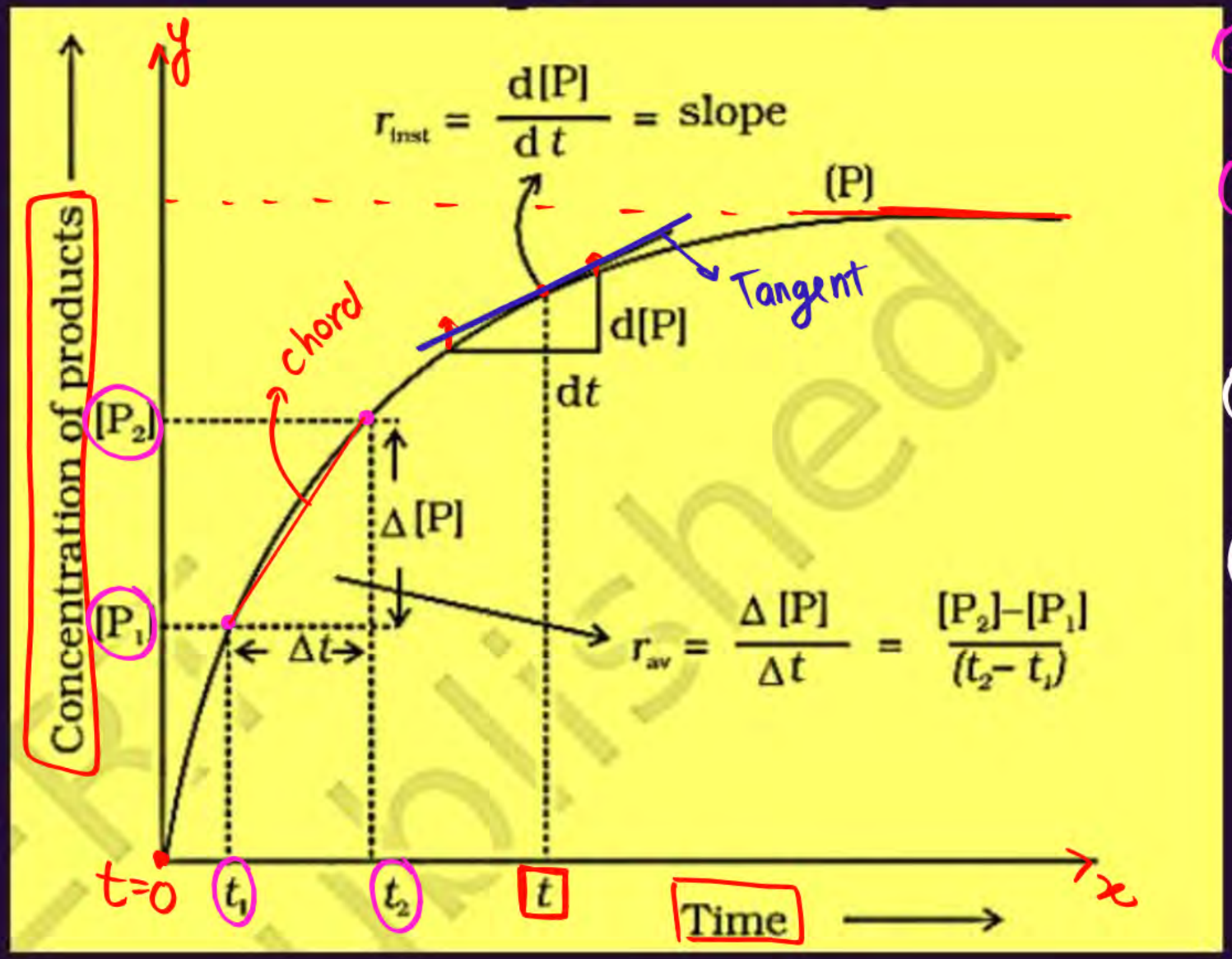
clear hai

LET'S DECODE NCERT



- Average rate is the **slope of the chord (secant line)** to the concentration vs. time graph over a **finite time interval**.
- Instantenous rate is the **slope of the tangent** to the concentration vs. time graph at a **particular instant**.

LET'S DECODE NCERT



(i) $r_{av} = \frac{P_2 - P_1}{t_2 - t_1} = +\frac{\Delta[P]}{\Delta t}$

(ii) graph, $m = \frac{\Delta y}{\Delta x} = +\frac{\Delta[P]}{\Delta t}$

(iii) $r_{ins} = \frac{d[P]}{dt}$

(iv) graph, $m = \frac{dy}{dx} = \frac{d[P]}{dt}$

LET'S PRACTICE

Assertion (A): Instantaneous rate is equal to the slope of the tangent to the concentration-time graph. (✓)

Reason (R): Instantaneous rate is calculated over a long interval of time (✗)

(A) A is true, R is false

(B) A is false, R is true

$\Delta t \to 0$

$\Delta t = t_2 - t_1 = dt$

DIFFERENCE BETWEEN AVERAGE AND INSTANTENOUS RATE

for this we consider a very small time interval (dt)

Feature	Average Rate	Instantaneous Rate
Time interval	Over a **finite** time interval Δt	At a **particular instant** of time t
Formula	$\text{Rate} = -\frac{\Delta[R]}{\Delta t} = \frac{\Delta[P]}{\Delta t}$	$\text{Rate} = -\frac{d[R]}{dt} = \frac{d[P]}{dt}$
Graphical meaning	Slope of the **chord** (secant line)	Slope of the **tangent** at a point
Accuracy	Less accurate	More accurate
Units	Same for both: **mol L^{-1} s^{-1}**	**mol L^{-1} s^{-1}**

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FACTORS AFFECTING RATE OF REACTION

FACTORS AFFECTING RATE OF REACTION

(i) Nature of Reactant: The rate of reaction is strongly influenced by the chemical nature of the reactant, i.e.

(a) Ionic or Covalent Compounds: Ionic compounds in aqueous solution react very fast because ions are readily available while covalent compounds react slowly as bonds need to be broken and reformed.

very fast rxn ← $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) \downarrow + NaNO_3(aq)$ → Ionic reactants

C A C A

slow rxn ← $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ → Covalent reactants

$(\ddot{N} \equiv \ddot{N})$ $(H-H)$ $\left[H-\ddot{N}(-H)-H\right]$

FACTORS AFFECTING RATE OF REACTION

(i) Nature of Reactant: The rate of reaction is strongly influenced by the chemical nature of the reactant, i.e.

(b) Physical State of Reactants and Products:

(b)-(i) Gases and liquids usually react faster than solids, as particles can collide more easily due to more kinetic energy.

(b)-(ii) Reactants in same phase (homogeneous reactants) react faster than the reactants in different phase (heterogeneous reactants).

(b)-(iii) Rate of reaction also depends upon the surface area of reactant(s). More the surface area more will be the rate of reaction.

FACTORS AFFECTING RATE OF REACTION

(i) Nature of Reactant: The rate of reaction is strongly influenced by the chemical nature of the reactant, i.e.

(c) Stability of reactant:

Rate of reaction is more if the reactant is unstable and converts easily to product(s).

$$\text{Rate} \propto \frac{1}{\text{Stability of Reactants}}$$

$$\underset{\text{(Unstable)}}{Na} + H_2O \rightarrow \underset{\text{Stable}}{NaOH} + H_2$$

FACTORS AFFECTING RATE OF REACTION

In detail in upcoming lectures

(ii) Concentration of Reactant: Higher concentration → more particles per volume → more collisions → higher rate. of formation of products.

In detail in upcoming lectures

(iii) Temperature: Increase in temperature → particles move faster → more collisions → higher energy → faster reaction.

→ *Rate roughly doubles for every 10°C rise (rule of thumb).*

(we will also talk about catalyst)

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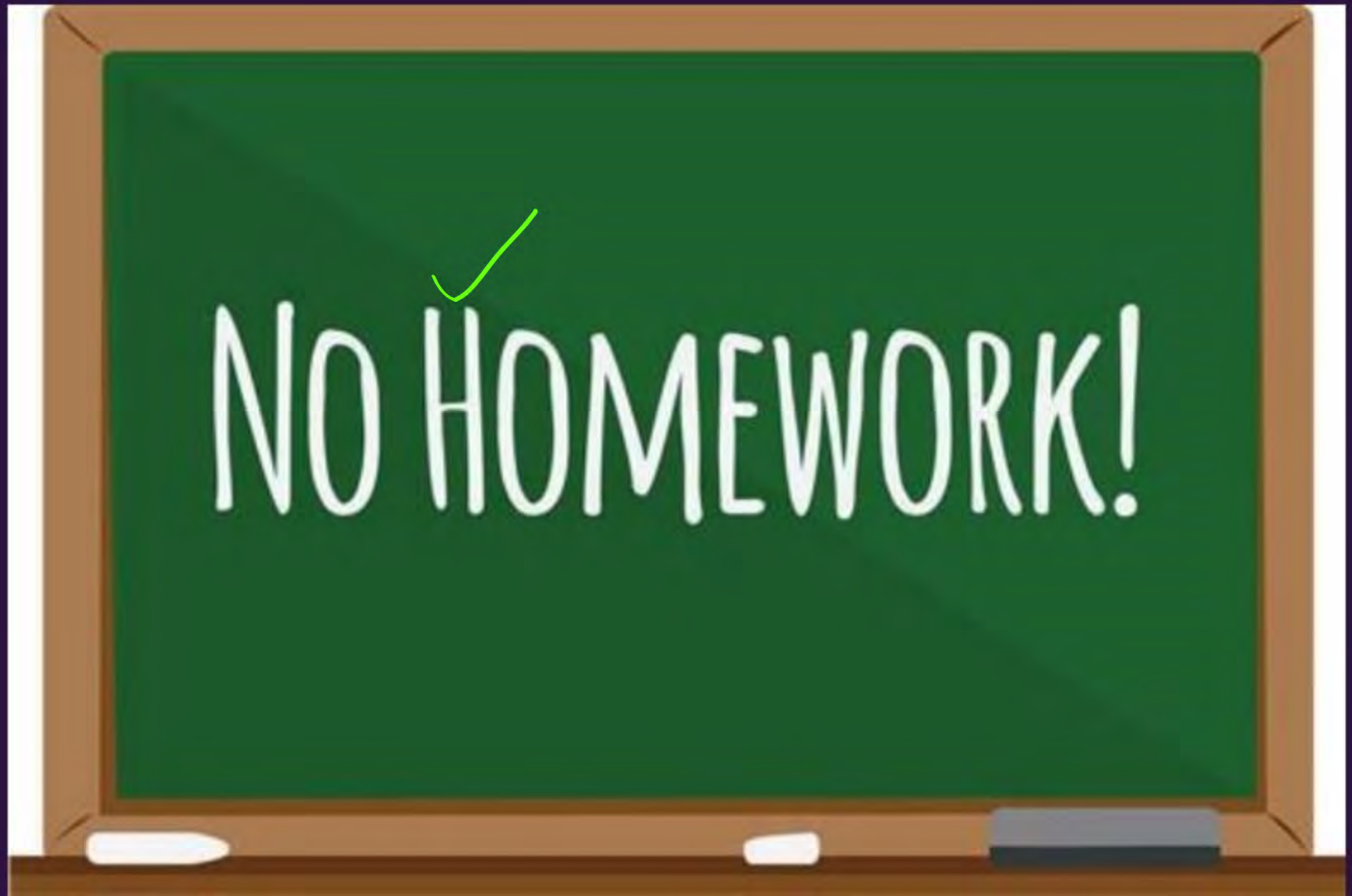
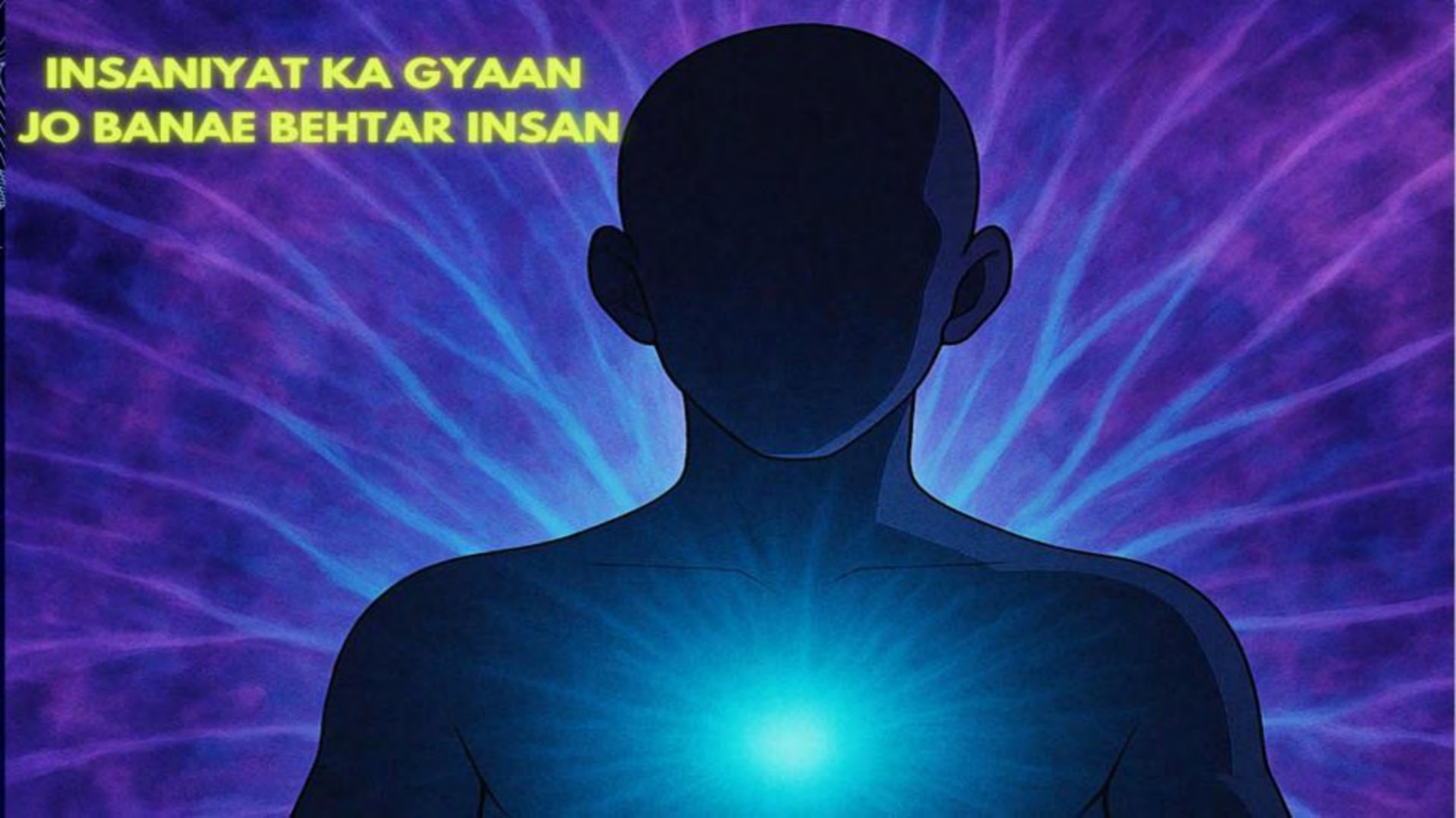
CONCEPT POLISH - HOMEWORK


NO HOMEWORK!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#pwsathhai ✓

Thank You

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CHEMISTRY

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Lecture 03

CHEMICAL KINETICS

(numerical practice)

Rate Law and Order of Reaction

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Rate Law or Rate Expression or Rate Equation

(ii) Order of Reaction

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- Rate of a reaction depends upon the concentration of reactants.
- The representation of rate of reaction in terms of concentration of the reactants is known as rate law. It is also called as rate equation or rate expression.

RATE LAW OR RATE EXPRESSION OR RATE EQUATION

RECAP – LAW OF MASS ACTION

(Class XI Chemical Equilibrium)

- Cato Guldberg and Peter Waage were colleagues and brother-in-law who formulated the Law of Mass Action in 1864.
- According to this law, the rate of reaction is directly proportional to the product of active masses of reactants (concentration of reactants) to the power of stoichiometric coefficients.

$$aA + bB \rightarrow cC + dD$$

Rate of rxn $\propto [A]^a [B]^b$

Rate of rxn $= k [A]^a [B]^b$

k → Constant of proportionality
OR
→ rate constant

THEORETICAL VS EXPERIMENTAL WAR

$$2NO(g) + 1O_2(g) \rightarrow 2NO_2(g)$$

(I) **THEORETICAL**

Rate of rxn = $k[NO]^2[O_2]^1$

(II) **EXPERIMENTAL**

Experiment	Initial [NO]/ mol L^{-1}	Initial $[O_2]$/ mol L^{-1}	Initial rate of formation of NO_2/ mol $L^{-1}s^{-1}$
1.	0.30	0.30	0.096
2.	0.60	0.30	0.384
3.	0.30	0.60	0.192
4.	0.60	0.60	0.768

Same, ×2 ; ×2, Same ; ×2, ×4

Rate of rxn = $k[NO]^2[O_2]^1$

Rate law

- When [NO] is doubled (keeping $[O_2]$ constant), rate becomes 4× → rate ∝ $[NO]^2$.

$R_1 = k[NO]^2[O_2]^1$

$R_2 = k[2NO]^2[O_2]^1 \Rightarrow R_2 = 4k\underbrace{[NO]^2[O_2]^1}_{R_1}$

- When $[O_2]$ is doubled (keeping [NO] constant), rate becomes 2× → rate ∝ $[O_2]^1$. Hence,

Do you think that everytime rate of rxn expression in terms of concentration of reactants will be same for theoretical & experimental ?

(A) Yes

(B) No

THEORETICAL VS EXPERIMENTAL WAR

THEORETICAL	EXPERIMENTAL
$1CHCl_3 + 1Cl_2 \rightarrow CCl_4 + HCl$	$CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$
$Rate = k[CHCl_3]^1[Cl_2]^1$	$Rate = k[CHCl_3]^1[Cl_2]^{\frac{1}{2}}$
$1CH_3COOC_2H_5 + 1H_2O \rightarrow CH_3COOH + C_2H_5OH$	$CH_3COOC_2H_5 + H_2O \rightarrow CH_3COOH + C_2H_5OH$
$Rate = k[CH_3COOC_2H_5]^1[H_2O]^1$	$Rate = k[CH_3COOC_2H_5]^1[H_2O]^0$

CONCLUSION

- While writing powers of concentration of reactants, the power may or may not be equal to stoichiometric coefficient. of reactant
- Hence, for any chemical reaction the **rate law or rate expression or rate equation** will be written as:

$$aA + bB \rightarrow cC + dD$$

Rate $\propto [A]^x [B]^y$

$$\text{Rate} = k[A]^x[B]^y$$

rate of rxn is directly proportional to product of concentration of reactants to the power of x & y (which may or may not be equal to stoichiometric coefficient)

CONCLUSION

- x and y in the rate law are the experimentally determined values.
- Some reactions happens in one/single step which are called elementary reactions and in those value of x and y are equal to stoichiometric coefficients of reactants.
- Some reactions occur in multiple steps, and the slowest (rate-determining) step controls the actual rate. These are called complex reactions and here value of x and y are not equal to stoichiometric coefficients of reactants.

GIVE A THOUGHT

Rate law for any reaction cannot be predicted by merely looking at the balanced chemical equation, i.e., theoretically but must be determined experimentally.

A. YES

B. NO

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ORDER OF A REACTION

ORDER OF A REACTION

$$\text{Rate} = k[A]^x[B]^y$$

- x and y: experimentally determined values which can be positive, negative, zero or even fractional
- x and y: sensitivity factor which tells how sensitive the rate is to changes in [A] and [B].
- x: order of reaction w.r.t A
- y: order of reaction w.r.t B
- n: overall order of reaction, n = x + y
- Order of reaction (n) can be 0, 1, 2, 3 or even a fraction.

ORDER OF A REACTION

	Order	Rate Law
(I)	0	Rate = k
(II)	1	Rate = k[A]^1
(III)	2	Rate = k[A]^2 ~~or k[A][B]~~

rate constant

OR

OR

$Rate = k[A]^0[B]^0$

$n = x + y = 0$

$Rate = k[A]^{.5}[B]^{.5}$

$n = x + y = .5 + .5 = 1$

$Rate = k[A]^1[B]^1$

(**) $n = x + y = 1 + 1 = 2$

- **Zero Order Reaction:** Rate of reaction does not depend on concentration of reactant.
- **First Order Reaction:** Rate of reaction is directly proportional to concentration of reactant. → only 1 reactant
- **Second Order Reaction:** Rate of reaction is proportional to the square of the reactant's concentration or to the product of two concentrations.

MORE ABOUBT RATE CONSTANT (K)

- Rate constant (k) is a proportionality constant in the rate law that relates the rate of reaction to the concentration of reactants.
- It depends only on temperature and catalyst.
- If k is large → reaction is fast
- If k is small → reaction is slow
- Unit of rate constant (k) is:

$$\text{Rate} = k[A]^x[B]^y$$

$$k = \frac{\text{Rate}}{[A]^x[B]^y}$$

$$k = \frac{mol \cdot L^{-1} \cdot s^{-1}}{(mol \cdot L^{-1})^x (mol \cdot L^{-1})^y}$$

$$\Rightarrow \frac{mol \cdot L^{-1} \cdot s^{-1}}{(mol \cdot L^{-1})^{x+y}}$$

$$\Rightarrow \frac{mol \cdot L^{-1} \cdot s^{-1}}{(mol \cdot L^{-1})^n}$$

$$\Rightarrow (mol \cdot L^{-1})^1 (mol \cdot L^{-1})^{-n} \cdot s^{-1}$$

$$\boxed{(mol \cdot L^{-1})^{1-n} \cdot s^{-1}}$$

(order of rxn)

$$n = x + y$$

UNITS OF RATE CONSTANT FOR DIFFERENT ORDER OF REACTION

$(mol \cdot L^{-1})^{1-n} \cdot s^{-1}$

Reaction	Order (n)	Units of rate constant
Zero order reaction	0	$(mol \cdot L^{-1})^{1-0} \cdot s^{-1} = \boxed{mol \cdot L^{-1} \cdot s^{-1}}$ (Unit of k = unit of rate of rxn)
First order reaction	1	$(mol \cdot L^{-1})^{1-1} \cdot s^{-1} = \boxed{s^{-1}}$
Second order reaction	2	$(mol \cdot L^{-1})^{1-2} \cdot s^{-1} \Rightarrow (mol \cdot L^{-1})^{-1} \cdot s^{-1}$ $\Rightarrow \boxed{mol^{-1} \cdot L^{1} \cdot s^{-1}}$

NCERT Expert

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LIKH DO.

AYE BHAIYA

NCERT WALLAH

NCERT SOLVED EXAMPLE 3.3

Calculate the overall order of a reaction which has the rate expression

(a) Rate = $k\,[A]^{1/2}\,[B]^{3/2}$

(b) Rate = $k\,[A]^{3/2}\,[B]^{-1}$

(a) $n = x + y$

$= \frac{1}{2} + \frac{3}{2} = \frac{4}{2} = 2 \rightarrow$ Second order rxn

(b) $n = x + y$

$= \frac{3}{2} - 1 = \frac{3-2}{2} = \frac{1}{2} = .5 \rightarrow$ Half order rxn

NCERT INTEXT 3.3

For a reaction, A + B → Product; the rate law is given by, $r = k[A]^{1/2}[B]^2$. What is the order of the reaction?

$$n = x + y$$

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

$$= \frac{1+4}{2}$$

$$= \frac{5}{2}$$

$$= 2.5$$

NCERT SOLVED EXAMPLE 3.4

Identify the reaction order from each of the following rate constants.

(i) $k = 2.3 \times 10^{-5}$ L mol^{-1} s^{-1} → second order rxn

(ii) $k = 3 \times 10^{-4}$ s^{-1} → first order rxn

NCERT EXERCISE 3.1

From the rate expression for the following reactions, determine their order of reaction and the dimensions of the rate constants.

(i) $3NO(g) \rightarrow N_2O\ (g)$ Rate $= k[NO]^2$

(ii) $H_2O_2\ (aq) + 3I^-\ (aq) + 2H^+ \rightarrow 2H_2O\ (l) + I_3^-$ Rate $= k[H_2O_2][I^-]$

(iii) $CH_3CHO\ (g) \rightarrow CH_4\ (g) + CO(g)$ Rate $= k[CH_3CHO]^{3/2}$

(iv) $C_2H_5Cl\ (g) \rightarrow C_2H_4\ (g) + HCl\ (g)$ Rate $= k[C_2H_5Cl]$

$(mol \cdot L^{-1})^{1-n} \cdot s^{-1}$

Rate Law Expression	Order of Reaction	Unit of Rate Constant k
Rate $= k[NO]^2$	$n = 2$	$mol^{-1} \cdot L^1 \cdot s^{-1}$
Rate $= k[H_2O_2]^1[I^-]^1$	$n = x + y = 2$	$mol^{-1} \cdot L^1 \cdot s^{-1}$ $= (mol \cdot L^{-1})^{-1.5} \cdot s^{-1}$
Rate $= k[CH_3CHO]^{3/2}$	$n = \frac{3}{2} = 1.5$	$[mol^{-.5} \cdot L^{.5} \cdot s^{-1}]$ $= (mol \cdot L^{-1})^{-.5} \cdot s^{-1}$
Rate $= k[C_2H_5Cl]^1$	$n = 1$	s^{-1}

NCERT EXERCISE 3.6

A reaction is second order with respect to a reactant. How is the rate of reaction affected if the concentration of the reactant is
(i) doubled (ii) reduced to half ?

$$\text{Rate} = k[A]^2$$

$$R_1 = k[A]^2$$

$[A] \rightarrow [2A]$

(i) $R_2 = k[2A]^2$

$R_2 = 4\overbrace{k[A]^2}^{R_1}$

$R_2 = 4R_1$

(Increased by 4 times)

$[A] = [\frac{1}{2}A]$

(ii) $R_3 = k[\frac{A}{2}]^2$

$R_3 = \frac{1}{4}k[A]^2$

$R_3 = \frac{1}{4}R_1$

(reduced by $\frac{1}{4}$ times)

NCERT EXERCISE 3.9

A reaction is first order in A and second order in B.

(i) Write the differential rate equation.

(ii) How is the rate affected on increasing the concentration of B three times?

(iii) How is the rate affected when the concentrations of both A and B are doubled?

$\text{Rate} = k[A]^1[B]^2 \rightarrow R_1 = k[A]^1[B]^2$

(i) $\boxed{-\frac{d[R]}{dt} = k[A]^1[B]^2}$ → Differential rate eqn

(ii) $[B] \rightarrow [3B]$

$R_2 = k[A]^1[3B]^2$

$R_2 = k[A]^1\, 9[B]^2$

$R_2 = 9\,k[A]^1[B]^2$

$\boxed{R_2 = 9R_1}$

(iii) $[A] \rightarrow [2A]$

$[B] \rightarrow [2B]$

$R_3 = k[2A]^1[2B]^2$

$R_3 = k\,2[A]^1\,4[B]^2$

$R_3 = 8\,k[A]^1[B]^2$

$\boxed{R_3 = 8R_1}$

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CONCEPT POLISH - HOMEWORK

NCERT EXERCISE 3.10

In a reaction between A and B, the initial rate of reaction (r_0) was measured for different initial concentrations of A and B as given below:

H.W.

A/ mol L^{-1}	0.20	0.20	0.40
B/ mol L^{-1}	0.30	0.10	0.05
r_0/mol $L^{-1}s^{-1}$	5.07×10^{-5}	5.07×10^{-5}	1.43×10^{-4}

What is the order of the reaction with respect to A and B?

$\log(2.82) = 0.4502$ and $\log(2) = 0.3010$

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CHEMISTRY

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Lecture 04

CHEMICAL KINETICS

CBSE PYQs' and Molecularity of Reaction

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) CBSE Previous Year Questions ✓

(ii) Molecularity of Reaction ✓

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CONCEPT POLISH – HOMEWORK DISCUSSION

NCERT EXERCISE 3.10

$$\text{Rate} = k[A]^x[B]^y$$

In a reaction between A and B, the initial rate of reaction (r_0) was measured for different initial concentrations of A and B as given below:

	I	II	III
A/ mol L^{-1}	0.20	0.20	0.40
B/ mol L^{-1}	0.30	0.10	0.05
r_0/mol $L^{-1}s^{-1}$	5.07×10^{-5}	5.07×10^{-5}	1.43×10^{-4}

same

different

change

What is the order of the reaction with respect to A and B?

$\log(2.82) = 0.4502$ and $\log(2) = 0.3010$

From observation:

Conc. of [A] in I & II → same

–"– [B] in I & II → different

Rate of rxn is same in I & II → Zero order w.r.t. B

$(y=0)$

From formula

$R_1 = k[A]^x[B_1]^y \Rightarrow k[A]^x = \frac{R_1}{[B_1]^y}$ — (1)

$R_2 = k[A]^x[B_2]^y \Rightarrow k[A]^x = \frac{R_2}{[B_2]^y}$ — (2)

$$\frac{R_1}{[B_1]^y} = \frac{R_2}{[B_2]^y}$$

$$\Rightarrow \frac{[B_2]^y}{[B_1]^y} = \frac{R_2}{R_1}$$

$$\Rightarrow \left[\frac{.10}{.30}\right]^y = \frac{5.07 \times 10^{-5}}{5.07 \times 10^{-5}}$$

$$\Rightarrow \left[\frac{.1\cancel{0} \times \cancel{100}}{.3\cancel{0} \times \cancel{100}}\right]^y = 1$$

$$\Rightarrow \frac{1}{3^y} = 1$$

$$\Rightarrow 1 = 3^y$$

$$\Rightarrow 3^0 = 3^y \Rightarrow y = 0$$

$$R = K[A]^x[B]^0 = K[A]^x$$

$$R_1 = K[A_1]^x, \quad R_3 = K[A_3]^x$$

$$\boxed{\frac{R_1}{[A_1]^x} = K} \quad (1) \qquad \boxed{\frac{R_3}{[A_3]^x} = K} \quad (2)$$

from (1) & (2)

$$\Rightarrow \frac{R_1}{[A_1]^x} = \frac{R_3}{[A_3]^x}$$

$$\Rightarrow \frac{R_1}{R_3} = \frac{[A_1]^x}{[A_3]^x}$$

$$\Rightarrow \frac{5.07 \times 10^{-5}}{1.43 \times 10^{-4}} = \left(\frac{.20}{.40}\right)^x$$

$$\frac{5.07 \times 10^{-1}}{1.43} = \frac{1}{(2^x)}$$

$$2^x = \frac{1.43}{5.07 \times 10^{-1}}$$

$$2^x = \frac{14.3}{5.07}$$

$$\boxed{2^x = 2.82}$$

Taking log both sides

$$x \log_{10} 2 = \log_{10} 2.82$$

$$x \times .3010 = .4502$$

$$x = \frac{.4502}{.3010}$$

$$\approx \frac{.45 \times 100}{.30 \times 100}$$

$$\approx \frac{45}{30}$$

$$\approx \boxed{1.5}$$

Order of rxn w.r.t A

$$1.5$$

$$\boxed{R = K[A]^{1.5}[B]^{0}}$$

CBSE PREVIOUS YEAR QUESTIONS

CBSE PYQ

For the reaction $A \rightarrow B$, the rate of reaction becomes three times when the concentration of A is increased by nine times. What is the order of reaction? [CBSE]

$R_1 = K[A_1]^x \Rightarrow \boxed{\frac{R_1}{[A_1]^x} = K}$ — (1)

$R_2 = K[A_2]^x \Rightarrow \boxed{\frac{R_2}{[A_2]^x} = K}$ — (2)

from (1) & (2)

$\Rightarrow \frac{R_1}{[A_1]^x} = \frac{R_2}{[A_2]^x}$

$\Rightarrow \frac{[A_2]^x}{[A_1]^x} = \frac{R_2}{R_1}$

$\Rightarrow \left(\frac{9\cancel{A_1}}{\cancel{A_1}}\right)^x = \frac{3\cancel{R_1}}{\cancel{R_1}}$

$\Rightarrow 9^x = 3$

$3^{2x} = 3^1$

$2x = 1$

$x = \frac{1}{2}$

CBSE PYQ

For the reaction $2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$, the rate of formation of $NO_2(g)$ is $2.8 \times 10^{-3}\ M\,s^{-1}$. Calculate the rate of disappearance of $N_2O_5(g)$. [CBSE]

$M = \frac{mol}{L} = mol \cdot L^{-1}$

$$\text{Rate of rxn} = -\frac{1}{2}\frac{\Delta[N_2O_5]}{\Delta t} = +\frac{1}{4}\frac{\Delta[NO_2]}{\Delta t} = +\frac{1}{1}\frac{\Delta[O_2]}{\Delta t}$$

$$\text{Rate of disappearance of } N_2O_5 = -\frac{\Delta[N_2O_5]}{\Delta t} = \frac{\cancel{2}^{1}}{\cancel{4}_{2}} \times 2.8 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1}$$

$$= \boxed{1.4 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1}}$$

MOLECULARITY OF REACTION

MOLECULARITY OF REACTION

- Molecularity of a reaction is the number of reactant molecules or species that collide simultaneously in an elementary reaction to bring about a chemical change.
(one step rxn)

HINGLISH MEANING: Kitne molecules ek saath takra ke reaction shuru karte hain (in one step)

Type	Example
Unimolecular (molecularity = 1)	$1O_3 \rightarrow O_2 + O$
Bimolecular (— " — = 2)	$1NO + 1O_3 \rightarrow NO_2 + O_2$
Trimolecular (— " — = 3)	$2NO + 1O_2 \rightarrow 2NO_2$

GIVE A THOUGHT

How molecularity can be one?

$$1O_3 \xrightarrow{U.V.} O_2 + O$$

(Ozone)

Here or in any unimolecular rxn there is absorption of energy (energy is not reactant).

GIVE A THOUGHT

The probability that more than three molecules can collide and react simultaneously is very small.

A. YES

B. NO

GIVE A THOUGHT

The probability that more than three molecules can collide and react simultaneously is very small.

A. YES

B. NO

YES!

(in 1 step)

For a reaction to occur all molecules should come very close, have enough energy and correct orientation. at one time.

Now, having all these with 4 or more molecules is a very rare. thing.

CHARACTERISTICS OF MOLECULARITY OF REACTION

- It is always a whole number 1, 2, 3, ...
- It can never be 0 or fractional.
- Only defined for elementary reactions (happens in one/single step) and not for overall reaction of a complex reaction (happens in multiple steps).

Complex Rxn

Molecularity (✓) ← Step I → $A + B \rightarrow C$

Molecularity (✓) ← Step II → $C \longrightarrow D + E$

Molecularity (x) ← Overall rxn → $\boxed{A + B \rightarrow D + E}$

WHY NOT FOR COMPLEX REACTIONS?

$$1KClO_3 + 6FeSO_4 + 3H_2SO_4 \rightarrow KCl + 3Fe_2(SO_4)_3 + 3H_2O$$

- After looking at the overall reactions it looks like 10 molecules are at colliding and reacting at the same time which is practically impossible.
- It is a complex reaction which happens in multiple steps.
- As this reaction occurs in multiple steps so the ones written in overall reaction are not the REAL REACTANTS colliding with each other. So, molecularity of the overall reaction can't be written.

LET'S ANALYSE!

(molecularity is not defined)

Overall Rxn →

$$KClO_3 + 6FeSO_4 + 3H_2SO_4 \rightarrow KCl + 3Fe_2(SO_4)_3 + 3H_2O$$

Step 1

Fe^{2+} → ferrous ion

ClO_3^- → chlorate ion

$$1Fe^{2+} + 1ClO_3^- + 1H^+ \rightarrow Fe^{3+} + ClO_2 + OH^-$$

ferric

- Involves collision of Fe^{2+}, ClO_3^-, H^+ (trimolecular)

Step 2

$$1ClO_2 + 1H_2O \rightarrow Cl^- + H^+ + O_2$$

- Involves collision of ClO_2 and H_2O (*bimolecular*)

ORDER OF REACTION

Relay Race

$$KClO_3 + 6FeSO_4 + 3H_2SO_4 \rightarrow KCl + 3Fe_2(SO_4)_3 + 3H_2O$$

Step 1 (Slow, Rate-Determining): → (RDS)

Step

$$Fe^{2+} + ClO_3^- + H^+ \rightarrow Fe^{3+} + ClO_2 + OH^-$$

- Involves collision of Fe^{2+}, ClO_3^-, H^+ (trimolecular)
- This step limits the overall reaction rate

- So, the rate law of this reaction will be: $Rate = K[Fe^{2+}]^1[ClO_3^{1-}]^1[H^+]^0$
- The order of this reaction will be: $n = 1 + 1 = 2$

LET'S ANALYSE!

- Consider the decomposition of hydrogen peroxide which is catalysed by iodide ion in an alkaline medium

we don't define molecularity of this

Rxn I

$$2H_2O_2 \xrightarrow[\text{Alkaline medium}]{I^-} 2H_2O + O_2$$

⟹ overall rxn of a complex rxn

Step 1 (slow):

(Rate determing step) $1H_2O_2 + 1I^- \rightarrow IO^- + H_2O$ (Molecularity = 2)

hypoiodite ion

Step 2 (fast):

$1IO^- + 1H_2O_2 \rightarrow I^- + H_2O + O_2$ (Molecularity = 2)

- So, the rate law of this reaction will be: (Rxn I) Rate = $K[H_2O_2]^1[I^-]^1$
- The order of this reaction will be: $n = 1+1 = 2$

DIFFERENCE BETWEEN ORDER AND MOLECULARITY OF REACTION

Feature	Molecularity	Order of Reaction
Definition	Number of reactant species colliding **simultaneously in an elementary step**	Sum of powers of concentration terms in the **experimentally determined rate law**
How it is obtained	From **reaction mechanism** (theoretical)	From **experimental data** (rate vs concentration)
Applicable to	Only **elementary reactions**	Both **elementary and complex reactions**
Value type	Always a **whole number** (1, 2, 3...)	Can be **0, fractional** or **whole number**
Zero possible?	❌ Never zero	☑ Can be zero
Fractional possible?	❌ Never fractional	☑ Can be fractional

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CBSE PREVIOUS YEAR QUESTIONS

CBSE PYQ

For a reaction: $2NH_3(g) \xrightarrow{Pt.} N_2(g) + 3H_2(g)$

$\boxed{\text{Rate} = k}$

(*i*) Write the order and molecularity of this reaction.

(*ii*) Write the unit of *k*.

(i) (a) Zero order rxn

(b) 2

(ii) Unit of K in zero order rxn → $mol \cdot L^{-1} \cdot s^{-1}$

CBSE PYQ

For a reaction: $H_2 + Cl_2 \xrightarrow{h\nu} 2HCl$

$\boxed{\text{Rate} = k}$

(i) Write the order and molecularity of this reaction.

(ii) Write the unit of k.

(i) (a) Zero order rxn

(b) 2

(ii) $k \rightarrow mol \cdot L^{-1} \cdot s^{-1}$

CBSE PYQ

Define order of reaction. How does order of a reaction differ from molecularity for a complex reaction? [AI Chennai]

Slide no. 21

CBSE PYQ

Write two differences between order of a reaction and molecularity of a reaction.

Slide no. 21

CBSE PYQ

1. For a reaction

$$2H_2O_2 \xrightarrow[\text{alkaline medium}]{I^-} 2H_2O + O_2$$

the proposed mechanism is as given below:

1. $H_2O_2 + I^- \longrightarrow H_2O + IO^-$ (slow)
2. $H_2O_2 + IO^- \longrightarrow H_2O + I^- + O_2$ (fast)

(*i*) Write rate law for the reaction.

(*ii*) Write the overall order of reaction.

(*iii*) Out of steps (1) and (2), which one is rate determining step?

(i) (Rate determining step) Rate $= k[H_2O_2]^1[I^-]^1$

(ii) Order of rxn (n) $= 1+1 = 2$

(iii) Step I is RDS

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CONCEPT POLISH - HOMEWORK

CBSE PYQ

The reaction between A and B is of first order with respect to A and zero order with respect to B. Fill in the blanks in the following table:

Experiment	[A]/mol L^{-1}	[B]/mol L^{-1}	Initial rate in mol L^{-1} $minute^{-1}$
I	0.1	0.1	2.0×10^{-2}
II		0.2	4.0×10^{-2}
III	0.4	0.4	
IV		0.2	2.0×10^{-2}

[AI Chandigarh]

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Sunday ko extra class?

(A) Thumbs-up

(B) -"- -down

Thank You

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Chemistry

Lecture 05

CHEMICAL KINETICS

Integrated Rate Equation – Introduction and Zero Order

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Integrated Rate Equation – Zero Order

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CONCEPT POLISH – HOMEWORK DISCUSSION

CBSE PYQ

(NCERT Exercise Question)

The reaction between A and B is of first order with respect to A and zero order with respect to B. Fill in the blanks in the following table:

Experiment	[A]/mol L^{-1}	[B]/mol L^{-1}	Initial rate in mol L^{-1} $minute^{-1}$
I	0.1	0.1	2.0×10^{-2}
II	.2	0.2	4.0×10^{-2}
III	0.4	0.4	8.0×10^{-2}
IV	.1	0.2	2.0×10^{-2}

[AI Chandigarh]

$Rate = K[A]^1[B]^0$

$\boxed{Rate = K[A]^1}$ ✓

Let's calculate 'k'

$R_1 = K[A_1]$

$\Rightarrow 2.0 \times 10^{-2} = k \times \frac{1}{10}$

$k = 2.0 \times 10^{-2} \times 10^1$

$= 2.0 \times 10^{-1}$

$= \frac{2}{10^1} = .2\ min^{-1}$

$$R_2 = k[A_2]$$

$$= \cancel{4}^{2} \times 10^{-2} = \frac{\cancel{2}^{1}}{10} \times [A_2]$$

$$= 2 \times 10^{-2} \times 10^{1} = [A_2]$$

$$= \frac{2}{10^{1}} = [A_2]$$

$$= \boxed{.2 \text{ mol} \cdot L^{-1} = [A_2]}$$

$$R_3 = k[A_3]$$

$$= .2 \times .4 = .08 = \frac{.08}{100} = \underline{\underline{8 \times 10^{-2}}} \text{ mol} \cdot L^{-1} \cdot \text{min}^{-1}$$

$$R_4 = k[A_4]$$

$$\cancel{2}.0 \times 10^{-2} = \frac{.\cancel{2}}{10} \times [A_4]$$

$$= 10^{-2} \times 10^{1} = [A_4]$$

$$= \frac{1}{10^{1}} = [A_4]$$

$$= \boxed{.1 \text{ mol} \cdot L^{-1} = [A_4]}$$

- Average rate ek time interval mein speed of reaction btata hai aur instantenous rate ek particular instant par lekin reaction ki speed ka concentration of reactant se relationship nahi btata (order of reaction) jo ki integrated rate eq^{n+} btata h.

I

Need

INTEGRATED RATE EQUATION – (ZERO ORDER)

Need

II

- Reactions ka rate time ke saath change hota rehta hai. Agar hum sirf **average rate** dekhen, toh hum yeh nahi samajh sakte ki reaction ka speed kaise slow ya fast ho raha hai. **Integrated rate law** se hum **har moment ka concentration** samajh sakte hain aur isse humein reaction ko **better control karne** mein madad milti hai.

ZERO ORDER REACTIONS - NCERT

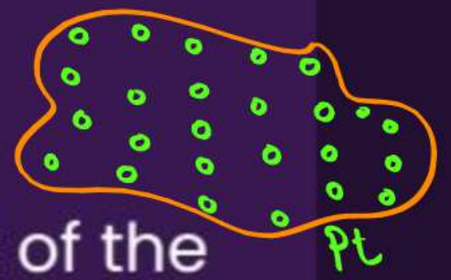
[Pt & Au] → 'catalyst'

(I) $2NH_3(g) \xrightarrow[\text{Pt catalyst}]{1130K} N_2(g) + 3H_2(g)$

Rate = $K[NH_3]^0$

(II) $2\,HI(g) \xrightarrow{\text{Au, heat}} H_2(g) + I_2(g)$

Rate = $K[HI]^0$



- In both the reactions, [rate of reaction is independent of the concentration of reactant molecules.]
- Here, rate of reaction depends on the surface area of catalyst, i.e. number of available sites on the catalyst.
- Once all the surface sites are filled, the reaction can't go any faster, regardless of how much more reactant is added.

INTEGRATED RATE EQⁿ – (ZERO ORDER)

- Zero order reaction means that the rate of the reaction is proportional to zero power of the concentration of reactants. Consider the reaction, R → P

$$\text{Rate} \propto [R]^0$$

$$\text{Rate} = k[R]^0 = k \times 1$$

$$-\frac{d[R]}{dt} = k$$

$$d[R] = -k\,dt$$

$t = 0$, conc. of reactant $\rightarrow [R_0]$

$t = t$, —"— $\rightarrow [R_t]$

$$\int_{[R_0]}^{[R_t]} d[R] = -k \int_0^t dt$$

$$\left\{[R]\right\}_{[R_0]}^{[R_t]} = -k\,[t]_0^t$$

$$[R_t] - [R_0] = -kt$$

multiply with (−) **

$$\boxed{[R_0] - [R_t] = kt} \quad (1)$$

**

$$\boxed{[R_0] - kt = [R_t]}$$

↓

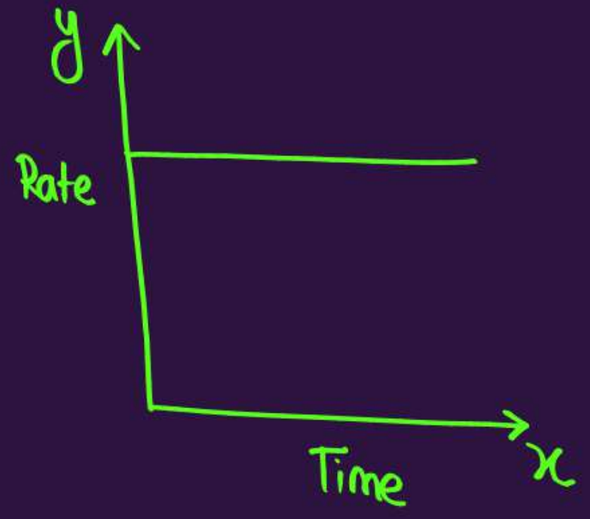
**

$$\boxed{[R_t] = -kt + [R_0]} \quad (2)$$

GRAPHS

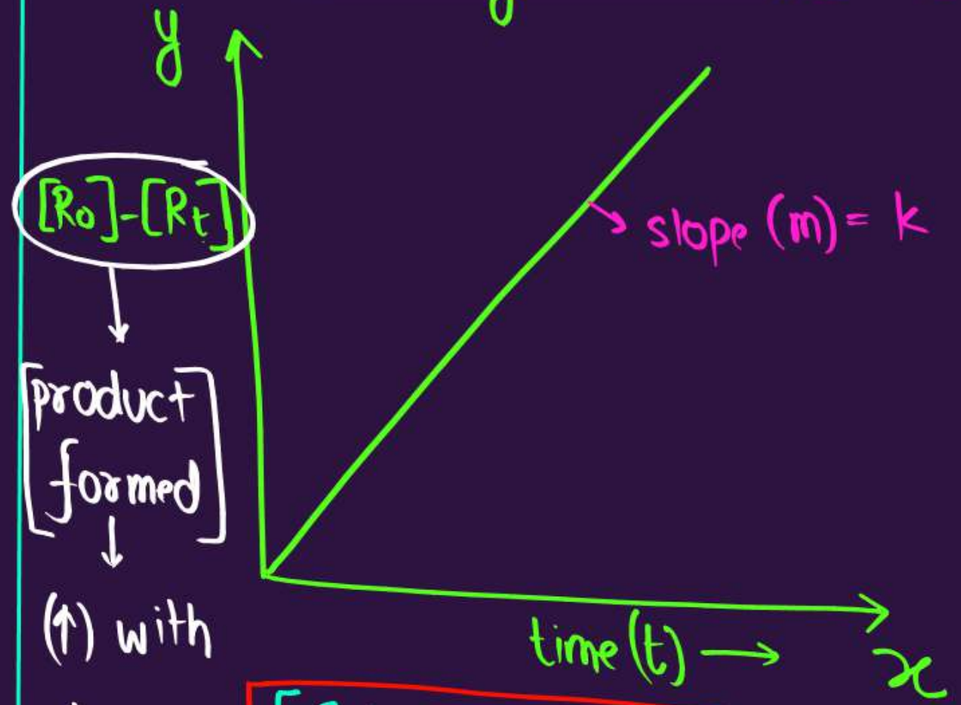
(I)

Rate = k



(II)

$$\underbrace{[R_o] - [R_t]}_{y} = \underbrace{k}_{m}\underbrace{t}_{x} + \underbrace{0}_{c}$$

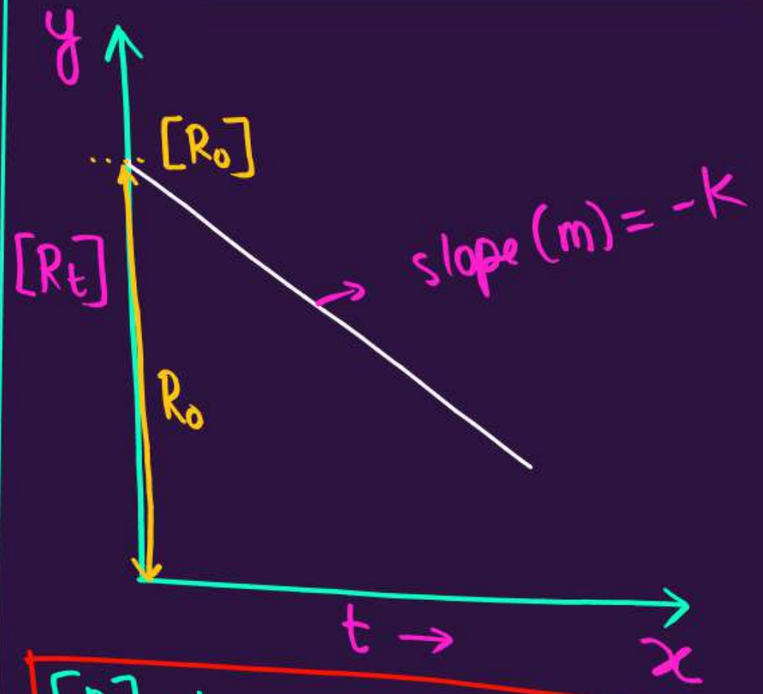


[P] (↑) with time

(III)

$$[R_t] = \underbrace{-k}_{m}\underbrace{t}_{x} + \underbrace{[R_o]}_{c}$$

$$\cancel{[R_t] = [R_o] - kt}$$



[R] decreases with time

GIVE A THOUGHT

What will be the concentration of product formed at time 't' when concentration of reactant is $[R_t]$?

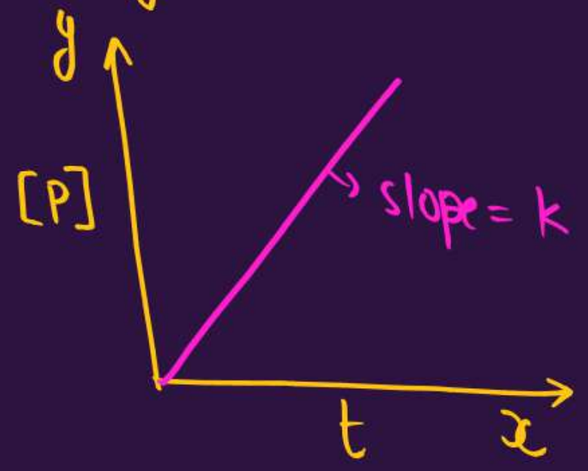
Given: Initial conc. of reactant → $[R_0]$

	[R]	[P]
$t=0$,	$[R_0]$	0
$t=t$,	$[R_t]=[R_0]-kt$	kt

$[P] = kt + 0$

$y = mx + c$

→ same as graph (II)



LET'S PRACTICE

QUESTION

Calculate the amount of product formed after t = 10 minute.
If rate constant (k) = 5 $mol \cdot L^{-1} \cdot s^{-1}$

$[R_0] - [R_t] = kt$

OR

$[P] = kt$

$= 5\ mol \cdot L^{-1} \cdot s^{-1} \times 600\ s$

$= 3000\ mol \cdot L^{-1}$

1 minute = 60 s

10 minute = 'x'

$x = 60 \times 10$

$= 600\ s$

QUESTION

For a zero-order reaction, the concentration of reactant decreases from 0.90 mol/L to 0.45 mol/L in 150 seconds. Calculate the rate constant k for the reaction.

$$[R_0] - [R_t] = kt$$

$$= .90 \frac{mol}{L} - .45 \frac{mol}{L} = k \times 150\ s$$

$$= \frac{.45}{150 \times 100}\ mol \cdot L^{-1} \cdot s^{-1} = k$$

$$= \frac{\cancel{45}^{3}}{\cancel{15} \times 10^{3}}\ mol \cdot L^{-1} \cdot s^{-1} = k$$

$$= \boxed{3 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1} = k}$$

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HALF LIFE OF A ZERO ORDER REACTION

The half-life of a reaction is the time in which the concentration of a reactant is reduced to one half of its initial concentration. It is represented as $t_{1/2}$ or $t_{50\%}$.

$$[R_0] - [R_t] = kt \quad -(1)$$

Now, $t = t_{1/2}$ or $t_{50\%}$ $-(2)$

$$[R_t] = \frac{[R_0]}{2} \quad -(3)$$

Putting value of (2) & (3) in (1)

$$\Rightarrow [R_0] - \frac{[R_0]}{2} = kt_{50\%}$$

$$\Rightarrow \frac{[R_0]}{2k} = t_{50\%}$$

$$t_{50\%} \propto [R_0]$$

$$t_{50\%} \propto \frac{1}{k}$$

$$[R_0] \xrightarrow{t_{50\%}} \frac{[R_0]}{2} \xrightarrow{t_{100\%}} 0$$

LET'S PRACTICE!

The completion time of a reaction is the time in which the concentration of a reactant is reduced to zero. It is represented as $t_{100\%}$.

$[R_0] - [R_t] = kt$ — (1)

$t = t_{100\%}$ — (2)

$[R_t] = 0$ — (3)

Putting values of (2) & (3) in (1)

$[R_0] - 0 = kt_{100\%}$

$$\boxed{\frac{[R_0]}{k} = t_{100\%}}$$

$t_{100\%} = \frac{[R_0]}{K}$, $t_{50\%} = \frac{[R_0]}{2K}$

$\Rightarrow t_{50\%} \times 2 = \frac{[R_0]}{K}$

$\Rightarrow \boxed{t_{50\%} \times 2 = t_{100\%}}$

LET'S PRACTICE!

Now calculate the same for $t_{75\%}$ → Rxn is 75% or $\frac{3}{4}$ completed & now only $\frac{1}{4}$ of reactant is left

$$[R_0]-[R_t] = kt \quad -(1)$$

$$\rightarrow \quad t = t_{75\%} \quad -(2)$$

$$\rightarrow \quad [R_t] = \frac{1}{4}[R_0] \quad -(3)$$

Put values of (2) & (3) in (1)

$$\Rightarrow \quad [R_0] - \frac{[R_0]}{4} = kt_{75\%}$$

$$\Rightarrow \quad \frac{3[R_0]}{4k} = t_{75\%}$$

$$\frac{3}{2}\underbrace{\frac{[R_0]}{2k}}_{t_{50\%}} = t_{75\%}$$

$$\boxed{\frac{3}{2}t_{50\%} = t_{75\%} \quad \text{OR} \quad 1.5 \times t_{50\%} = t_{75\%}}$$

LET'S PRACTICE!

Now calculate the same for $t_{25\%}$ → Rxn is 25% complete & $\frac{3}{4}$ or 75% of reactant is still left

$[R_0] - [R_t] = kt$ — (1)

$t = t_{25\%}$ — (2)

$[R_t] = \frac{3}{4}[R_0]$ — (3)

Putting values of (2) & (3) in (1)

$\Rightarrow [R_0] - \frac{3}{4}[R_0] = k t_{25\%}$

$\Rightarrow \frac{[R_0]}{4k} = t_{25\%}$

$\frac{1}{2}\frac{[R_0]}{2k} = t_{25\%}$

$t_{50\%}$

$\frac{1}{2} \times t_{50\%} = t_{25\%}$

OR

$.5 \times t_{50\%} = t_{25\%}$

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LET'S PRACTICE

QUESTION

The half-life of a zero-order reaction is 50 seconds. If the initial concentration of the reactant is 0.60 mol/L calculate the rate constant k.

$t_{1/2}$ or $t_{50\%} = 50\,s$

$[R_0] = .60\ mol/L$

$k = ?$

$$\boxed{t_{50\%} = \frac{[R_0]}{2k}}$$

$$k = \frac{[R_0]}{2 \times t_{50\%}} = \frac{6\cancel{0}}{2 \times 5\cancel{0}} \times \frac{1}{100} = \frac{6}{10 \times 100} = \boxed{6 \times 10^{-3}\ mol \cdot L^{-1} \cdot s^{-1}}$$

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CONCEPT POLISH - HOMEWORK

(NO HOMEWORK)

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CHEMISTRY

Bharat Mata Ki Jai ♡

Lecture 06

CHEMICAL KINETICS

Integrated Rate Equation – First Order

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Integrated Rate Equation – First Order ✓

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INTEGRATED RATE EQUATION – FIRST ORDER

FIRST ORDER REACTIONS

addition of H_2 on ethene

for boards

I. Hydrogenation of ethene is an example of first order reaction.

$$C_2H_4(g) + H_2(g) \xrightarrow[200\,^\circ C]{Ni/Pd/Pt} C_2H_6(g)$$

(ethene) (ethane)

$$\boxed{Rate = k[C_2H_4]^1}$$

II. All natural and artificial radioactive decay of unstable nuclei take place by first order kinetics.

α- Decay

α-particle OR

$$^{226}_{88}Ra \rightarrow {}^{4}_{2}He^{2+} + {}^{222}_{86}Rn$$

$$\boxed{Rate = k[Ra]^1}$$

- **EXTRA (Out-of-syllabus):** Hydrogenation of ethene does not follow first-order kinetics and is better explained on the basis of Langmuir-Hinshelwood mechanism.

INTEGRATED RATE EQ^n – FIRST ORDER

- First order reaction means that the rate of the reaction is proportional to first power of the concentration of reactants. Consider the reaction, $\boxed{R \rightarrow P}$

$$\text{Rate} \propto [R]^1$$

$$\text{Rate} = k[R]$$

$$-\frac{d[R]}{dt} = k[R]$$

$$\frac{1}{[R]}d[R] = -k\,dt$$

$t=0$, conc. of reactant $\rightarrow [R_0]$

$t=t$, —"— $\rightarrow [R_t]$

$$\Rightarrow \int_{[R_0]}^{[R_t]} \frac{1}{[R]}\,d[R] = -k\int_0^t dt$$

$$\Rightarrow \Big\{\ln[R]\Big\}_{[R_0]}^{[R_t]} = -kt$$

$$\Rightarrow \boxed{\ln[R_t] - \ln[R_0] = -kt} \quad (**)$$

$\ln \rightarrow$ natural log

$$\boxed{\ln x = \log_e x}$$

$$\ln[R_t] - \ln[R_0] = -kt$$

multiply entire eqn by (-1)

$$\boxed{\ln[R_0] - \ln[R_t] = kt} \quad (*)$$

$$\boxed{\ln\left(\frac{[R_0]}{[R_t]}\right) = kt} \quad (**)$$

$$\ln[R_0] - kt = \ln[R_t]$$

OR

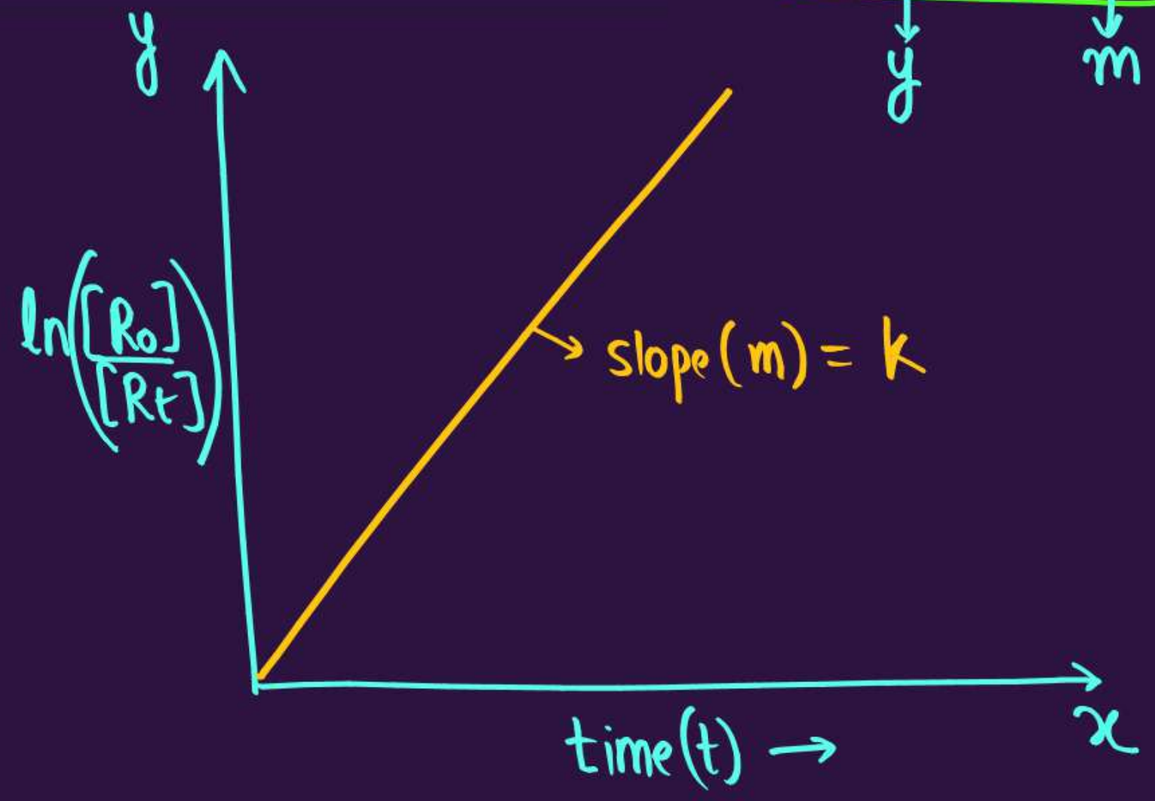
$$\boxed{\ln[R_t] = -kt + \ln[R_0]} \quad (*)$$

GRAPHS

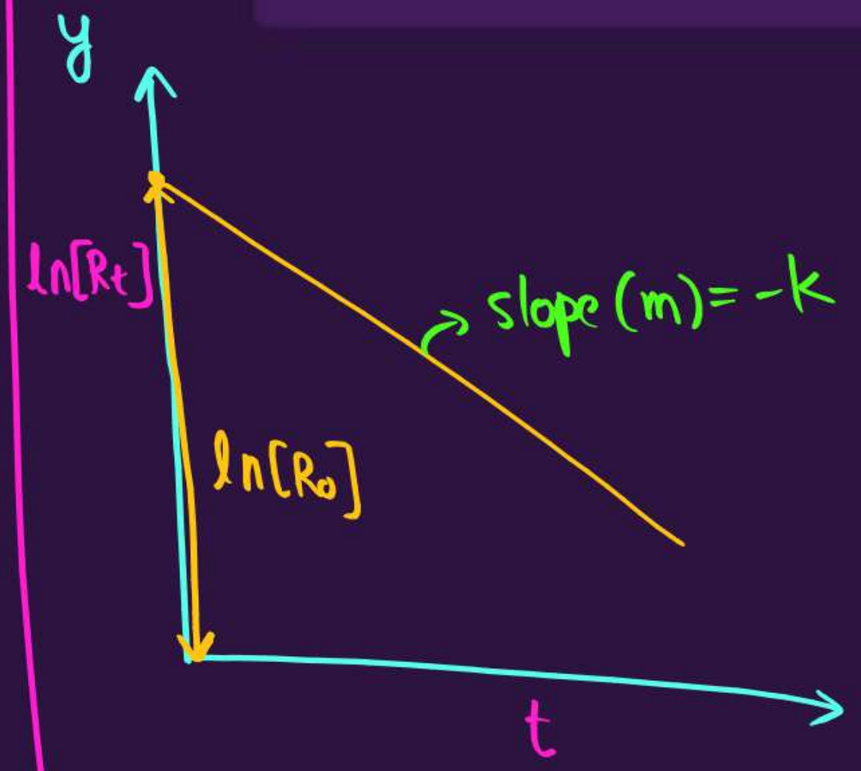
$$\ln[R_o] - \ln[R_t] = kt$$

OR

$$\underbrace{\ln\left(\frac{[R_o]}{[R_t]}\right)}_{y} = \underbrace{k}_{m}\,\underbrace{t}_{x} + \underbrace{0}_{c}$$



$$\underbrace{\ln[R_t]}_{y} = \underbrace{-k}_{m}\,\underbrace{t}_{x} + \underbrace{\ln[R_o]}_{c}$$



DIFFERENT FORMS

$\ln(x)$

$\ln\frac{[R_0]}{[R_t]} = kt$ This can also be written as:

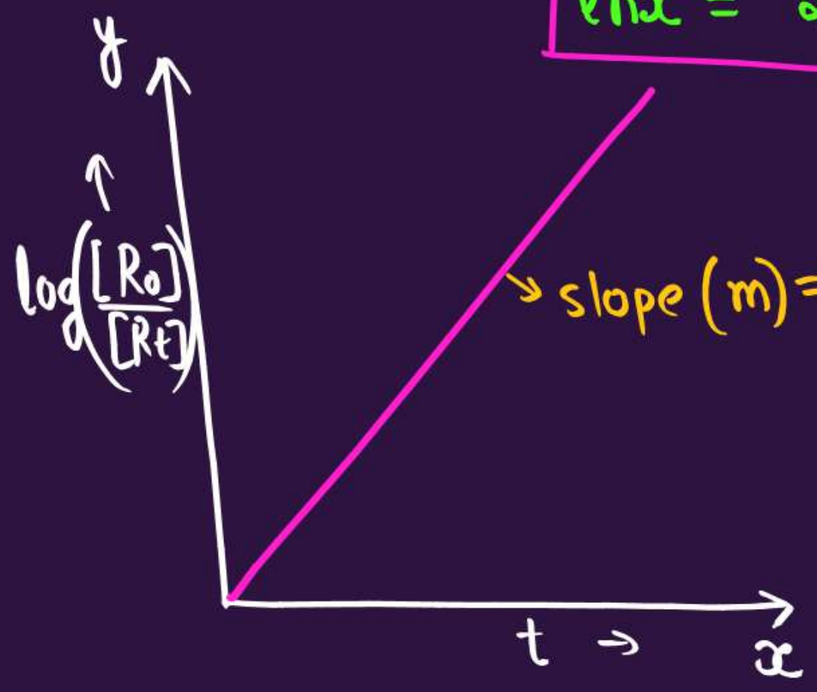
$$2.303 \log_{10}\left(\frac{[R_0]}{[R_t]}\right) = kt$$

$$\underbrace{\log\left(\frac{[R_0]}{[R_t]}\right)}_{y} = \underbrace{\frac{kt}{2.303}}_{m} + \underbrace{0}_{c}$$

$kt \rightarrow x$

$\ln x = \log_e x$

$\ln x = 2.303 \log_{10} x$



slope $(m) = \frac{k}{2.303}$

DIFFERENT FORMS

$\ln\frac{[R_0]}{[R_t]} = kt$ This can also be written as:

$\ln x$

$\Rightarrow \log_e\left(\frac{[R_0]}{[R_t]}\right) = kt$

$\Rightarrow$ Taking antilog with base e on both sides

$\Rightarrow \frac{[R_0]}{[R_t]} = \text{antilog}_e(kt)$

$\Rightarrow \frac{[R_0]}{[R_t]} = e^{kt}$

$\boxed{\frac{[R_0]}{e^{kt}} = [R_t]} \rightarrow \boxed{[R_0]e^{-kt} = [R_t]}$ (*)

$\boxed{\ln x = \log_e x}$

$10^3 = 1000$

$\rightarrow \log_{10} 1000 = 3$

$\rightarrow \text{antilog}_{10} 3 = 10^3 = 1000$

FIRST ORDER – GAS PHASE RXⁿ

- Let us consider a typical first order gas phase reaction

$$A(g) \rightarrow B(g) + C(g)$$

	A(g)	→	B(g)	+	C(g)
At $t = 0$	p_i atm		0 atm		0 atm
At time t	$(p_i - x)$ atm		x atm		x atm

$$\log\left(\frac{[R_0]}{[R_t]}\right) = \frac{kt}{2.303} \Longrightarrow \log\frac{(P_i)}{(P_i - x)} = \frac{kt}{2.303} \quad (1)$$

At time 't'

$$P_t = P_A + P_B + P_C$$

$$= (P_i - x) + x + x$$

$$P_t = P_i + x$$

$$P_t - P_i = x \quad (2)$$

Put value of (2) in (1)

$$\log\frac{(P_i)}{(P_i - P_t + P_i)} = \frac{kt}{2.303}$$

$$\log\frac{(P_i)}{(2P_i - P_t)} = \frac{kt}{2.303}$$

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HALF LIFE OF A FIRST ORDER REACTION

The half-life of a reaction is the time in which the concentration of a reactant is reduced to one half of its initial concentration. It is represented as $t_{1/2}$ or $t_{50\%}$.

$\rightarrow \ln\left(\frac{[R_0]}{[R_t]}\right) = kt$ — (1)

$\rightarrow t = t_{1/2}$ or $t_{50\%}$ — (2)

$\rightarrow [R_t] = \frac{[R_0]}{2}$ — (3)

Put value of (2) & (3) in (1)

$\ln \frac{[\cancel{R_0}]}{\left[\frac{\cancel{R_0}}{2}\right]} = k t_{50\%}$

$\boxed{\frac{\ln(2)}{k} = t_{50\%}}$ **

$\Rightarrow \frac{2.303 \log_{10}(2)}{k} = t_{50\%} \Rightarrow \frac{2.303 \times .3010}{k} = t_{50\%}$

$\boxed{\frac{.693}{k} = t_{50\%}}$ **

GIVE A THOUGHT

The half-life of a first order reaction depends on initial concentration of reactant.

A. YES

B. NO

$$t_{50\%} = \frac{.693}{K}$$

Zero order

$$t_{1/2} \text{ or } t_{50\%} = \frac{[R_0]}{2K}$$

GIVE A THOUGHT

Can a first-order reaction completes in finite time or concentration of reactant becomes zero in finite time?

A. YES

B. NO ✓

initial concentration can't be zero

$[R_t] = [R_0]e^{-kt}$

if rxn completes, $[R_t] = 0$, How?

$\Rightarrow t = \infty$

$\Rightarrow e^{-\infty} \Rightarrow \frac{1}{e^{\infty}} = \frac{1}{\infty} = '0'$

$\Rightarrow [R_0] \times 0$

$\Rightarrow 0$

Conclusion

→ At infinite time, first order rxn completes & $[R_t] = 0$

→ (It cannot be completed in a finite time)

COMPARISON BETWEEN FIRST AND ZERO ORDER

[Zero Order]

$$100\ M \xrightarrow[t_{1/2}]{} \underset{(50\%\ \text{rxn complete})}{50\ M} \xrightarrow[t_{1/2}]{} \underset{(100\%\ \text{rxn complete})}{0}$$

[First order]

$$100\ M \xrightarrow[t_{1/2}]{} \underset{(50\%\ \text{rxn complete})}{50\ M} \xrightarrow[t_{1/2}]{} \underset{(75\%\ \text{rxn complete})}{25\ M} \xrightarrow[t_{1/2}]{} \underset{(87.5\%\ \text{rxn complete})}{12.5\ M} \xrightarrow[t_{1/2}]{} \underset{(93.75\%\ \text{rxn complete})}{6.25\ M} \xrightarrow{t_{1/2}} \ldots\ldots 3.125\ M$$

LET'S PRACTICE!

$\log_{10} 10^2 = 2\log_{10} 10$
$= 2 \times 1$

Now calculate the same for $t_{99\%}$

$$2.303 \log_{10} \frac{[R_0]}{[R_t]} = kt \quad (1)$$

$t = t_{99.1}$ (2)

$[R_0] = 100\ M$

$[R_t] = 1\ M$ (3)

Put value of (2) & (3) in (1)

$$= 2.303 \log\left(\frac{100}{1}\right) = k t_{99.1}$$

$$= 2.303 \log 10^2 = k t_{99.1}$$

$$\frac{2.303 \times 2 \times 1}{k} = t_{99.1}$$

$$\boxed{\frac{4.606}{k} = t_{99.1}}$$

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LET'S PRACTICE

(NCERT WALLAH)

NCERT SOLVED EXAMPLE 3.7

A first order reaction is found to have a rate constant, $k = 5.5 \times 10^{-14}\ s^{-1}$. Find the half-life of the reaction.

$$t_{50\%} = \frac{.693}{k} = \frac{693 \times 10^{1}}{5.5 \times 10^{-14} \times 10^{3}}\ s$$

$$= \frac{693 \times 10^{12}}{55}\ s$$

$$= \frac{69.3 \times 10^{1} \times 10^{12}}{55}\ s$$

$$\approx \frac{\cancel{70} \times 10^{13}}{\cancel{55}}\ s \qquad \boxed{\approx 1.2 \times 10^{13}\ s}$$

NCERT SOLVED EXAMPLE 3.8

Show that in a first order reaction, time required for completion of 99.9% is 10 times of half-life ($t_{1/2}$) of the reaction.

$\rightarrow 2.303 \log \frac{[R_0]}{[R_t]} = kt$ — (1)

$\rightarrow t = t_{99.9\%}$ — (2)

$\rightarrow [R_0] = 100\ M$

$[R_t] = .1\ M$ — (3)

Put value of (2) & (3) in (1)

$2.303 \log \left(\frac{100 \times 10}{.1}\right) = k t_{99.9\%}$

$2.303 \log 10^3 = k t_{99.9\%}$

$2.303 \times 3 \times 1 = k t_{99.9\%}$

$6.909 = k t_{99.9\%}$

$\boxed{\frac{6.909}{k} = t_{99.9\%}}$

$t_{50\%} = \frac{.693}{k}$

$t_{50\%} \times 10 = \frac{6.93}{k}$

$\boxed{t_{50\%} \times 10 = t_{99.9\%}}$

H.P.

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CONCEPT POLISH - HOMEWORK

NCERT INTEXT 3.5

A first order reaction has a rate constant $1.15 \times 10^{-3}\ s^{-1}$. How long will 5 g of this reactant take to reduce to 3 g?

NCERT EXERCISE 3.17

The rate constant for a first order reaction is 60 s^{-1}. How much time will it take to reduce the initial concentration of the reactant to its $1/16^{th}$ value?

NCERT EXERCISE 3.19

A first order reaction takes 40 min for 30% decomposition. Calculate $t_{1/2}$.

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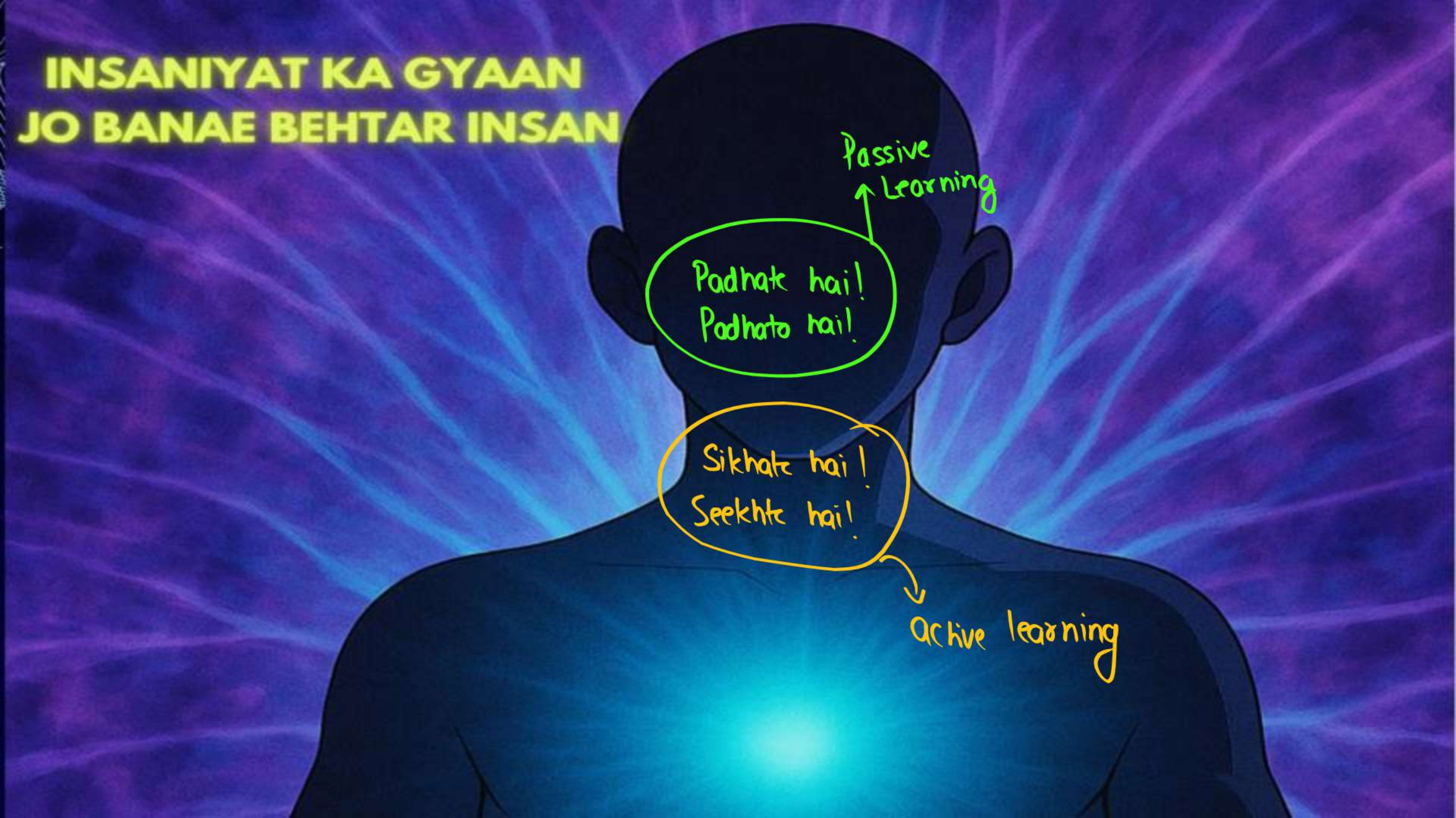

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CHEMISTRY

Lecture 07

CHEMICAL KINETICS

Numerical Practice, Pseudo First Order and Temperature Dependence on Rate of Reaction – Part I

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Numerical Practice

(ii) Pseudo First Order

(iii) Temperature Dependence on Rate of Reaction – Part I

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NUMERICAL PRACTICE

(Last Class H.W.)

NCERT INTEXT 3.5

A first order reaction has a rate constant $1.15 \times 10^{-3}\ s^{-1}$. How long will 5 g of this reactant take to reduce to 3 g?

k

$$2.303 \log\left(\frac{[R_0]}{[R_t]}\right) = kt$$

$[R_0] = 5\ g$

$[R_t] = 3\ g$

$\text{Log}_{10}5 = .6990$

$\text{Log}_{10}3 = .4771$

$$\Rightarrow 2.303 \log\left(\frac{5}{3}\right) = 1.15 \times 10^{-3} \times t$$

$$\Rightarrow 2.303 \times [\log(5) - \log(3)] = 1.15 \times 10^{-3} \times t$$

$$\Rightarrow 2.303 \times (.6990 - .4771) = 1.15 \times 10^{-3} \times t$$

$$\Rightarrow 2.303 \times (.2219) = 1.15 \times 10^{-3} \times t$$

$$\frac{2.303 \times .2219}{1.15 \times 10^{-3}} = t$$

$$\Rightarrow \frac{2303 \times 2219 \times 100}{115 \times 10^{4}} = t$$

$$\Rightarrow \frac{2303 \times 2219 \times 10^{-2}}{115} = t$$

$$\approx \frac{2300 \times 2200 \times 10^{-2}}{115 \ s^{-1}} = t$$

$$\approx \frac{\cancel{23} \times 22 \times \cancel{100}^{20}}{\cancel{115}_{\cancel{23}} \ s^{-1}} = t$$

$$\approx 440 \ s = t$$

NCERT EXERCISE 3.16

The rate constant for a first order reaction is 60 s^{-1}. How much time will it take to reduce the initial concentration of the reactant to its $1/16^{th}$ value?

$$[R_0] \xrightarrow[t=?]{} \left[\frac{R_0}{16}\right] = [R_t]$$

$$\Rightarrow 2.303 \log \frac{[R_0]}{[R_0/16]} = 60 \times t$$

$$\Rightarrow 2.303 \log(16) = 60 \times t$$

$$\Rightarrow 2.303 \log(2^4) = 60 \times t$$

$$\Rightarrow 2.303 \times 4 \log 2 = 60 \times t$$

$$\Rightarrow \frac{2.303 \times 4 \times .3010}{60} = t$$

$$\frac{2.303 \times 3010}{15 \times 10^4 \times 10^3}$$

$$\approx \frac{2300 \times 600}{3 \times 10^7}$$

$$\approx \frac{23 \times 2 \times 10^4}{10^3}$$

$$\approx \frac{46}{10^3} \quad \boxed{\approx .046\ s}$$

NCERT EXERCISE 3.19

A first order reaction takes 40 min for 30% decomposition. Calculate $t_{1/2}$.

$\log_{10}7 = .8451$

$[R_0] = 100\ g$

$[R_t] = 70\ g$

$\Rightarrow 2.303 \log\left(\frac{100}{70}\right) = k \times 40 \times 60\ s$

$\Rightarrow 2.303(\log 10 - \log 7) = k \times 40 \times 60\ s$

$\Rightarrow 2.303(1 - .8451) = k \times 40 \times 60\ s$

$\Rightarrow 2.303 \times .1549 = k \times 40 \times 60\ s$

$\Rightarrow \frac{2.303 \times .1549}{40 \times 60} = k$

$\frac{2.303 \times .1549}{40 \times 60 \times 10^3 \times 10^4} = k$

$\approx \frac{2301 \times 1600}{40 \times 60 \times 10^3 \times 10^4} = k$

$\approx \frac{767 \times 2}{107} = k$

$\approx \frac{1534}{107} = k$

$t_{1/2} = \frac{.693}{k}$ — (1)

$\approx \frac{.693 \times 10^7}{1534 \times 10^3}$

$\approx \frac{693 \times 10^4}{1534}$

$\approx \frac{6930 \times 10^3}{1534}$

$\approx 4.5 \times 10^3$

$\approx 4500\ s$

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'false'

PSEUDO FIRST ORDER REACTION

PSEUDO FIRST ORDER REACTION

A pseudo first order reaction is a reaction which is actually of higher order but behaves as first order reaction is called a pseudo first order reaction.

(ACIDIC HYDROLYSIS OF ESTER)

(I) (ethyl acetate) ← $CH_3COOC_2H_5 + H_2O \xrightarrow{H^+} CH_3COOH + C_2H_5OH$

	$CH_3COOC_2H_5$	H_2O (excess)	CH_3COOH	C_2H_5OH
$t = 0$	0.01 mol	10 mol	0 mol	0 mol
t	0 mol	9.99 mol	0.01 mol	0.01 mol

negligible change in conc. of reactant even after rxn so, we can say conc. of this reactant is constant. Hence, it does not affect rate of rxn.

Rate = $k[CH_3COOC_2H_5]^1$

Rate of rxn only depends on:

(ACIDIC HYDROLYSIS OF CANE SUGAR)

(II) Inversion of cane sugar is another pseudo first order reaction.

$$C_{12}H_{22}O_{11} + H_2O \xrightarrow{H^+} C_6H_{12}O_6 + C_6H_{12}O_6$$

(excess)

Cane sugar (Sucrose) — Glucose — Fructose

Rate = $k[C_{12}H_{22}O_{11}]^1$

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TEMPERATURE DEPENDENCE OF RATE OF REACTION – PART I

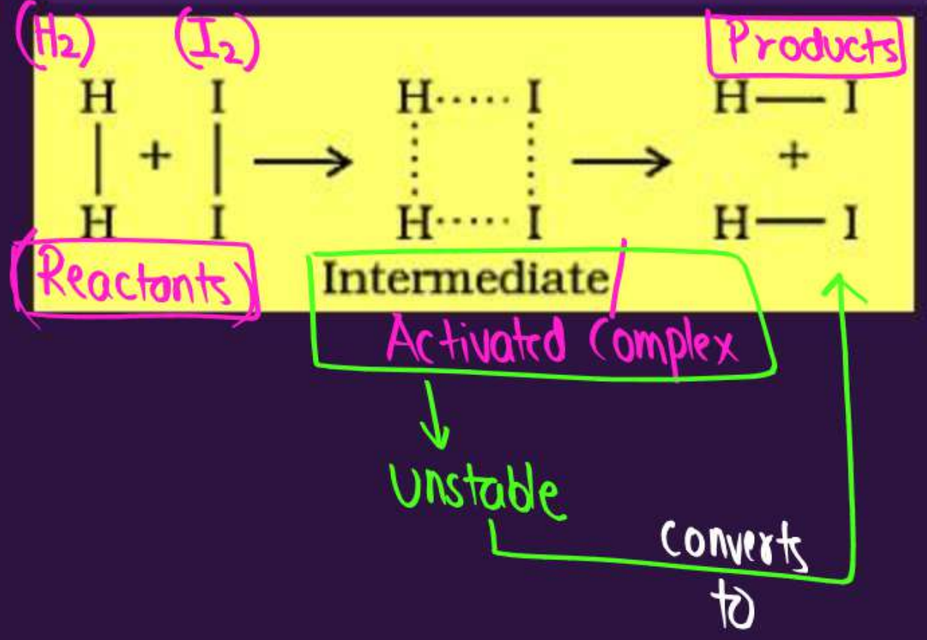
ACTIVATION ENERGY

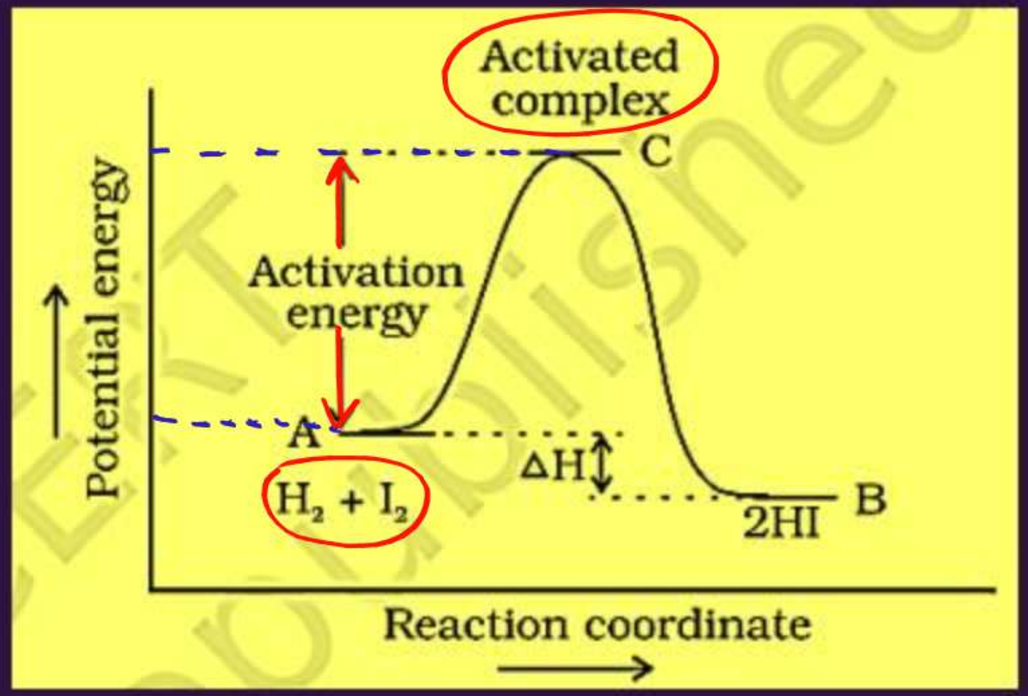
(Ea)

C-I

Activation Energy (Ea) is the minimum amount of energy that reactant molecules must have to successfully undergo a chemical reaction.

C-II





ARRHENIUS EQUATION

The temperature dependence of the rate of a chemical reaction can be accurately explained by Arrhenius equation

$$k = Ae^{-E_a/RT}$$

Symbol	Meaning
k	Rate constant
A	Frequency factor (collision frequency + orientation factor)
E_a	Activation energy (J/mol)
R	Gas constant (8.314 J/mol·K)
T	Temperature in Kelvin
e	Euler's number (≈ 2.718)

ARRHENIUS EQUATION

$$k = A e^{-E_a/RT}$$

(i) $e^{-E_a/RT}$: Exponential Factor

- Yeh batata hai ki kitne molecules ke paas enough energy ($\geq$ Ea) hai.
- Ab T badhta hai, E_a/RT chhota hota hai, _____ aur bada hota hai → **Rate constant k badhta hai**

It has been found that for a chemical reaction with rise in temperature by 10 °C, the rate constant is nearly doubled.

ARRHENIUS EQUATION

$$k = Ae^{-E_a/RT}$$

(ii) A: Frequency Factor

'A' btata hai ki kitne molecules collide kr rhe hai with correct orientation.

$$k = A \cdot e^{-\frac{E_a}{RT}} = (\text{Collisions with correct orientation}) \times (\text{Molecules with sufficient energy})$$

IMPORTANT TO REMEMBER

$$k = Ae^{-Ea/RT}$$

In syllabus

Factor	Effect on Rate Constant (k)
T ↑	k ↑
Ea ↑	k ↓
A ↑	k ↑ (but depends on orientation)

GIVE A THOUGHT

Rate constant (k) depends on T, E_a and A.

A. YES

B. NO

ARRHENIUS EQUATION – NATURAL LOG FORM

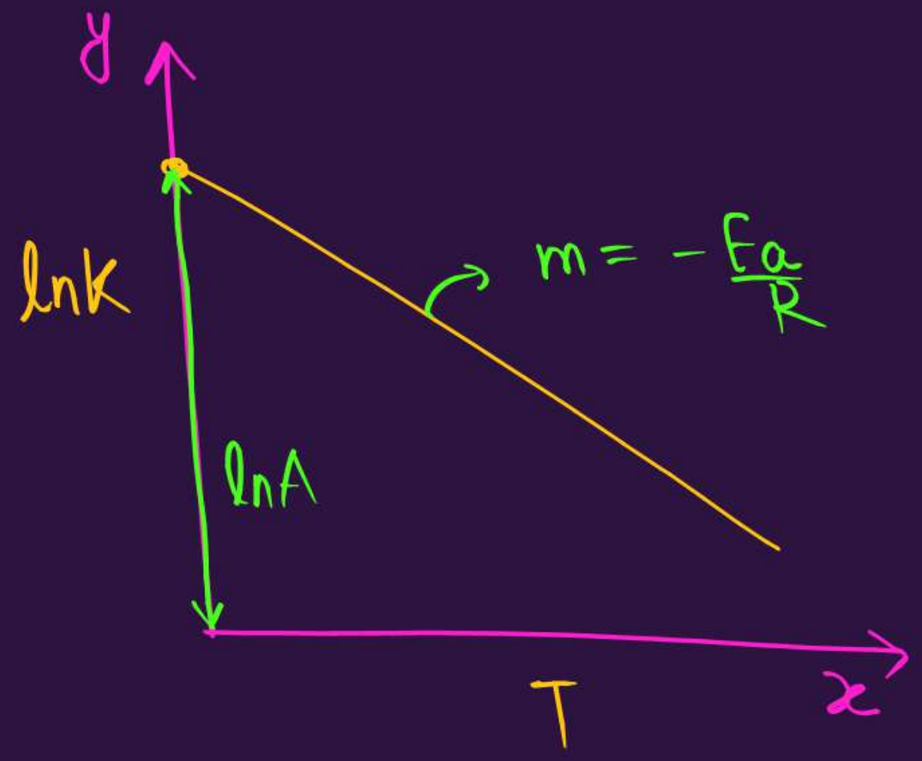
$$k = Ae^{-E_a/RT}$$

$$\Rightarrow \ln k = \ln A + \ln e^{-E_a/RT}$$

$$\Rightarrow \ln k = \ln A - \frac{E_a}{RT}\cancel{\ln e}^{\,1}$$

$$\Rightarrow \ln k = \ln A - \frac{E_a}{RT}$$

$$\boxed{\underbrace{\ln k}_{y} = \underbrace{-\frac{E_a}{RT}}_{m\,x} + \underbrace{\ln A}_{c}}$$



$$\ln K = -\frac{E_a}{RT} + \ln A$$

$$2.303 \log K = -\frac{E_a}{RT} + 2.303 \log A$$

Divide entire eqn by $\boxed{2.303}$

$$\boxed{\log K = \log A - \frac{E_a}{2.303\,RT}}$$

Arrhenius Eqn → Common Log form

ARRHENIUS EQUATION – COMMON LOG FORM

k_1, T_1
k_2, T_2

$$k = Ae^{-E_a/RT}$$

$$\ln x = 2.303 \log_{10} x$$

$$\ln k_1 = -\frac{E_a}{RT_1} + \ln A \quad -(1)$$

$$\ln k_2 = -\frac{E_a}{RT_2} + \ln A \quad -(2)$$

(2) − (1)

$$\ln k_2 - \ln k_1 = -\frac{E_a}{RT_2} + \frac{E_a}{RT_1}$$

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

$$\Rightarrow 2.303 \log \frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

$$\Rightarrow \boxed{\log \frac{k_2}{k_1} = \frac{E_a}{2.303R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]}$$

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CONCEPT POLISH - HOMEWORK

NCERT SOLVED EXAMPLE 3.9

The rate constants of a reaction at 500K and 700K are $0.02s^{-1}$ and $0.07s^{-1}$ respectively. Calculate the values of E_a and A.

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Thank You

PARISHRAM

2026

CHEMISTRY

Bharat Mata Ki Jai ♡

Lecture 08

CHEMICAL KINETICS

Temperature Dependence on Rate of Reaction – Part II, Effect of Catalyst and CBSE PYQs'

BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Temperature Dependence on Rate of Reaction – Part II (✓)

(ii) Collision Theory of Chemical Rxn (✓)

(iii) Effect of catalyst on R.o.R. (✓)

(iv) CBSE PYQs' (✓)

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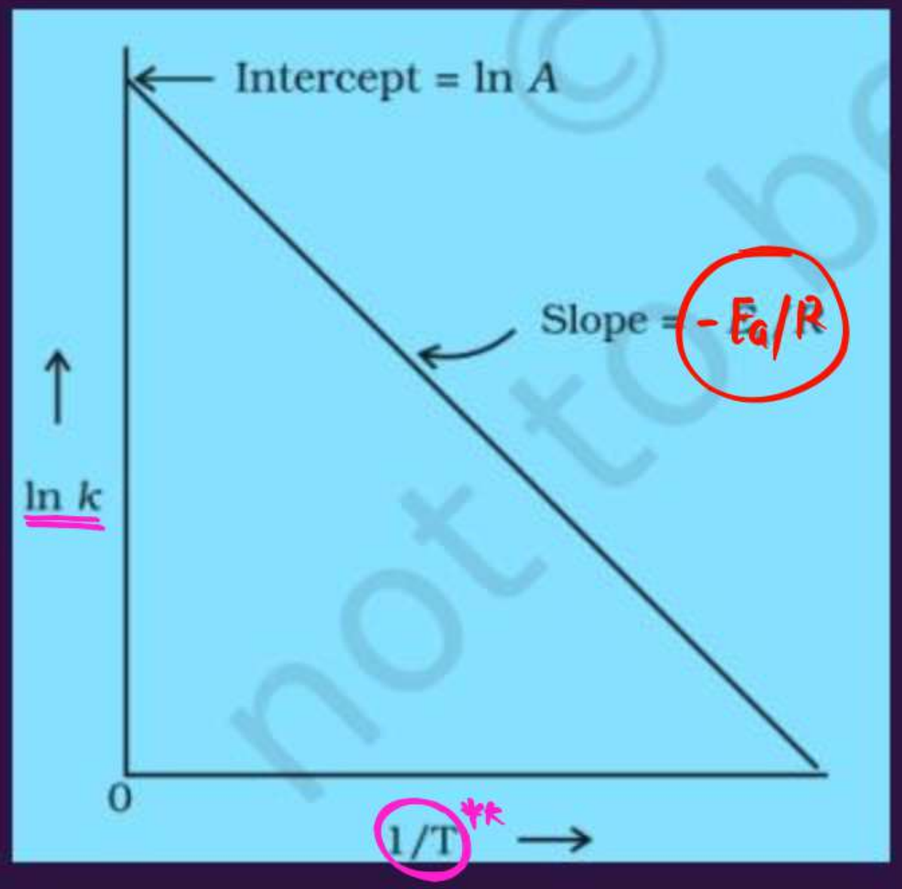
Excellence deserves recognition!

In the upcoming week, the **Top 10 Students** on the Weekly Test leaderboard will receive exclusive giveaway books.

NOTE: You will be eligible to receive the books only if you attempt the test in real time.

IMPORTANT CHANGE

$$\ln k = -\frac{E_a}{R}\frac{1}{T} + \ln A$$



CONCEPT POLISH – HOMEWORK DISCUSSION

NCERT SOLVED EXAMPLE 3.9

very complex calculation

In Exam → easy

The rate constants of a reaction at 500K and 700K are $0.02 s^{-1}$ and $0.07 s^{-1}$ respectively. Calculate the values of E_a and A.

$R = 8.314 \text{ J mol}^{-1}\text{K}^{-1}$ | $\log 2 = 0.3010$ | $\log 7 = 0.8450$ (approx)

$T_1 = 500 K$, $K_1 = .02 s^{-1}$

$T_2 = 700 K$, $K_2 = .07 s^{-1}$

$$\log\left(\frac{K_2}{K_1}\right) = \frac{E_a}{2.303R}\left(\frac{T_2 - T_1}{T_1 T_2}\right)$$

$$\Rightarrow \log\left(\frac{.07}{.02}\right) = \frac{E_a}{2.303 \times 8.314}\left(\frac{700-500}{700 \times 500}\right)$$

$$\Rightarrow \log\left(\frac{7}{2}\right) = \frac{E_a}{2.303 \times 8.314}\left(\frac{200}{350000}\right)$$

1750

$$\Rightarrow \log(7) - \log(2) = \frac{E_a}{2.303 \times 8.314}\left(\frac{1}{1750}\right)$$

$$\Rightarrow .8450 - .3010 = \frac{E_a}{2.303 \times 8.314 \times 1750}$$

$$\Rightarrow \frac{544 \times 2.303 \times 8.314 \times 175}{100 \times 1000 \times 1000} = E_a$$

$$\Rightarrow \frac{544 \times 2303 \times 8314 \times 175}{100000000} = E_a$$

$$\approx \frac{5\cancel{00} \times 23\cancel{00} \times 83\cancel{00} \times 2\cancel{00}}{1\cancel{000000000}} = E_a$$

$$\approx 5 \times 2 \times 23 \times 83 = E_a$$

$$\approx 10 \times 23 \times 83 = E_a$$

$$\approx \boxed{19090 \ \frac{J}{mol} = E_a}$$

Just Try

$$\boxed{\log K = \log A - \frac{E_a}{2.303 RT}}$$

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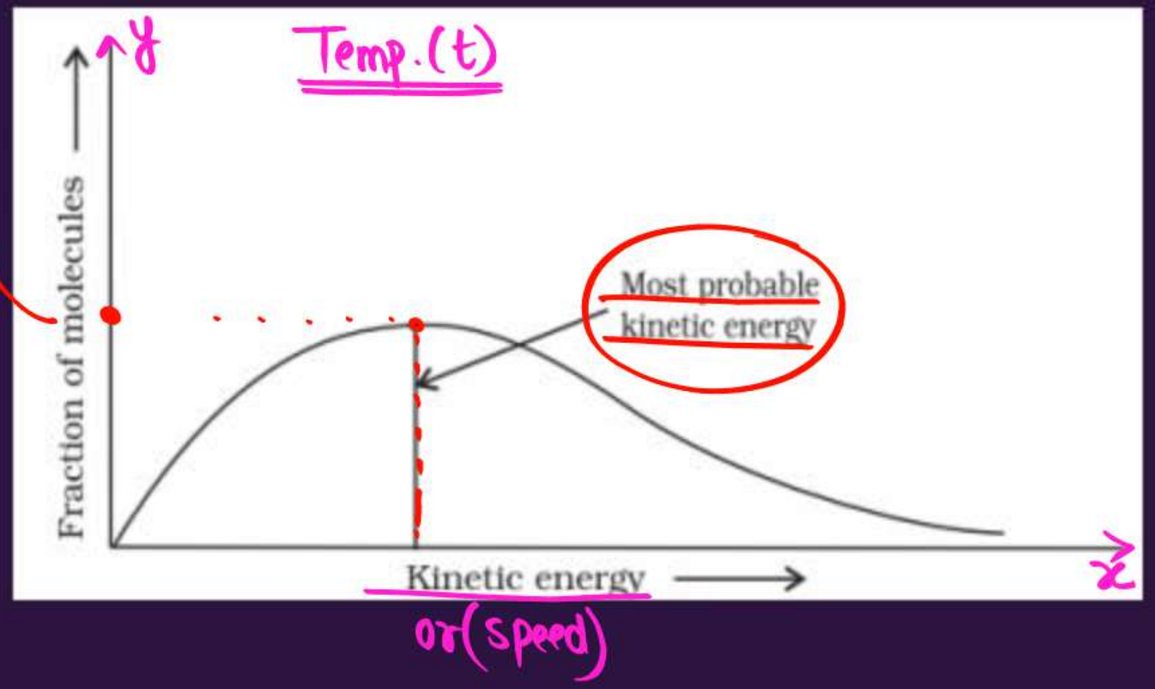
TEMPERATURE DEPENDENCE OF RATE OF REACTION – PART II

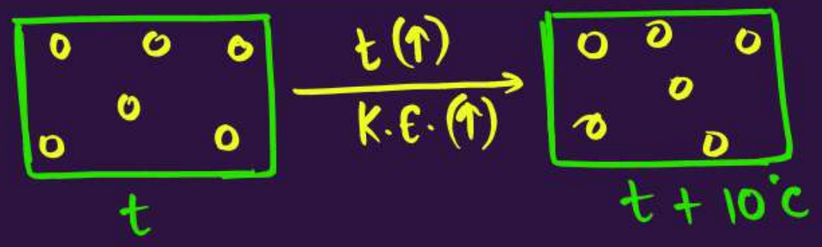
MAXWELL-BOLTZMANN CURVE

$\left(\frac{N_E}{N_T}\right)$ N_E → molecules with energy (E) kinetic
N_T → Total no. of molecules

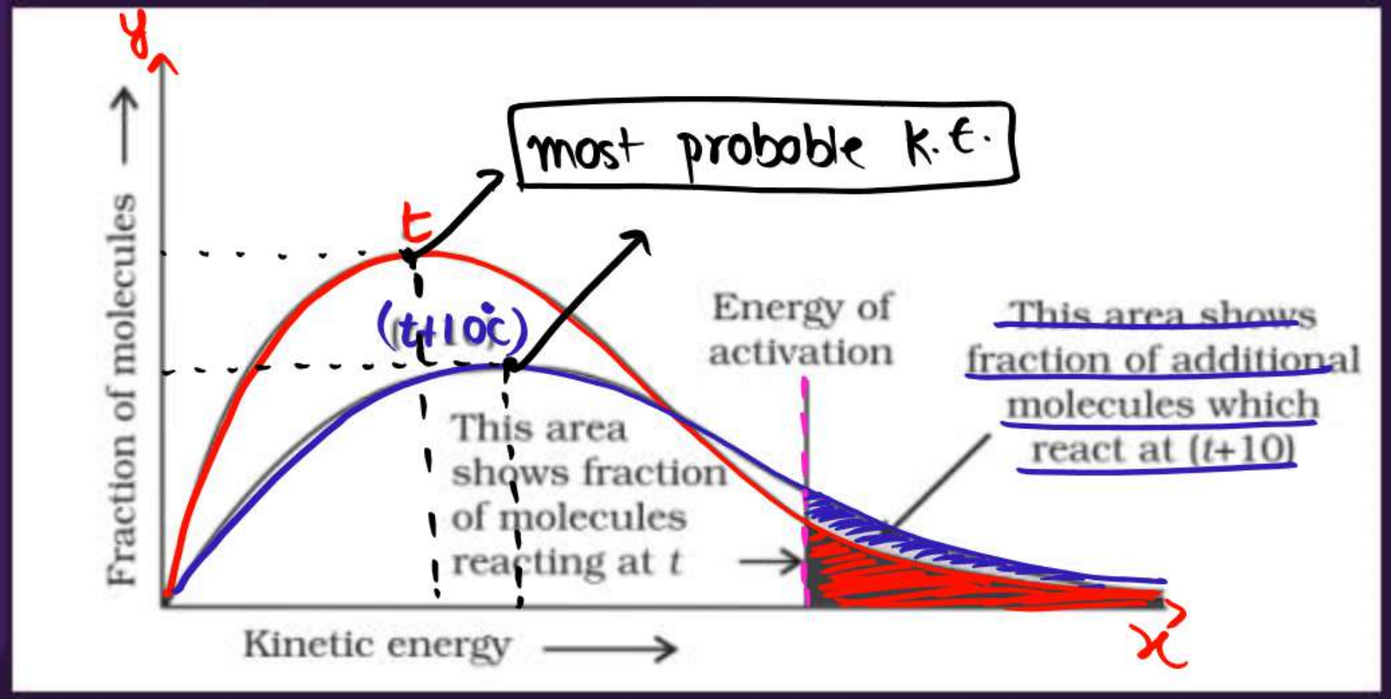
- Temperature is directly proportional to average kinetic energy of the molecules.
- **Shows distribution of kinetic energy** among molecules at a given temperature. (t)
- **X-axis** → Kinetic Energy (or speed)
  Y-axis → ~~Number~~ Fraction of molecules (with that energy)
- **Most probable energy**: Peak point of the curve (maximum molecules have this energy).



MAXWELL-BOLTZMANN CURVE



- Temperature badhne par, zyada molecules ke paas activation energy se zyada energy hoti hai.
- At (t + 10°), the area showing the fraction of molecules having energy equal to or greater than activation energy gets doubled leading to doubling the rate of a reaction.



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COLLISION THEORY OF CHEMICAL REACTIONS

COLLISION THEORY

According to this theory:

- Reactant molecules are assumed to be hard spheres. A + B C
- Reactant molecules must collide to form products.
- For a bimolecular elementary reaction A + B → Products

$$\text{Rate} = Z_{AB} e^{-E_a / RT}$$

(only 1 parameter, i.e. energy of reactants ≥ Ea)

The number of collisions per second per unit volume of the reaction mixture is known as collision frequency (Z).

- The above expression gives accurate results for simple molecules/atomic species.

COLLISION THEORY

- For complex molecules, all collisions do not lead to the formation of products.
- The collisions in which molecules collide with sufficient kinetic energy (called threshold energy) and proper orientation, so as to facilitate breaking of bonds between reacting species and formation of new bonds to form products are called as effective collisions.

$$\text{Rate} = PZ_{AB}e^{-E_a/RT}$$

energy of reactants $\geq E_a$

orientation of reactant molecules

orientation factor

collision frequency

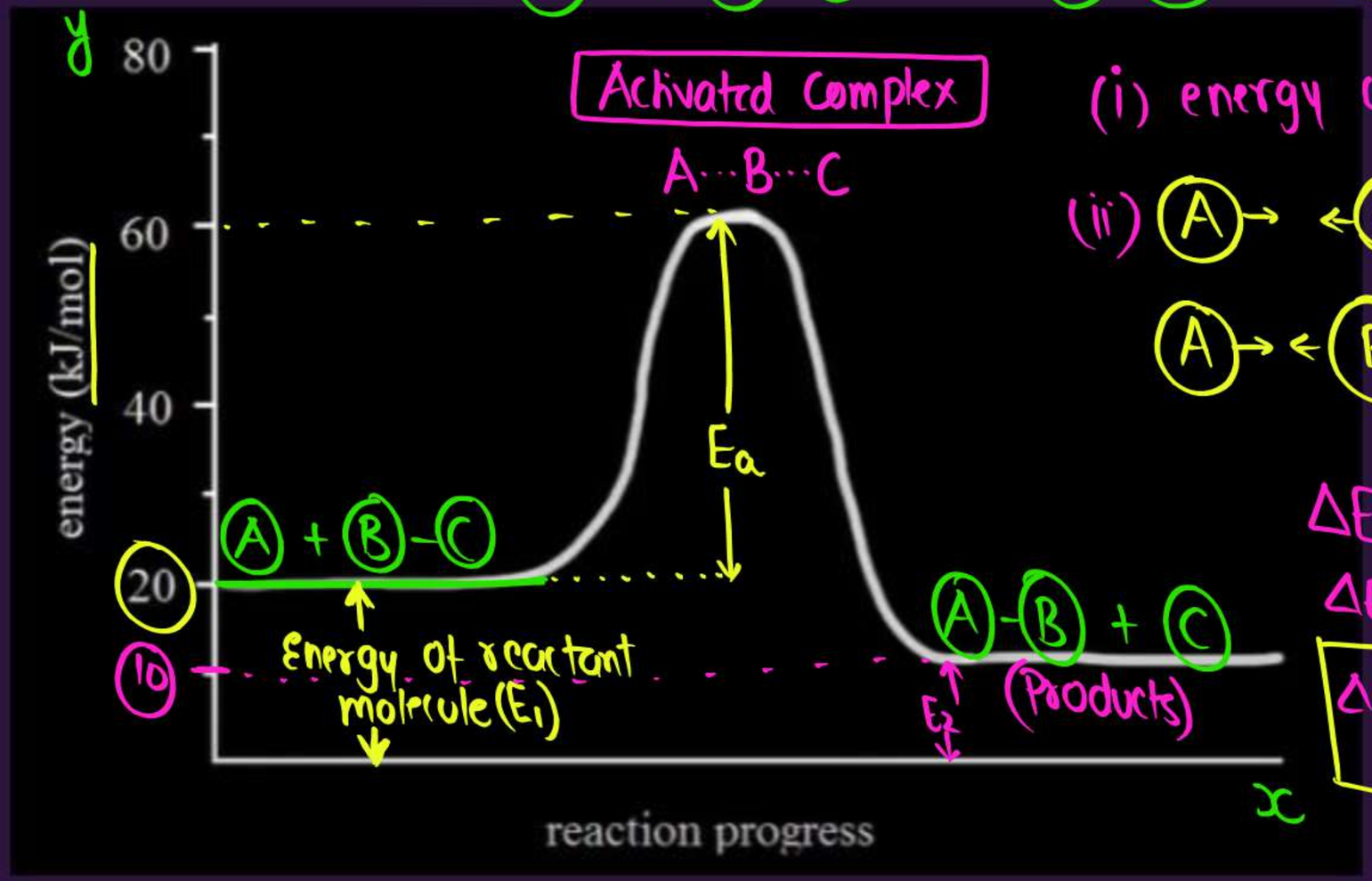
In Arrhenius eqn

$A = PZ$

COLLISION THEORY

(Threshold energy = $E_a + E_1$)

$A + B-C \longrightarrow A-B + C$



(i) energy of reactants $\geq E_a$

(ii) A→ ←C—B (✗)

A→ ←B—C (✓)

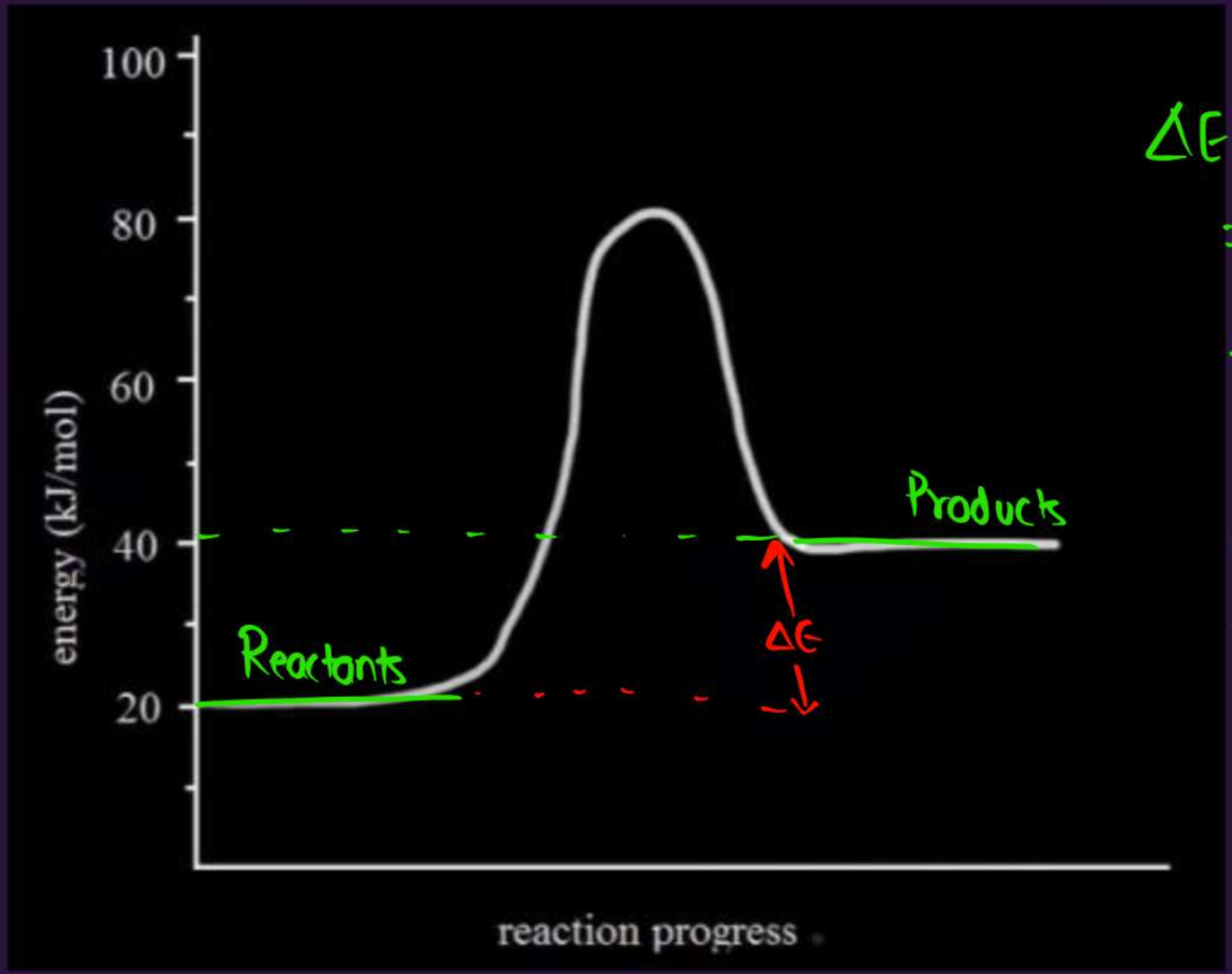
$\Delta E = E_2 - E_1$

$\Delta E = 10 - 20$

$\Delta E = -10 \ \frac{kJ}{mol}$

Exothermic Rxn

COLLISION THEORY



$\Delta E = E_2 - E_1$

$= 40 - 20$

$= +20 \ \frac{kJ}{mol}$

Endothermic

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EFFECT OF CATALYST ON RATE OF REACTION

EFFECT OF CATALYST

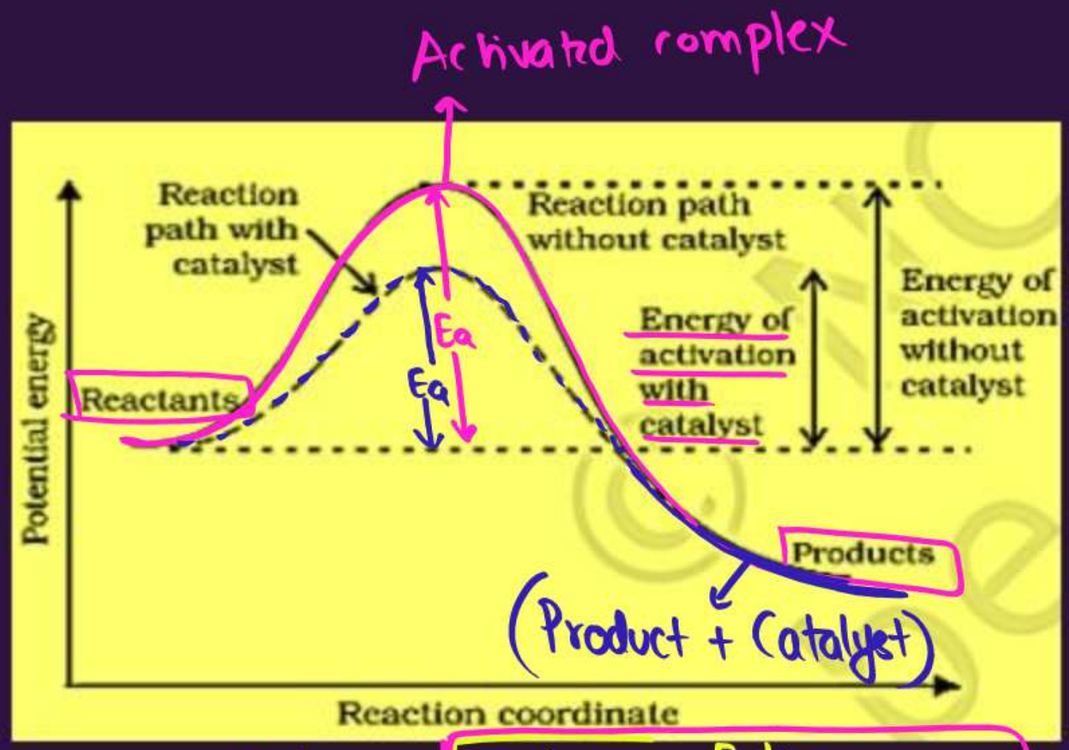
Inhibitor → reduces rate of rxn

- A catalyst increases the rate of a reaction without itself undergoing permanent chemical change.

$$2KClO_3 \xrightarrow[\text{heat}]{MnO_2} 2\,KCl + 3O_2$$

- Intermediate Complex Theory: Catalyst forms a temporary complex with reactants → decomposes to give product + catalyst recovered.

- Catalyst provides an alternate path with lower activation energy, as per Arrhenius equation ⇒ faster reaction.



$K = Ae^{-Ea/RT}$, $E_a(\downarrow)$, $e^{-Ea/RT}(\uparrow)$, $K(\uparrow)$

GIVE A THOUGHT

A catalyst does not alter Gibbs energy, ΔG of a reaction.

A. YES

B. NO

GIVE A THOUGHT

Catalyst

It catalyses the spontaneous reactions but does not catalyse non-spontaneous reactions.

A. YES

B. NO

GIVE A THOUGHT

A catalyst does not change the equilibrium constant of a reaction rather, it helps in attaining the equilibrium faster.

A. YES

B. NO

Does catalyst change potential energy barrier or Ea of reactants?

(A) Yes

(B) No

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CBSE PREVIOUS YEAR QUESTIONS

CBSE 2025

Which among the following is a false statement ?

(A) Rate of zero order reaction is independent of initial concentration of reactant. (✓)

Rate = K

(B) Half-life of a zero order reaction is inversely proportional to the rate constant. (✓)

$t_{1/2} = \frac{[R_0]}{2K}$

(C) Molecularity of a reaction may be zero.

(D) For a first order reaction, $t_{1/2} = 0.693/k$.

CBSE 2025

The number of molecules that react with each other in an elementary reaction is a measure of the :

(A) activation energy of the reaction
(B) stoichiometry of the reaction
(C) molecularity of the reaction
(D) order of the reaction

(comment section)

CBSE 2025

$t_{1/2} = \frac{.693}{k}$

Assertion (A) : In a first order reaction, if the concentration of the reactant is doubled, its half-life is also doubled. (x)

Reason (R) : The half-life of a reaction does not depend upon the initial concentration of the reactant in a first order reaction. (✓)

(A) Both Assertion (A) and Reason (R) are true and Reason (R) is the correct explanation of the Assertion (A).

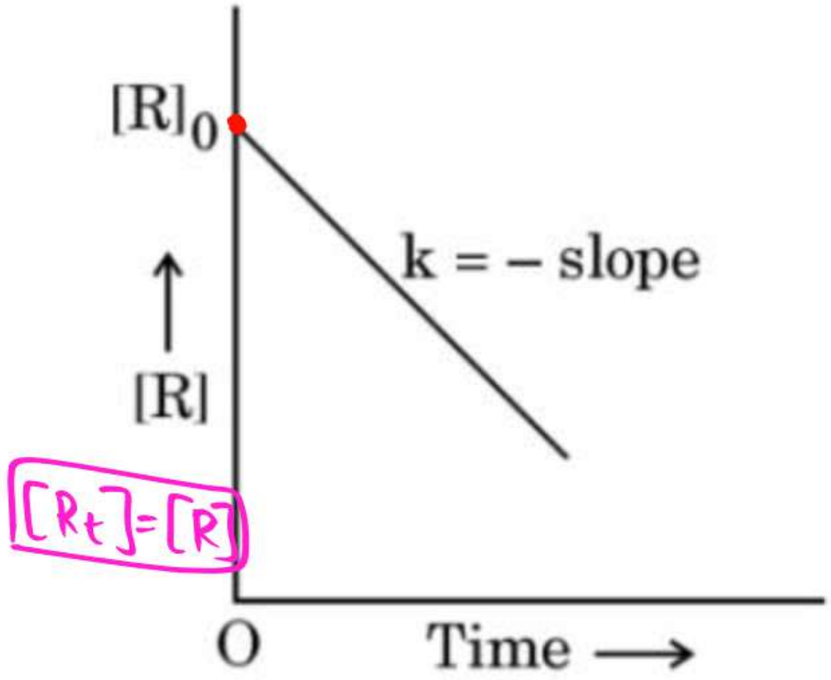
(B) Both Assertion (A) and Reason (R) are true, but Reason (R) is ***not*** the correct explanation of the Assertion (A).

(C) Assertion (A) is true, but Reason (R) is false.

(D) Assertion (A) is false, but Reason (R) is true.

CBSE 2025

A plot between concentration of reactant [R] and time 't' is shown below. Which of the given order of reaction is indicated by the graph ?



$[R_t] = -kt + [R_0]$

$y \quad m \quad x \quad c$

$[R_t] = [R]$

(A) Third order

(B) Second order

(C) First order

(D) Zero order

CBSE 2025

Which of the following is the slope of the first order reaction in the plot of $ln[R]$ vs. time ?

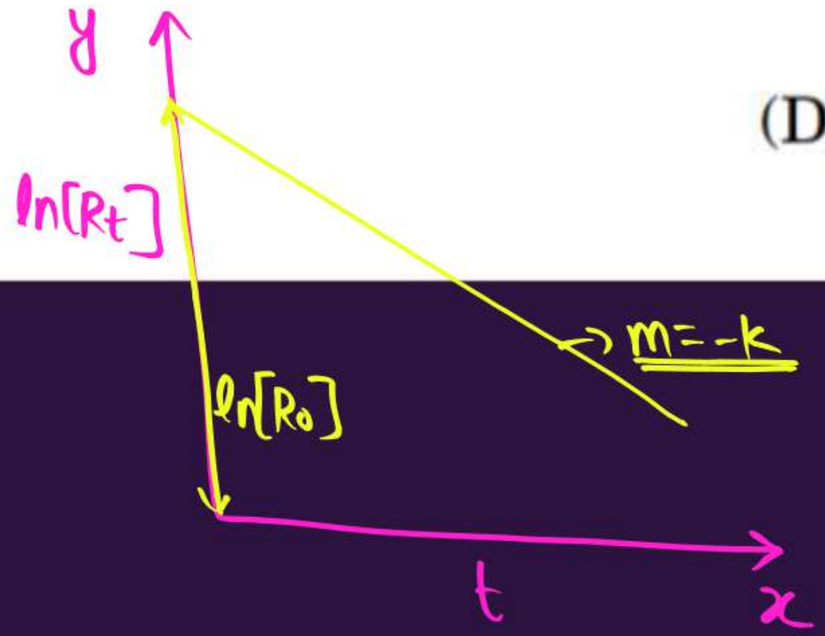
$$\underbrace{\ln[R_t]}_{y} = \underset{m}{-k}\,\underset{x}{t} + \underset{c}{\ln[R_0]}$$

(A) $+k$

(B) $\dfrac{+k}{2.303}$

(C) $-k$

(D) $\dfrac{-k}{2.303}$



CBSE 2025

The activation energy (E_a) of a reaction can be determined from the slope of which of the following plots ?

(A) $ln\ k$ vs. T

(B) $\frac{ln\ k}{T}$ vs. T

(C) $ln\ k$ vs. $\frac{1}{T}$

(D) $\frac{T}{ln\ k}$ vs. $\frac{1}{T}$

CBSE 2025

Which of the following represents the fraction of molecules with energies equal to or greater than E_a ?

(A) $\frac{-E_a}{RT}$

(B) $e^{-E_a/RT}$

(C) $e^{+E_a/RT}$

(D) $\frac{+E_a}{RT}$

$K = A e^{-Ea/RT}$

A → frequency factor

$e^{-Ea/RT}$ → Exponential factor

CBSE 2025

A certain reaction is 50% complete in 20 minutes at 300 K and the same reaction is 50% complete in 5 minutes at 350 K. Calculate the activation energy if it is a first order reaction.

[R = 8.314 J K^{-1} mol^{-1}; log 4 = 0.602]

$$\log\left(\frac{K_2}{K_1}\right) = \frac{E_a}{2.303R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

CONCEPT POLISH - HOMEWORK

Slide 29 & 35

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→ Marks don't matter in life.

→ Knowledge, skills & intent → MATTER

Thank You