ULTIMATE KCET

CRASH COURSE 2026

Chemistry | Lecture - 01

Chemical kinetics – II + Thermodynamics

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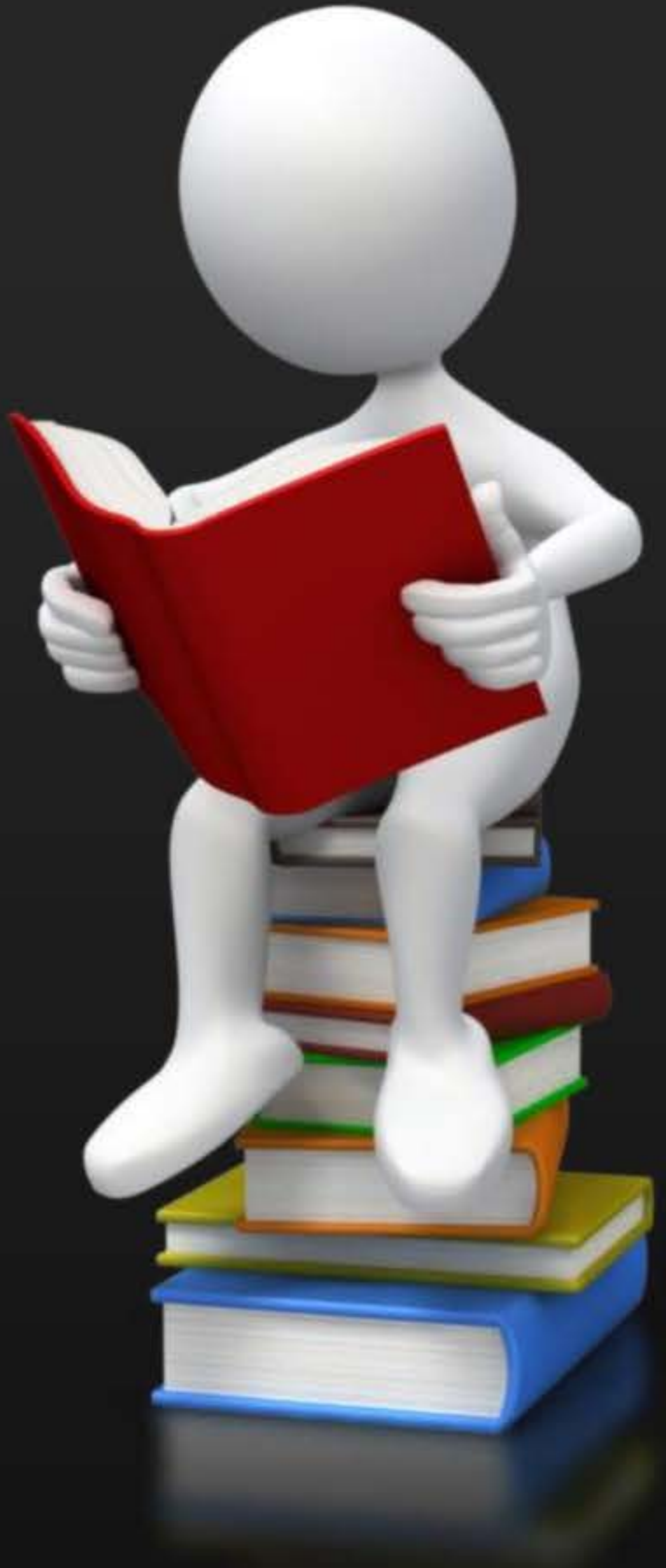
Recap *of previous lecture*

① **Batteries and Mcqs on redox reaction + Chemical kinetics**

Topics *to be covered*

1. Chemical kinetics - part II ✓ + MCQ ✓
2. Thermodynamics – synopsis ✓ + MCQ ✓



Integrated rate expression.

first order →

$$K = \frac{2.303}{t}\log\frac{[R]_0}{[R]}$$ ✓

or

or

$$K = \frac{2.303}{t}\log\frac{a}{(a-x)}$$

or

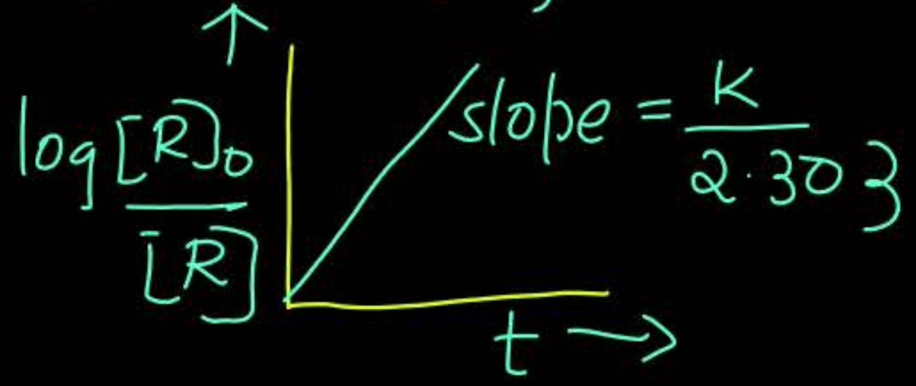
$$k = \frac{2.303}{t}\log\frac{c_0}{c_t}$$

$$\ln[R] = -kt + \ln[R]_0$$

$$kt = \ln[R]_0 - \ln[P]$$

$$kt = 2.303\log\frac{[R]_0}{[R]}$$

$$\frac{kt}{2.303} = \log\frac{[R]_0}{[R]}$$

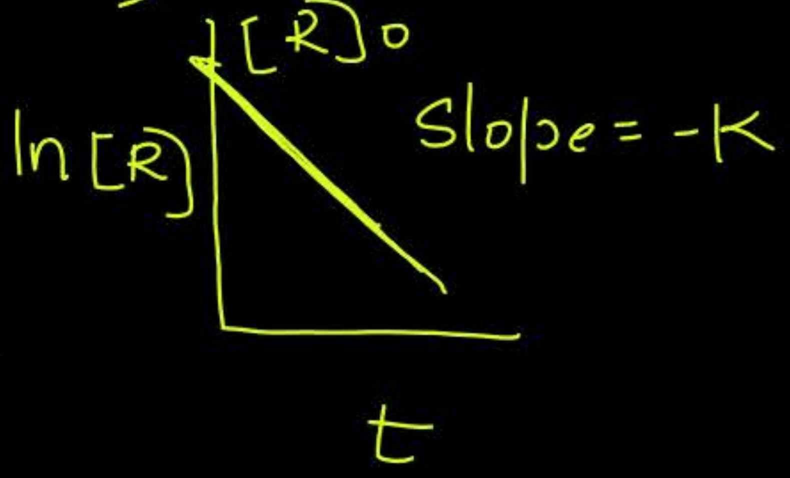


$$\ln[R] = -kt + \ln[R]_0$$

$$\ln[R] - \ln[R]_0 = -kt$$

$$\frac{\ln[R]}{\ln[R]_0} = -kt$$

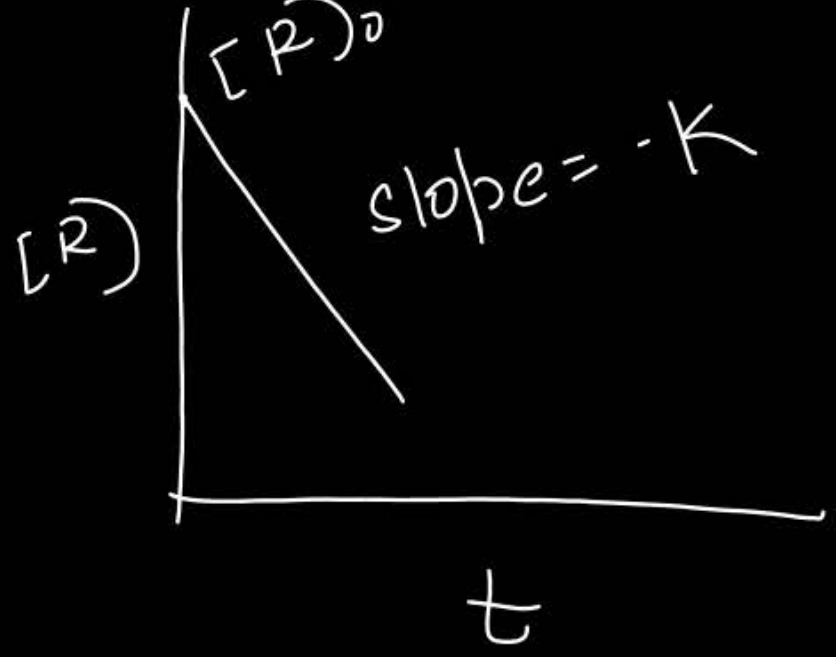
$$\frac{[R]}{[R]_0} = e^{-kt}$$ ✓



$$t_{1/2} = \frac{0.693}{K}$$

$$k = \frac{[R]_0 - [R]}{t}$$

$$t_{1/2} = \frac{[R]_0}{2K}$$



→ zero order

Order	Zero zero	1st 1st	2nd 2nd	nth nth.
Differential rate law	Rate = k	Rate= k[A]	Rate = $k[A]^2$	Rate = $k[A]^n$
Integrated rate law	$[A_0]-[A]_t = kt$	$kt = \ln\frac{[A]_0}{[A]_t}$	$kt = \frac{1}{[A]_t} - \frac{1}{[A]_0}$	$kt = \frac{1}{(n-1)}\left[\frac{1}{[A]_t^{n-1}} - \frac{1}{[A]_0^{n-1}}\right]$
Half life ($t_{1/2}$)	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{\ln 2}{k}$	$t_{1/2} = \frac{1}{[A]_0 k}$	$t_{1/2} = \cancel{\frac{1}{k(n-1)}}\left[\frac{\cancel{2^{n-1}-1}}{[A_0]^{n-1}}\right]$
($t_{3/4}$)	$t_{3/4}=1.5\ t_{1/2}$	$t_{3/4}= 2\ t_{1/2}$	$t_{3/4} = 3\ t_{1/2}$	$\boxed{t_{3/4}= (2^{n-1}+1)\ t_{1/2}}$

$t_{1/2} \propto \frac{1}{[A]_0^{n-1}}$

$A \longrightarrow B.$

Order	Zero zero	1st first	2nd 2^{nd}	nth n^{th}.
Differential rate law	Rate = k $k = mol\,L^{-1}\,s^{-1}$	Rate= k[A] $k = \frac{Rate}{[A]} = s^{-1}$	Rate = $k[A]^2$ $k = L\,mol^{-1}\,s^{-1}$	Rate = $k[A]^n$ $k = \frac{[Rate]}{[A]^n} \rightarrow \frac{mol\,L^{-1}\,s^{-1}}{mol^n\,L^{-n}} = mol^{1-n}\,L^{n-1}\,s^{-1}$
Integrated rate law	$[A_0]-[A]_t = kt$ $k = \frac{[A]_0 - [A]_t}{t}$	$kt = \ln\frac{[A]_0}{[A]_t}$ $\frac{kt}{2.303} = \log\frac{[A]_0}{[A_t]}$	Extra $kt = \frac{1}{[A]_t} - \frac{1}{[A]_0}$	$kt = \frac{1}{(n-1)}\left[\frac{1}{[A]_t^{n-1}} - \frac{1}{[A]_0^{n-1}}\right]$
Half life $(t_{1/2})$	$t_{1/2} \rightarrow = \frac{[A]_0}{2k}$ ✓	$t_{1/2} = \frac{0.693}{K}$. $t_{1/2} = \frac{\ln 2}{k}$	$t_{1/2} = \frac{1}{[A]_0 k}$	✓ [3000 KCET] $t_{1/2} = \frac{1}{k(\underset{2}{n}-1)}\left[\frac{2^{n-1}-1}{[A_0]^{n-1}}\right]$
$(t_{3/4})$	$t_{3/4} = 1.5\ t_{1/2}$	$t_{3/4} = 2\ t_{1/2}$	$t_{3/4} = 3\ t_{1/2}$	$\boxed{t_{3/4} = (2^{n-1}+1)\ t_{1/2}}$

$A \longrightarrow B$

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

$mol\,L^{-1}\,s^{-1}$

$mol\,dm^{-3}\,s^{-1}$

$$\ln 2 = 2.303 \log 2 = 2.303 \times 0.3$$
$$= 0.693$$

First Order gaseous phase Reactions

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

where, p_i is the initial pressure at time $t = 0$.

$$p_t = (p_i - x) + x + x = p_i + x$$

$$x = (p_t - p_i)$$

where, $\boxed{p_A = p_i - x = p_i - (p_t - p_i)}$

$$= 2p_i - p_t$$

$$\boxed{k = \left(\frac{2.303}{t}\right)\left(\log\frac{p_i}{p_A}\right)}$$

$$\boxed{= \frac{2.303}{t}\log\frac{p_i}{(2p_i - p_t)}}$$

Question

$$A \longrightarrow B + C$$

$$\begin{array}{ccc} p_i^o & 0 & 0 \\ (P_i^o - x) & x & x \end{array}$$

$P_t = P_i^o - \cancel{x} + \cancel{x} + x$

$P_t = P_i^o + x$

✓ $x = P_t - P_i^o$

Consider a first order gas phase decomposition reaction given below:

$\boxed{A(g) \rightarrow B(g) + C(g)}$

The initial pressure of the system before decomposition of A was p_i. After lapse of time 't', total pressure of the system increased by x units and became 'p_t'. The rate constant k for the reaction is given as given as ....

A $k = \dfrac{2.303}{t}\log\dfrac{P_i}{P_i - x}$

B $\boxed{k = \dfrac{2.303}{t}\log\dfrac{P_i}{2P_i - P_t}}$ ✓

C $k = \dfrac{2.303}{t}\log\dfrac{P_i}{2P_i + P_t}$

D $k = \dfrac{2.303}{t}\log\dfrac{P_i}{2P_i + P_t}$

Pseudo first order reactions:-

→ Inversion of cane sugar is another pseudo first order reaction.

$$\underset{\text{Cane sugar}}{C_{12}H_{22}O_{11}} + \underbrace{H_2O}_{} + H^+ \rightarrow C_6H_{12}O_6 + C_6H_{12}O_6$$

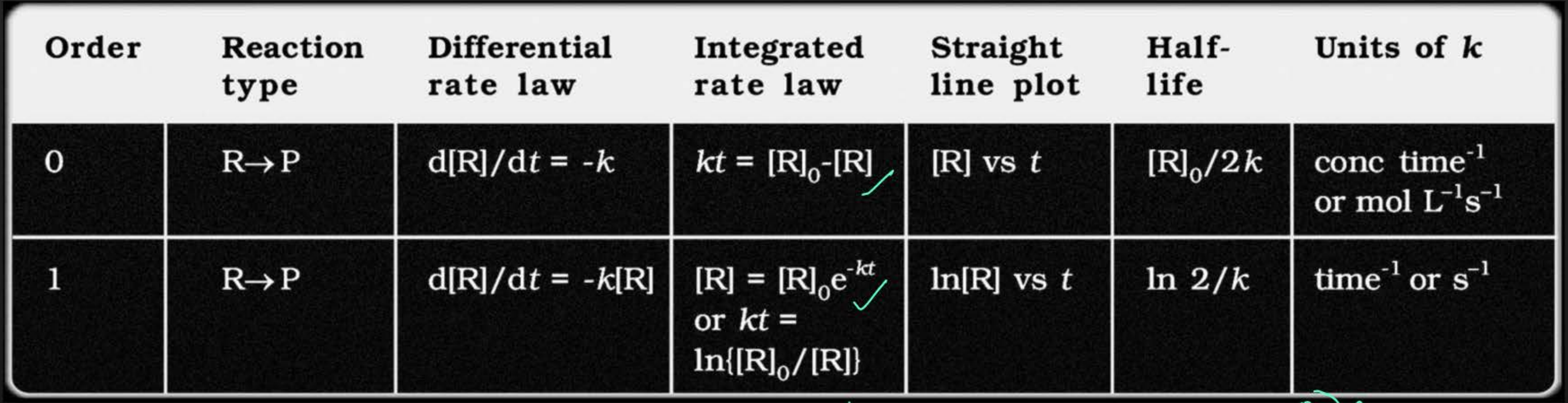
$$\text{Rate} = k\,[C_{12}H_{22}O_{11}]^1\,[H_2O]^0$$

→ hydrolysis of ester

$$H_3C-COOC_2H_5 + \boxed{H_2O} \longrightarrow CH_3COOH + C_2H_5OH$$

$$\text{Rate} = k\,[CH_3COOC_2H_5]^1\,[H_2O]^0 .$$

Integrated Rate Laws for the Reactions of Zero and First Order



Order	Reaction type	Differential rate law	Integrated rate law	Straight line plot	Half-life	Units of k
0	$R \rightarrow P$	$d[R]/dt = -k$	$kt = [R]_0 - [R]$ ✓	$[R]$ vs t	$[R]_0/2k$	conc time^{-1} or mol $L^{-1}s^{-1}$
1	$R \rightarrow P$	$d[R]/dt = -k[R]$	$[R] = [R]_0 e^{-kt}$ ✓ or $kt = \ln\{[R]_0/[R]\}$	$\ln[R]$ vs t	$\ln 2/k$	time^{-1} or s^{-1}

$$\frac{[R]}{[R]_0} = e^{-kt}$$

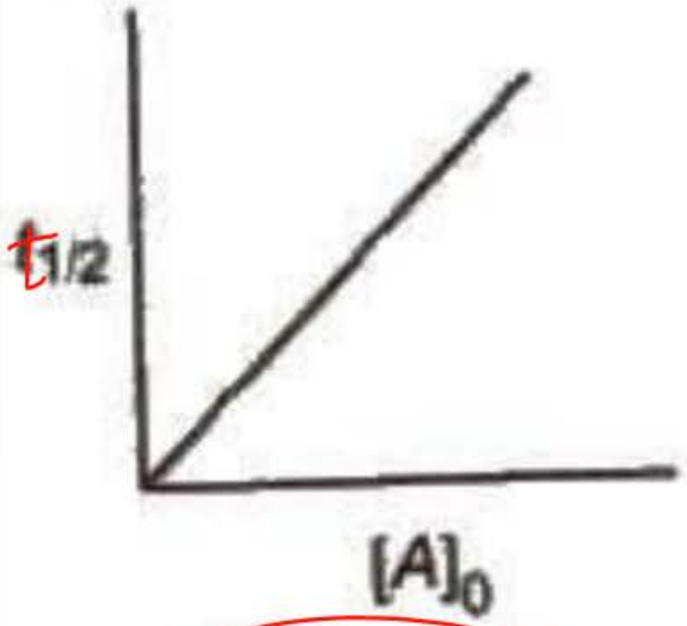
$$K = \frac{2.303}{t}\log\frac{[R]_0}{[R]}$$

$t_{1/2} = \frac{[A]_0}{2K}$

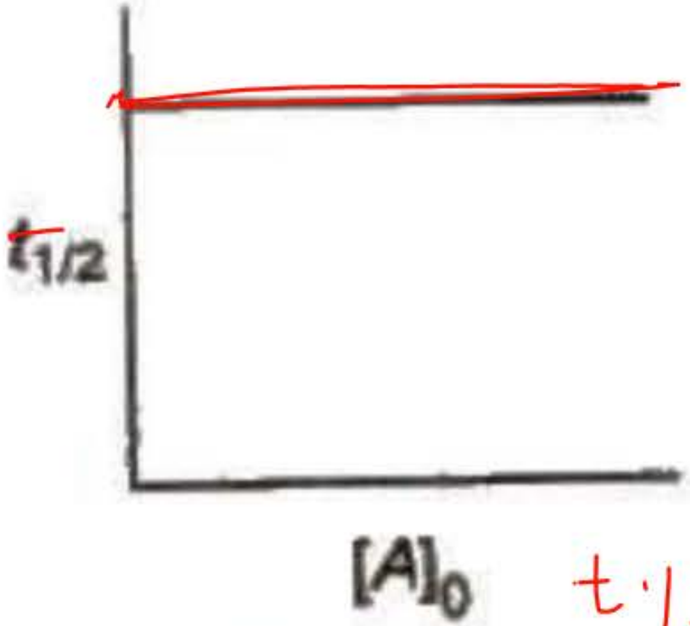
$$\boxed{t_{1/2} \propto \frac{1}{[A]_0^{n-1}}}$$

$t_{1/2} \alpha \frac{1}{[A]_0^{n-1}}$

$t_{1/2} \alpha \frac{1}{[A]_0^{1}}$

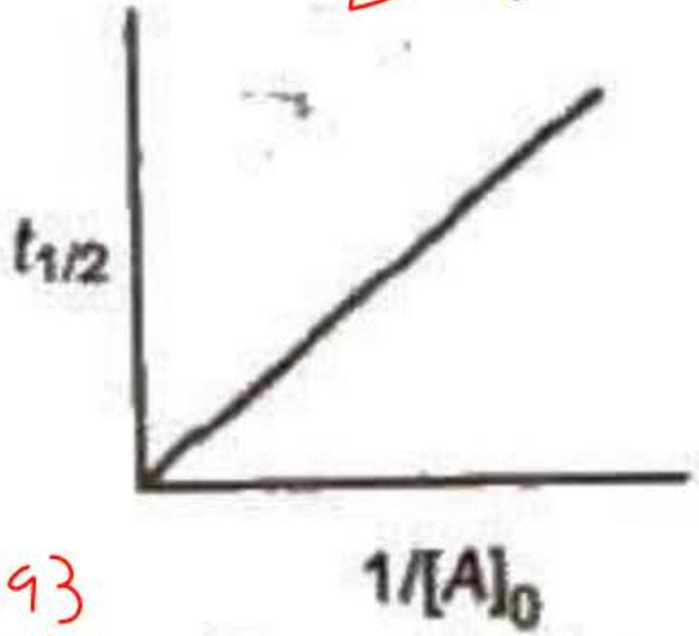


Zero order



First order

$t_{1/2} = \frac{0.693}{K}$



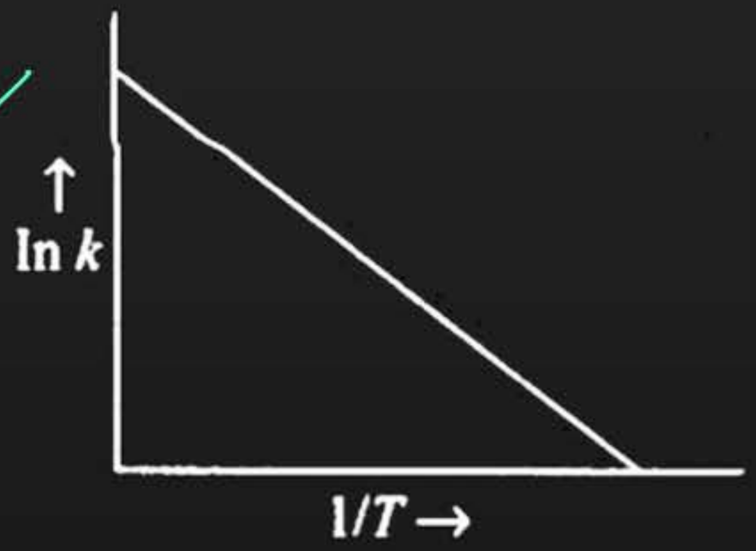
Second order

Arrhenius equation.

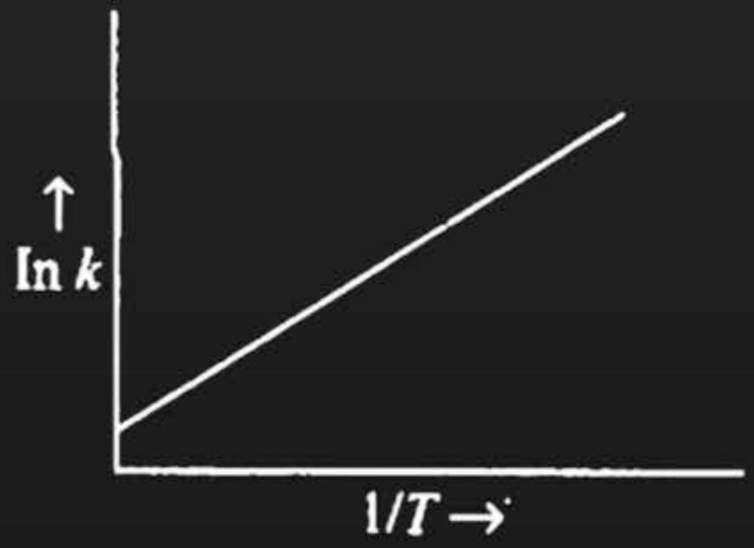
Question

According to Arrhenius equation rate constant k is equal to $Ae^{-Ea/RT}$. Which of the following options represents the graph of In k versus $1/T$?

[✓ A]



B



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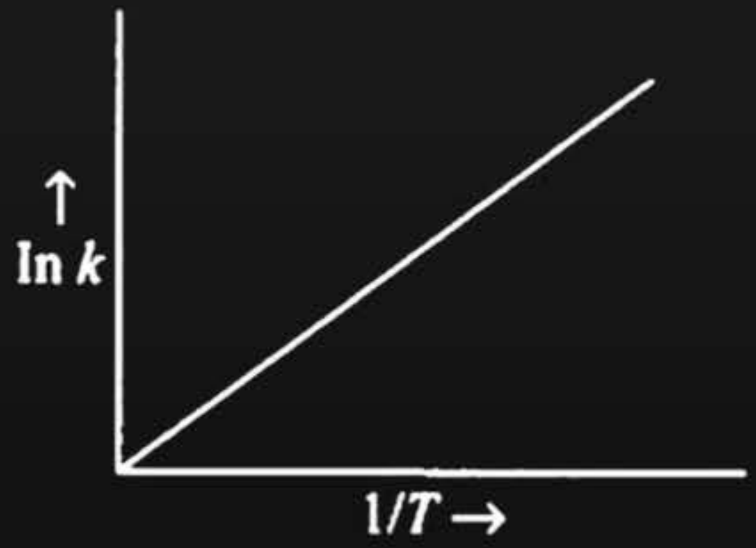

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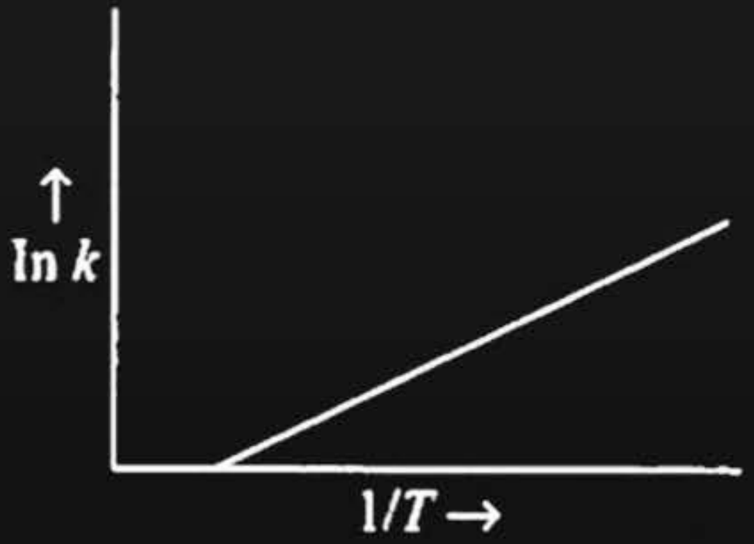
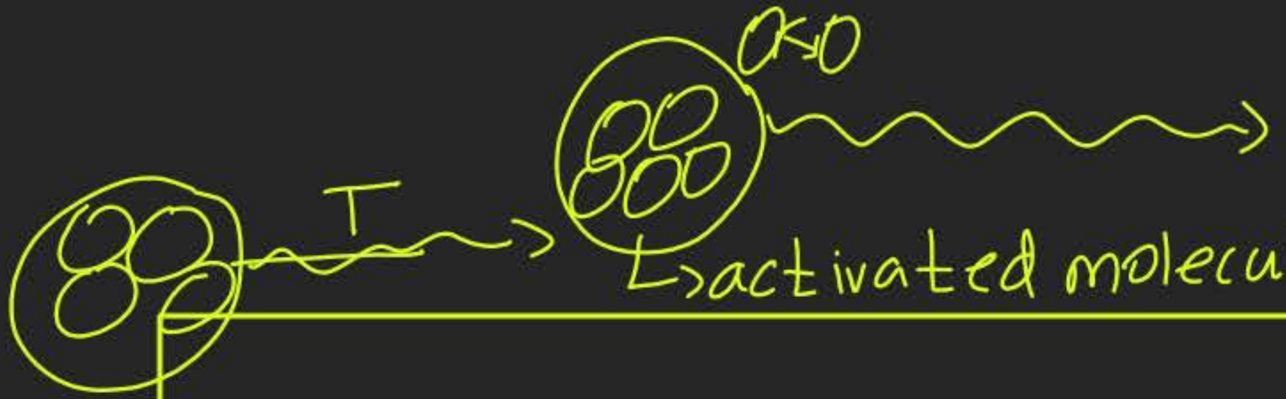




↳ activated molecule.

{ Minimum energy } activated molecule → should have to bring about effective collision to form products.

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

➢ $\boxed{k = Ae^{-Ea/RT}}$ → $\ln_e k = \ln_e A + \ln_e e^{-Ea/RT}$

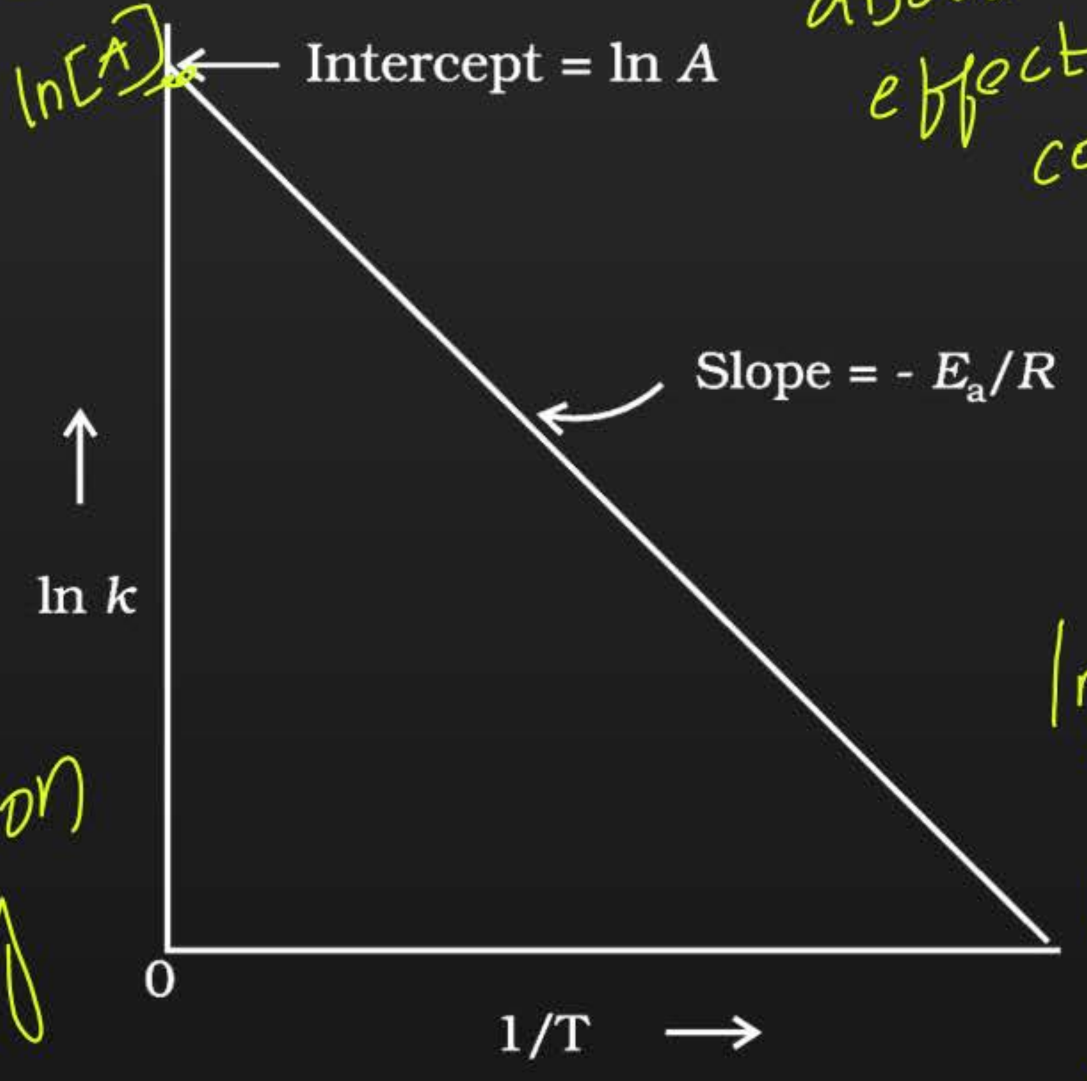
$\ln k = \ln A - \frac{Ea}{RT}$

→ preexponential factor

→ frequency factor

Ea = activation energy

$e^{-Ea/RT}$ → fraction of molecule having energy equal or greater than activation energy.



$\ln_e e = 1$.

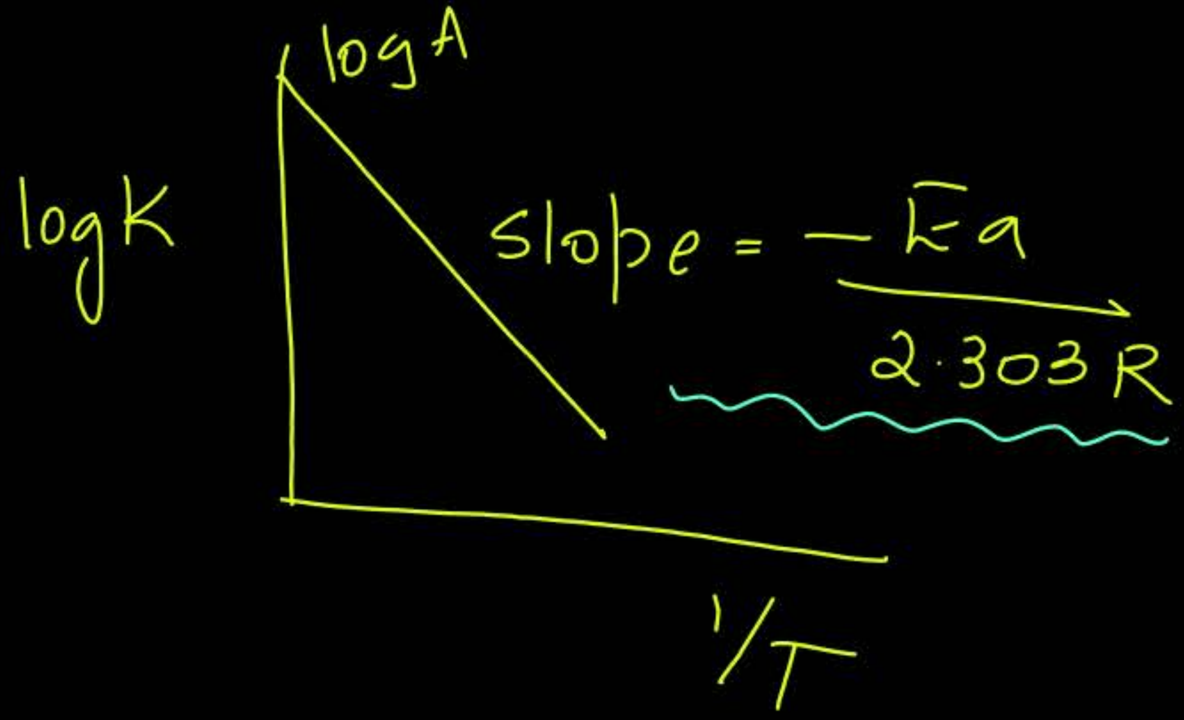
$\ln_e e^a$

$a \ln_e e = 1$

$\boxed{a}$

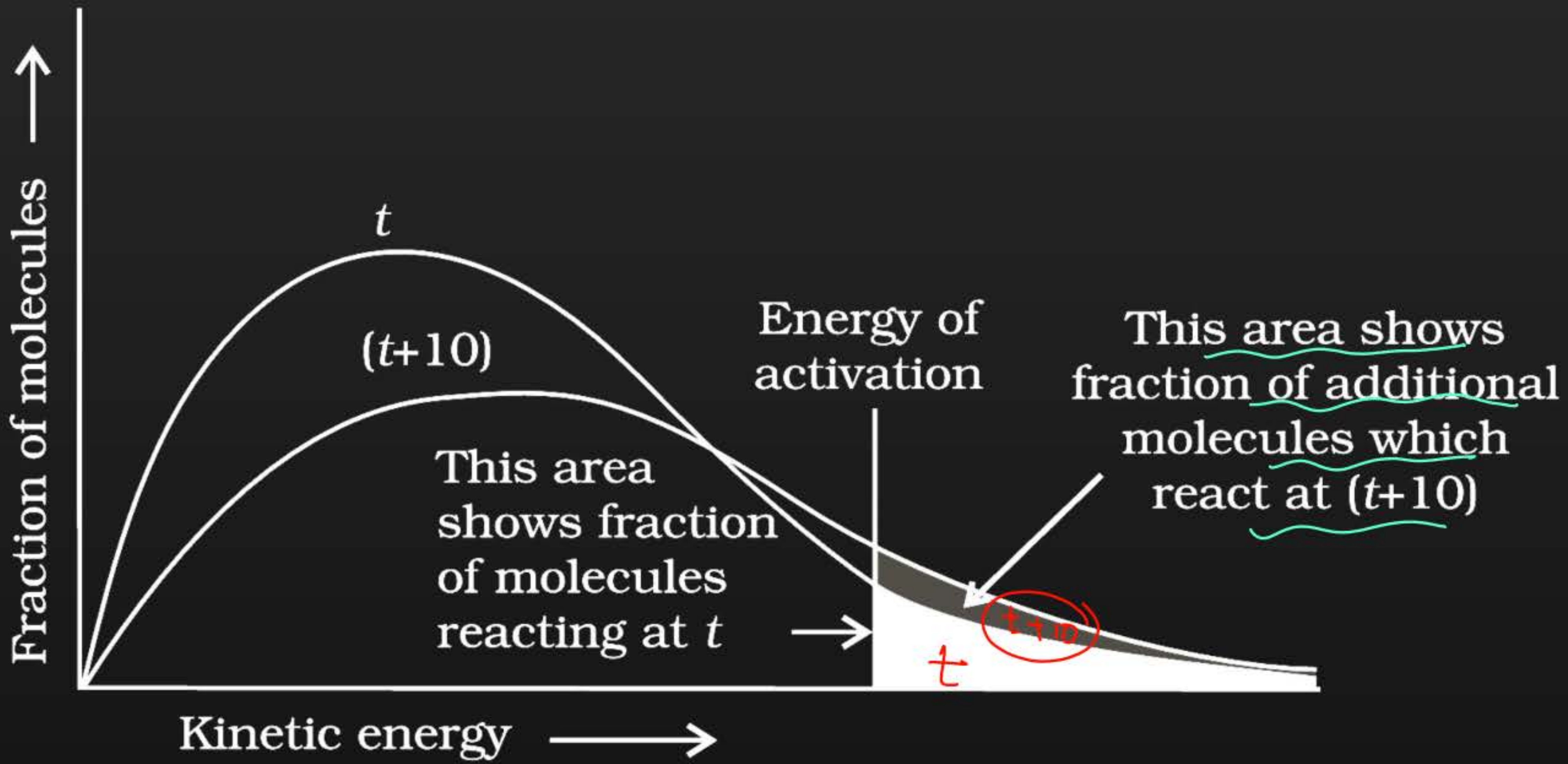
$$\ln k = \ln A - \frac{E_a}{RT}$$

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

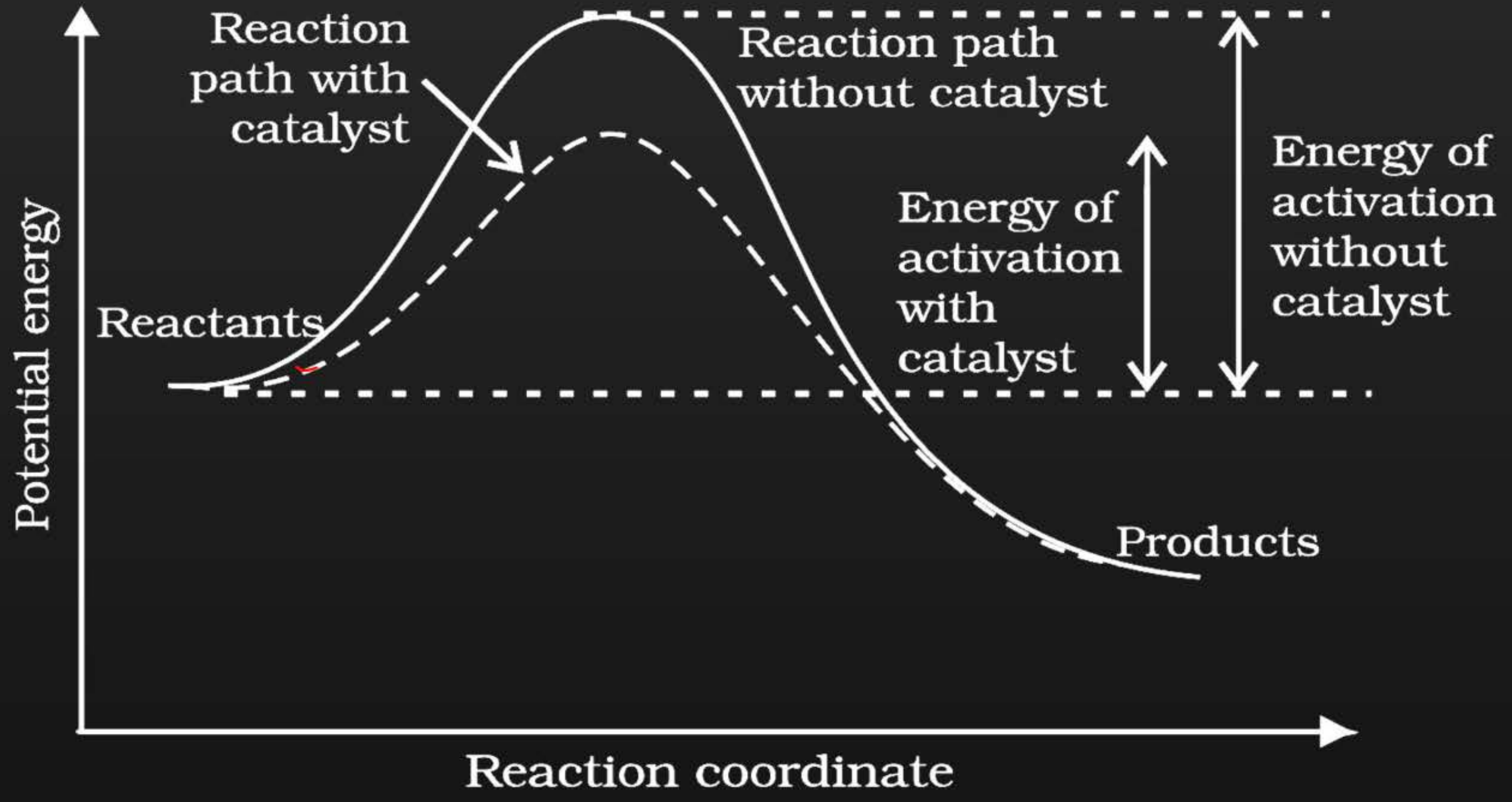


$$\text{slope} = -\frac{E_a}{2.303\,R}$$

Distribution curve showing temperature dependence of rate of a reaction



Effect of catalyst on activation energy



Ea

Effect of Catalyst

$$\boxed{2\mathrm{KClO_3} \xrightarrow{\mathrm{MnO_2}} 2\mathrm{KCl} + 3\mathrm{O_2}}$$

- It is clear from Arrhenius equation , that lower the value of activation energy faster will be the rate of a reaction.
- A small amount of the catalyst can catalyse a large amount of reactants.
- A catalyst does not alter Gibbs energy, G of a reaction.
- It catalyses the spontaneous reactions but does not catalyse non-spontaneous reactions.

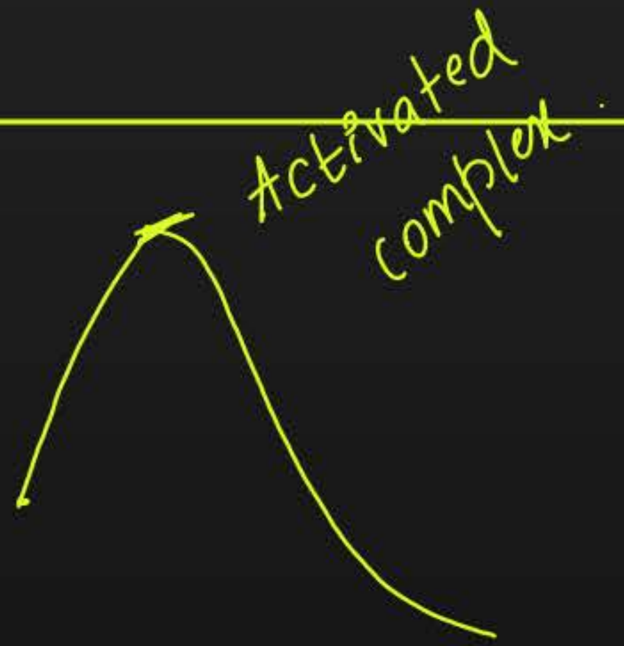
Kc

- It is also found that a catalyst does not change the equilibrium constant of a reaction rather, it helps in attaining the equilibrium faster.
- it catalyses the forward as well as the backward reactions to the same extent so that the equilibrium state remains same but is reached earlier.

Collision Theory of Chemical Reactions

According to this theory, the reactant molecules are assumed to be hard spheres and reaction is postulated to occur when molecules collide with each other.

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



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

where Z_{AB} represents the collision frequency of reactants, A and B and $e^{-Ea/RT}$ represents the fraction of molecules with energies equal to or greater than E_a.

Comparing with Arrhenius equation, we can say that A is related to collision frequency.

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.

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

collision frequency factor.

Orientation of molecule

Question

Which of the following graph represents exothermic reaction?

- **A** (i) only
- **B** (ii) only
- **C** (iii) only
- **D** both (i) and (ii)

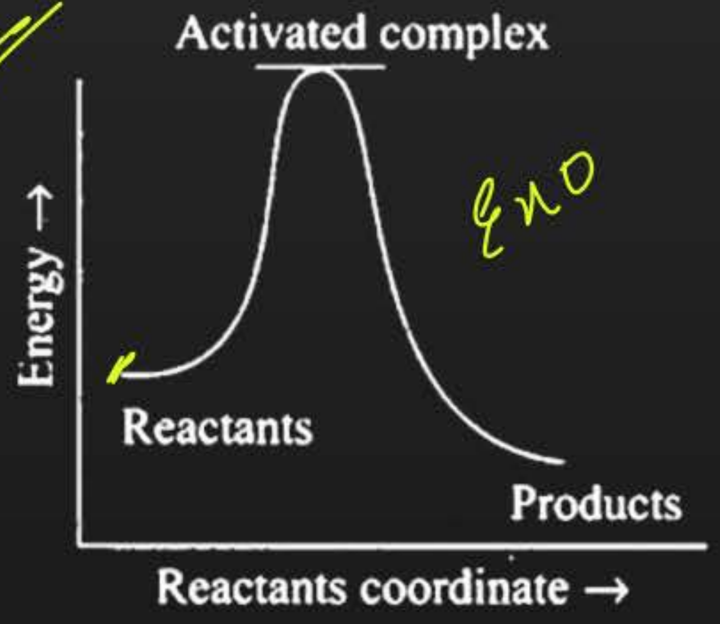
Exothermic

[R] → energy
↓
more

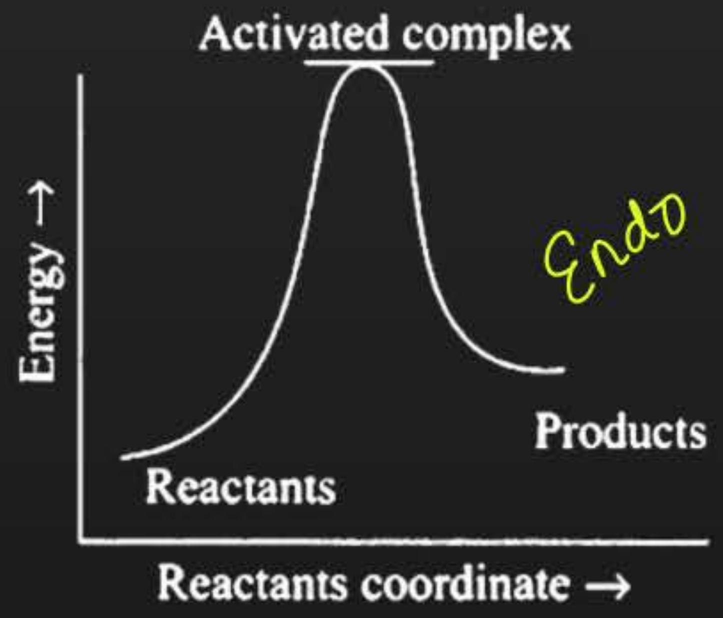
Endothermic

[P] → energy
↓
more

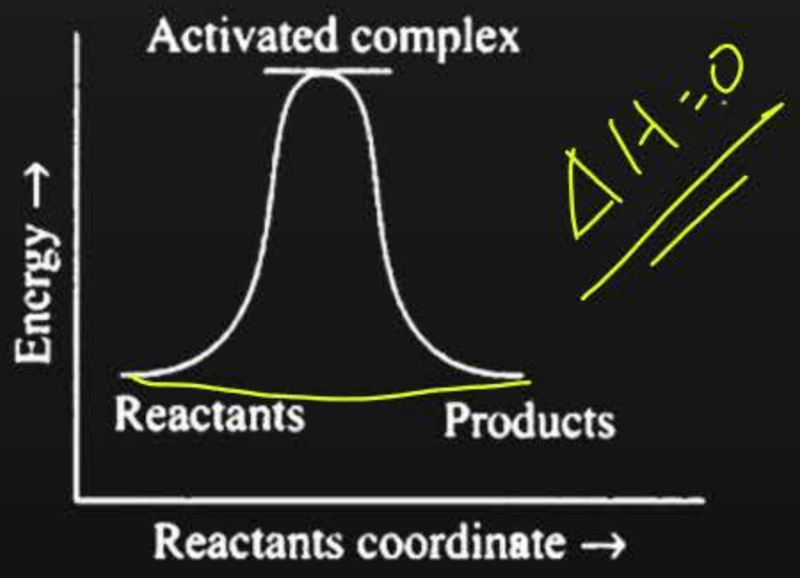
(i)



(ii)



(iii)



- A (A) → (iv), (B) → (iii), (C) → (ii), (D) → (i)
- B (A) → (i), (B) → (ii), (C) → (iii), (D) → (iv)
- C (A) → (ii), (B) → (i), (C) → (iv), (D) → (iii)
- D (A) → (i), (B) → (ii), (C) → (iv), (D) → (iii)

$X \longrightarrow Y$

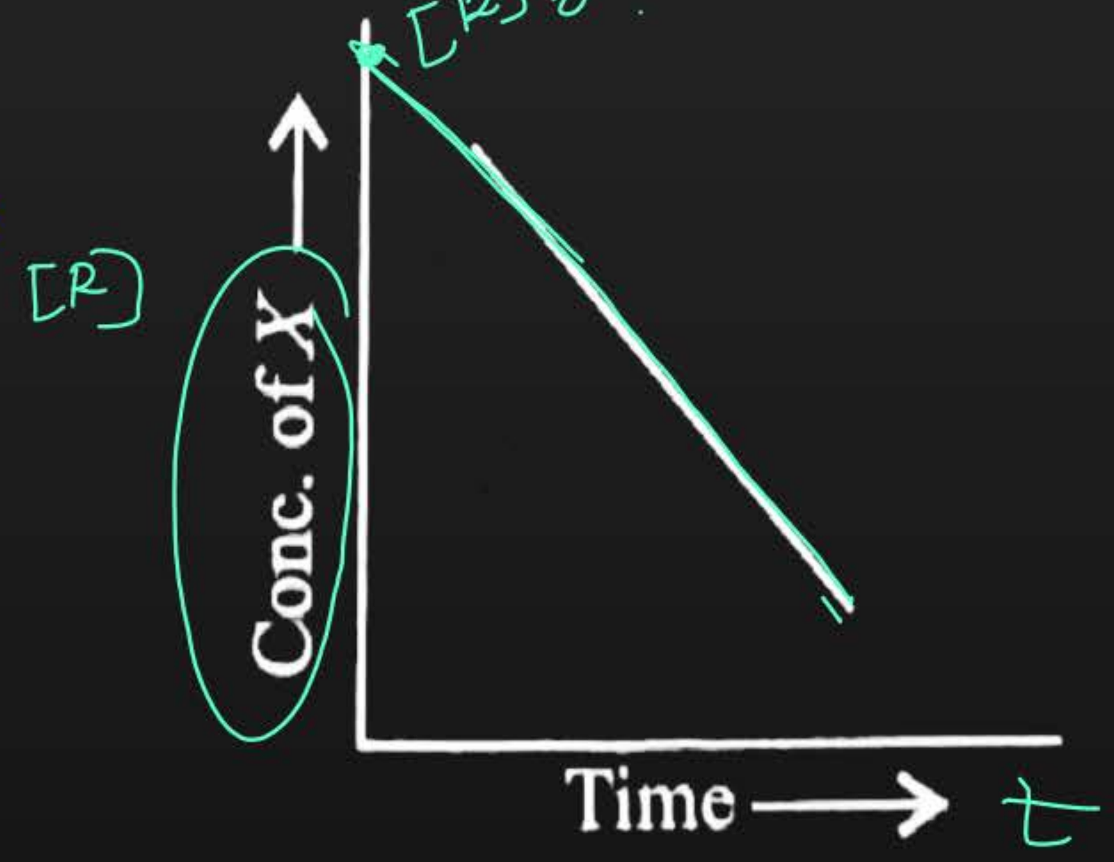
QUESTION

For a general reaction $X \rightarrow Y$, the plot of conc. of X vs time is given in the figure. What is the order of the reaction and what are the units of rate constant ?

$\text{Rate} = k[X]^0$

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

- **A** Zero, mol L^{-1} s^{-1}
- **B** First, mol L^{-1} s^{-1}
- **C** First, s^{-1}
- **D** Zero, L mol^{-1} s^{-1}



QUESTION

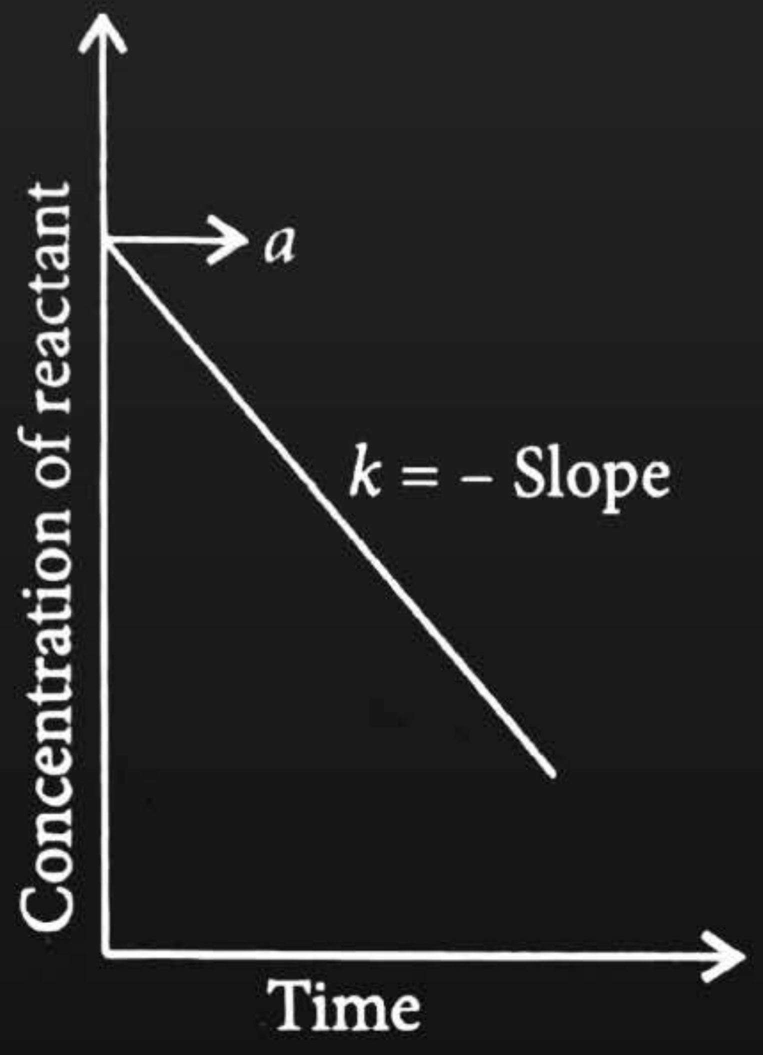
The above plot is for ________ order reaction to calculate value of rate constant.

A second

B first

C zero

D first and zero



QUESTION

In a reaction, X → Y the concentration of X decreases from 0.50 M to 0.38 M in 10 min. What is the rate of reaction in M s^{-1} during this interval ?

$= mol\,L^{-1}\,s^{-1} \longrightarrow$ zero order.

- (A) 2×10^{-4}
- (B) 4×10^{-2}
- (C) 2×10^{-2}
- (D) 1×10^{-2}

$$k = \frac{[R]_0 - [R]}{t} = \frac{[0.50] - [0.38]}{10 \times 60}$$

$$= \frac{0.12}{600} = \frac{\cancel{12}\,2}{100 \times \cancel{6} \times 100} \rightarrow \underline{2 \times 10^{-4}}\ mol\,L^{-1}\,s^{-1}$$

QUESTION

For the reaction $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O_2$ if the rate of disappearance of NH_3 is 3.6×10^{-3} mol $L^{-1}s^{-1}$, what is the rate of formation of H_2O ?

- **A** ✓ 5.4×10^{-3} mol L^{-1} s^{-1}
- **B** 3.6×10^{-3} mol L^{-1} s^{-1} ✗
- **C** 4×10^{-x} mol L^{-1} s^{-1}
- **D** 0.6×10^{-4} mol L^{-1} s^{-1}

$$-\frac{1}{4}\frac{d[NH_3]}{dt} = \frac{1}{6}\frac{d[H_2O]}{dt}$$

$$\frac{d[H_2O]}{dt} = -\frac{6}{4}\frac{d[NH_3]}{dt}$$

$$= \frac{-6}{4} \times 3.6 \times 10^{-3} = 5.4 \times 10^{-3}\ mol\,L^{-1}s^{-1}$$

QUESTION

Given $t_{1/2}$ = 3 hours, then how many gram of a substance will remain after 18 hours from 300 gram of a substance?

- **A** 4.6 gram
- **B** 5.6 gram
- **C** 9.2 gram
- **D** 6.4 gram

$t_{1/2} = 3$ hours.

$T = 18$ hours

$[R]_0 = 300g$

$[R] = ?$

$$n = \frac{\text{Total Time}}{\text{Half-life } (t_{1/2})} = \frac{18 \text{ hours}}{3 \text{ hours}} = \underline{6 \text{ half-lives}}$$

✓

The remaining amount (N_t) is calculated by taking the initial amount (N_0) and multiplying it by $\frac{1}{2}$ raised to the power of the number of half-lives (n).

$$\boxed{N_t = N_0 \times \left(\frac{1}{2}\right)^n}$$

or

$2 \times 2 \times 2 \times 2 \times 2 \times 2$

✓✓

$$N_t = \underline{300} \times \left(\frac{1}{2}\right)^{⑥}$$

$$N_t = 300 \times \frac{1}{64}$$

$$N_t = 4.6875 \text{ g}$$

or

$$[R] = [R]_0 \times \left(1/2\right)^n \longrightarrow \boxed{\text{No. of half life.}}$$

$[R]$ ↓ final conc.

$[R]_0$ ↓ intial conc

QUESTION

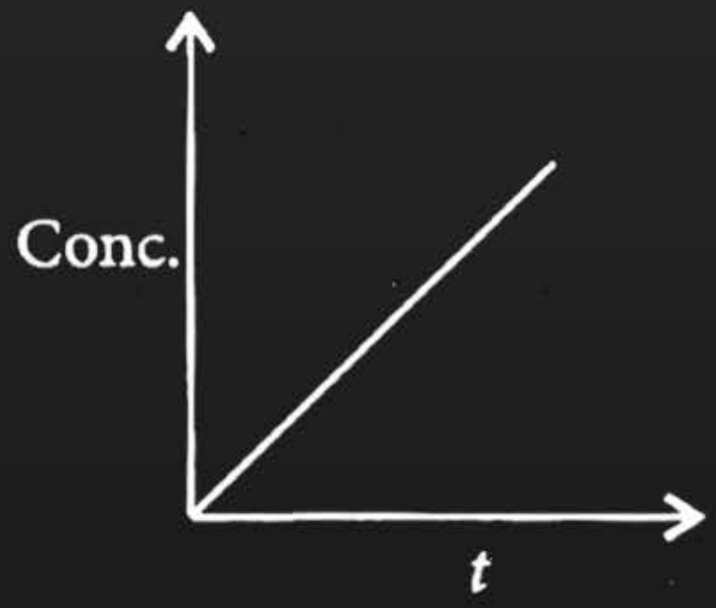
A graph corresponding to a first order reaction is

$$\boxed{\text{Rate} = k[\text{conc}]} + 0$$

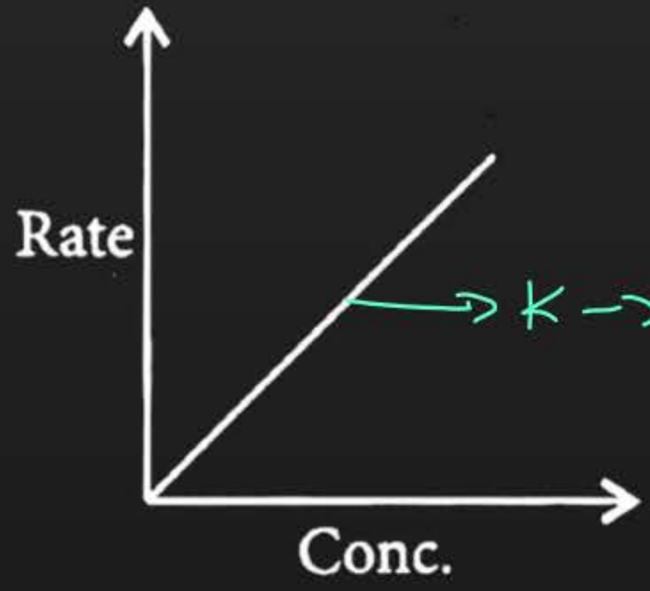
Rate ↓ y conc ↓ x

$y = \boxed{m}x$

A

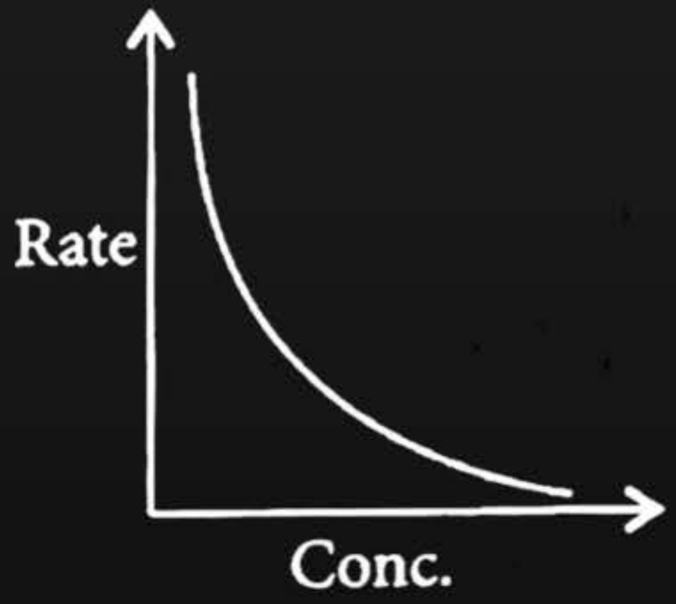


[✓ B]

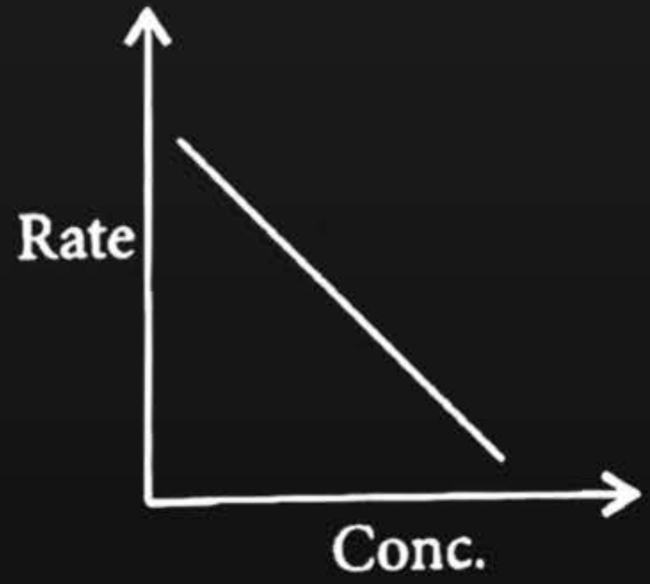


→ k → slope.

C



D



QUESTION

Which of the following relation is correct for zero order reaction?

(A) $t_{3/4} = 2t_{1/2}$

[✓ (B)] $t_{3/4} = 1.5\ t_{1/2}$

(C) $t_{3/4} = \frac{1}{2}t_{1/2}$

(D) $t_{3/4} = \frac{1}{3}t_{1/2}$

$t_{3/4} = \left[2^{n-1}+1\right] t_{1/2}.$ $n=0$

$t_{3/4} = \left[2^{0-1}+1\right] t_{1/2}$

$t_{3/4} = \left[2^{-1}+1\right] \cdot t_{1/2}$

$= \left[\frac{1}{2}+1\right] t_{1/2} \longrightarrow 3/2\, t_{1/2}.$

$t_{3/4} = 1.5 \cdot t_{1/2}$

zero order

$$\boxed{t_{1/2} = \frac{[A]_0}{2K}}$$

$$100\,g \xrightarrow{t_{1/2}=20\,min} 50\,g \xrightarrow{t_{1/2}=10\,min} 25\,g \xrightarrow{t_{1/2}=5\,min} 12.5\,g \xrightarrow{2.5\,min} 6.25\,g$$

3.125

$$\frac{\cancel{6}.25}{\cancel{2}\ 1}$$

first order

$$\boxed{t_{1/2} = \frac{0.693}{K}}$$

$$100\,g \xrightarrow{t_{1/2}=20\,min} 50\,g \xrightarrow{t_{1/2}=20\,min} 25\,g \xrightarrow{t_{1/2}=20\,min} 12.5\,g \xrightarrow{t_{1/2}=20\,min} 6.25\,g \xrightarrow{t_{1/2}=20} 3.125\,g$$

$$\boxed{t_{75\%} = 2\,t_{1/2}}$$

$$t_{87.5\%} = 3\,t_{1/2}.$$

$$t_{93.85\%} = 4\,t_{1/2}.$$

$$t_{96.875\%} = 5\,t_{1/2}.$$

$$t_{99.9\%} = $$

$$\boxed{t_{99\%} = \qquad t_{1/2}} \qquad \boxed{t_{99.9\%} = \qquad t_{1/2}}$$

Order	Zero	1st	2nd	nth
Differential rate law	Rate = k	Rate= k[A]	Rate = $k[A]^2$	Rate = $k[A]^n$
Integrated rate law	$[A_0]-[A]_t = kt$	$kt = \ln\frac{[A]_0}{[A]_t}$	$kt = \frac{1}{[A]_t} - \frac{1}{[A]_0}$	$kt = \frac{1}{(n-1)}\left[\frac{1}{[A]_t^{n-1}} - \frac{1}{[A]_0^{n-1}}\right]$
Half life ($t_{1/2}$)	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{\ln 2}{k}$	$t_{1/2} = \frac{1}{[A]_0 k}$	$t_{1/2} = \frac{1}{k(n-1)}\left[\frac{2^{n-1}-1}{[A_0]^{n-1}}\right]$
($t_{3/4}$)	$t_{3/4}$=1.5 $t_{1/2}$ ✓ $t_{75\%}$	$t_{3/4}$= 2 $t_{1/2}$ ✓ $t_{75\%}$	$t_{3/4}$ = 3 $t_{1/2}$ ✓	$t_{3/4}$= $(2^{n-1}+1)$ $t_{1/2}$ ✓

QUESTION

The half life of a certain first order reaction is 60-minutes. How long will it take for 75% reaction to occur ?

[redacted]

$$t_{75\%} = 2 \cdot \boxed{t_{1/2}}$$

$$= 2 \times 60 \text{ min}$$

$$t_{75\%} = 120 \text{ min}.$$

QUESTION

The rate constant is given by the equation $k = Ae^{-E_a/RT}$ Which factor should register a decrease for the reaction to proceed more rapidly ?

A T

B Z

C A

D E_a

$$\overset{\checkmark}{K} = A \cdot e^{\frac{-Ea}{RT}}$$

$Ea \downarrow$ $K \uparrow$ rate$\uparrow$ $e^{\frac{-Ea}{RT}}$ =

Ea → less
T → more

$e^{-Ea/RT}$ = less -ve

$$\boxed{W = -P\Delta V}$$

$$\boxed{\begin{aligned} W &= -2.303nRT\log\frac{V_2}{V_1} \\ W &= -2.303nRT\log\frac{P_1}{P_2} \end{aligned}}$$

Thermodynamics. ⟶ 02-03 ↓ range of Questions

└→ 2025 → 03 Qs

heat capacity

$C_p - C_v = R$

specific heat capacity

Molar heat capacity

→ first law of thermodynamics

→ Entropy → * Statement based

→ Thermochemistry → enthalpy of

- → formation
- → Combustion
- → Solution
- → lattice energy

Hess's law

System → part of universe under observation.
Surrounding → region near to system.
universe → system + Surroundings.

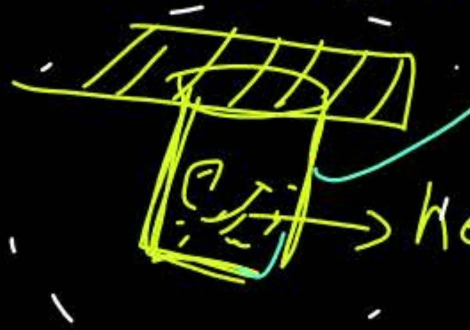
State of the system
→ present condition.

Types of System

open system ✓

→ matter
→ heat
→ conducting wall or diathermic

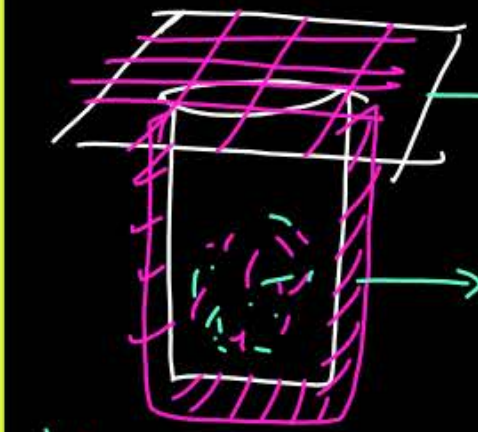
Closed system



→ conducting wall or Diathermic wall
→ heat (exchanges)

Matter → do not exchange.

isolated System



→ closed
→ adiabatic wall

[Isolated system]

State functions.

→ property depends only on state but not path

→ pressure
→ Entropy
→ Enthalpy
→ internal energy
→ Temp.

Path function
work
heat

$$\boxed{W = -P_{ext}\Delta V}$$

First law of thermodynamics

Can be states as – “The energy of an isolated system is constant” - $\Delta U = q + w$

It is commonly stated as the law of conservation of energy i.e., energy can neither be created nor be destroyed.

$$\boxed{\Delta U = q + w}$$

heat absorbed by the system → $+q$ ✓

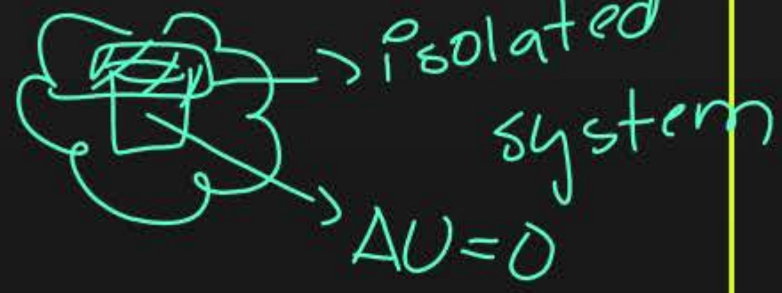
heat released from the system $= -q$ ✓

Expansion

work done by the system $-w$

compression

work done on the system → $w = +ve$.



Thermodynamic processes

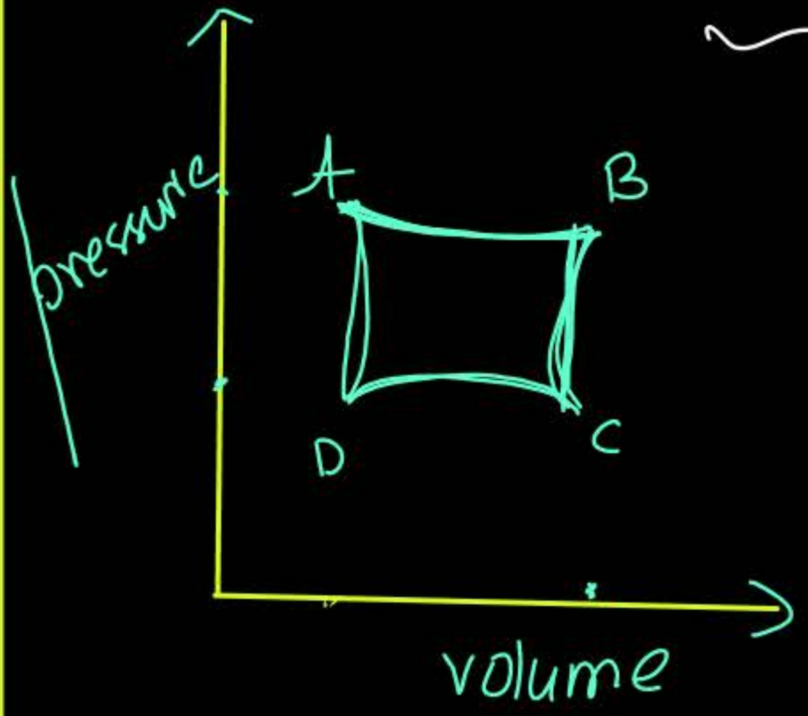
$\boxed{\Delta U = \Delta T}$

→ Adiabatic process → $q=0$

$\boxed{\Delta U = W}$ heat constan

isothermal process * Temp is constant, $\Delta T = 0$, $\boxed{\Delta U = 0}$

$\Delta U = q + w$



*Cyclic system → A → B → C → D → A

$\Delta V = 0$ $\Delta T = 0$

$\Delta P = 0$

* pressure is constant → isobaric process → A-B

D to C

$\Delta U = q + w \rightsquigarrow \boxed{\Delta U = q}$

$W = -P_{ext} \Delta V$

* volume is constant → isochoric process

$\Delta V = 0$

$W = -P_{ext} \Delta V$

$W = 0$

$\boxed{\Delta U = q}$

B → C
A → D

$w = 0$

Extensive property → Mass of an apple
→ amount of substance.

Intensive property → taste of an apple.
→ does not depend upon the amount

Extensive property	Intensive property
Mass	Temperature
Volume	Pressure $= \frac{\text{Force}}{\text{area}} = \frac{E}{E}$
Heat capacity → 1°C / 1K ↑	Density $= \frac{\text{Mass}}{\text{Volume}} = \frac{E}{E} = I$
Enthalpy	Refractive index = 
Work energy ✓	Viscosity
Entropy ✓	Specific heat 1g of substance → 1°C → heat
Gibbs's free energy ✓	Molarity → Moles
Internal energy ✓	Molality → Moles
	Mole fraction
	Molar entropy

QUESTION

The final temperature in an adiabatic expansion is

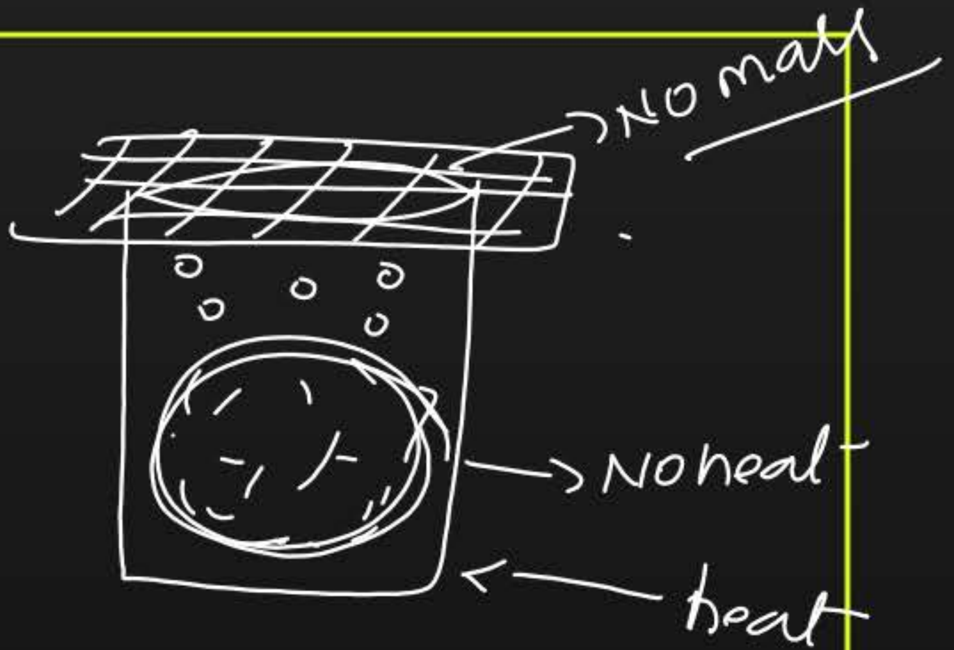
A greater than the initial temperature

B same as the initial temperature

C half of the initial temperature

D less than the initial temperature.

$\boxed{\Delta U = W}$ → expansion



QUESTION

1 KJ = 1 Nm.

10 = deca

$\frac{1}{10}$ = deci

The external pressure needed to compress an ideal gas from 22 dm^3 to 8 dm^3 if work done is of 4.545 kJ is

A 3.24×10^5 N m^{-2}

B 2.03×10^5 N m^{-2}

C 1×10^5 N m^{-2}

D -3.4×10^5 N m^{-2}

$$W = -P_{ext}\,\Delta V$$

$$P_{ext} = -\frac{W}{\Delta V}$$

$$P_{ext} = -\left[\frac{4.545\times10^{3}\,J}{(8-22)\,dm^3}\right]$$

$$P_{ext} = \left[\frac{4.545\times10^{3}}{14\times10^{-3}}\right] = \frac{4.545\times10^{3}}{14\times10^{-3}}$$

$$0.324\times10^{6}$$

$$= 3.24\times10^{-1}\times10^{6}$$

$$= 3.24\times10^{5}$$

$W = 4.545$ KJ

$P_{ext} = ?$

$\Delta V = V_2 - V_1$

$V_1 = 22 dm^3$

$V_2 = 8 dm3$

1 decimeter = 0.1 m

$1 dm^3 = (0.1)^3 m^3$

$(10^{-1})^3 m^3$

$1 dm^3 = 1 L = 10^{-3} m^3$

$$= \frac{4.545\times10^{3}\ \mathrm{Nm}}{(8-22)\ \mathrm{dm}^{3}}$$

$$= \frac{+\,4.545\times10^{3}\ \mathrm{Nm}}{14\times10^{-3}\ \mathrm{m}^{+3}}$$

$$= \overset{\curvearrowright}{0.3}2\times10^{\textcircled{+6}}\ \mathrm{N}\,\mathrm{m}^{\cancel{+1}}\,\mathrm{m}^{\cancel{-3}}$$

$$= \boxed{3.2\times10^{+5}\ \mathrm{N}\,\mathrm{m}^{-2}}$$

QUESTION

Two grams of hydrogen gas at STP expands so that the volume is doubled. The work done during this process is

A –11.2 L atm

[✓ B] –22.4 L atm

C +44.8 L atm

D 11.2 L atm

$$\boxed{W = -P_{ext}\Delta V}$$

STP

→ 273 K = 0°C

→ P = 1 atm

$2g\ H_2 \longrightarrow 22.4\,L \times 2$

$= 44.8\,L$

$22.4\,L \rightarrow V_1$

$44.8\,L \rightarrow V_2$

$$W = -\left[1\,\text{atm} \times V_2 - V_1\right]$$

$$W = -\left[1\,\text{atm} \times (44.8 - 22.4)\right]$$

$$\boxed{W = -22.4\ \text{L atm}} \longleftrightarrow$$

work done by the system → Expansion.

QUESTION

In an adiabatic expansion of ideal gas

A $W = -\Delta E$

B $W = \Delta E$

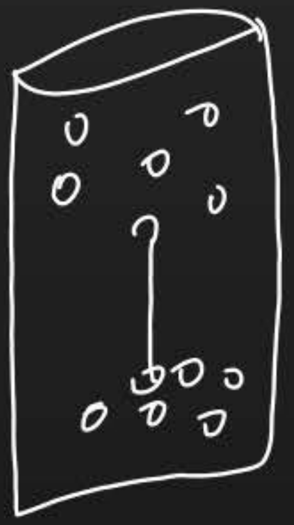
$\underset{\Delta E}{\Delta U} = W$

C $\Delta E = 0$

D $W = 0$

$\Delta E = \Delta U$

↳ internal energy



$\underline{\underline{\Delta U = W}}$

$\underset{\Delta E}{\Delta U} = \overset{\checkmark}{W} + \overset{\checkmark}{q}$

QUESTION

$R = 8.314\ \boxed{J}\ K^{-1} mol^{-1}$

$= 8.314 \times 10^{7}$

The work done (in erg) for the reversible expansion of one mole of an ideal gas from a volume of 10 L to 20 L at 298 K is

V_1 V_2

$-2.303 \times n \times R \times T \log \frac{V_2}{V_1}$

$= -2.303 \times 1 \times 8.314 \times 298 \times \log \frac{\cancel{20}^{2}}{\cancel{10}_{1}}$

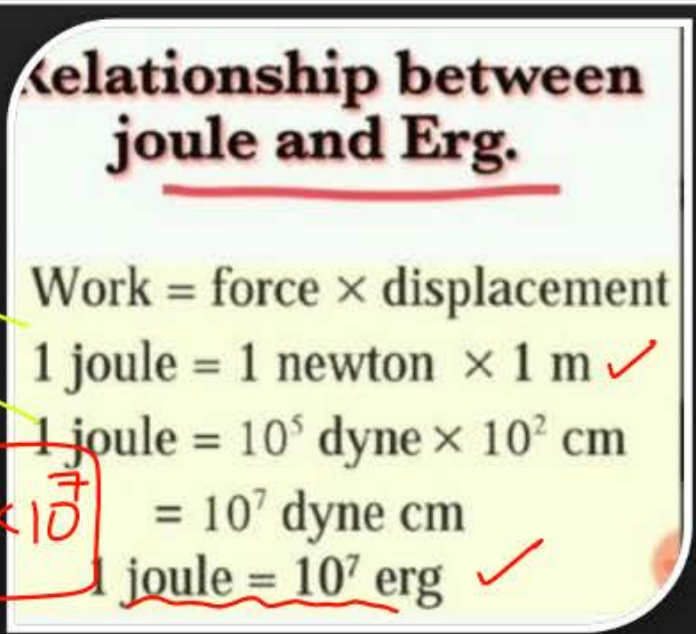
$= -2.303 \times 8.314 \times 298 \times \log_{10} 2 \times 10^{7}$

A $-2.303 \times 298 \times 0.082 \times \log_{10}2$

B $-2.303 \times 298 \times 0.082 \times \log_{10}0.5$

C ✓ $-2.303 \times 298 \times 8.2 \times 10^{7} \times \log_{10}2$

D $-2.303 \times 298 \times 2 \times \log_{10}2$



elationship between joule and Erg.

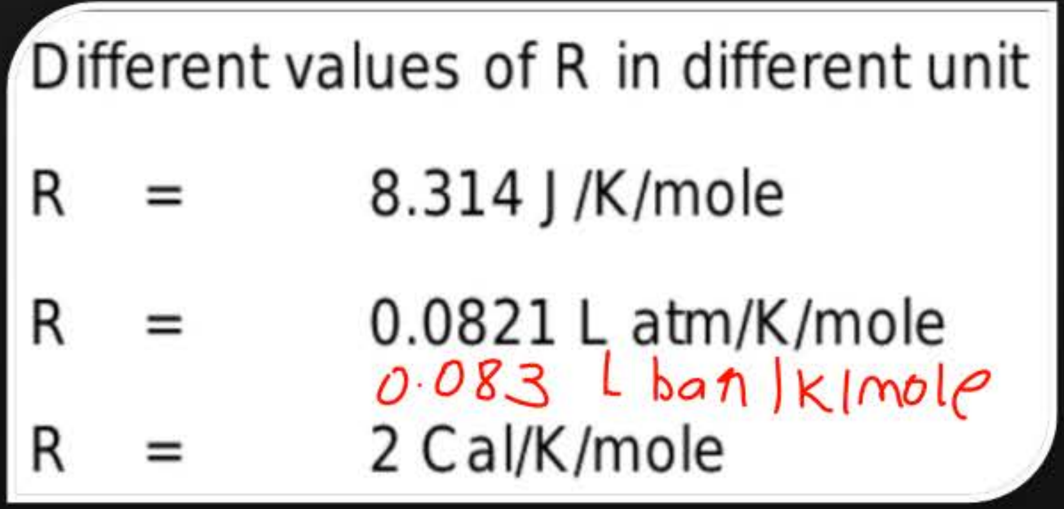
Work = force × displacement

1 joule = 1 newton × 1 m ✓

1 joule = 10^5 dyne × 10^2 cm

= 10^7 dyne cm

1 joule = 10^7 erg ✓



Different values of R in different unit

R = 8.314 J /K/mole

R = 0.0821 L atm/K/mole

0.083 L bar/K/mole

R = 2 Cal/K/mole

QUESTION

3 moles of an ideal gas are expanded isothermally and reversibly from 10 m^3 to 20 m^3 at 300 K. The [redacted] work done is (R = 8.314 J K^{-1} mol^{-1})

20 V_1

Ⓐ +5.187 kJ ✗

Ⓑ −5.187 kJ

Ⓒ −2.175 kJ

Ⓓ +3.750 kJ ✗

$$W = -2.303\, nRT \log \frac{V_2}{V_1}$$

$$= -2.303 \times 3 \times 8.314 \times 300 \times \log \frac{\overset{2}{\cancel{20}}}{\cancel{10}}$$

$$= -2.3 \times ③ \times 8.3 \times ③ \times 10^{2} \times ③ \times 10^{-1}$$

$$= 27 \times 8 \times 2$$

$$= 27 \times 16 \Rightarrow -432 \times 10^{1}$$

$$\begin{array}{r} 162 \\ 27+ \\ \hline 432 \end{array}$$

Thank you