Chemical Kinetics

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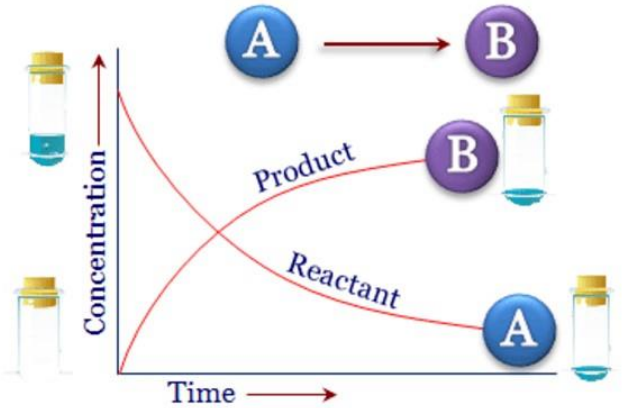

The contents of this presentation is prepared to provide a brief idea about the topics, details will be discussed in the classes.
Contents have been collected from multiple textbooks and internet.

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Rate of a Reaction

- **Reaction rate** is the change in the concentration of a reactant or a product with time (M/s).

$$A \longrightarrow B$$

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

$\Delta[A]$ = change in concentration of A over time period Δt

$$\text{rate} = \frac{\Delta[B]}{\Delta t}$$

$\Delta[B]$ = change in concentration of B over time period Δt

Because [A] decreases with time, **$\Delta[A]$ is negative**.

Unit: mole L^{-1} s^{-1}, mole L^{-1} min^{-1}, mole L^{-1} h^{-1} etc.

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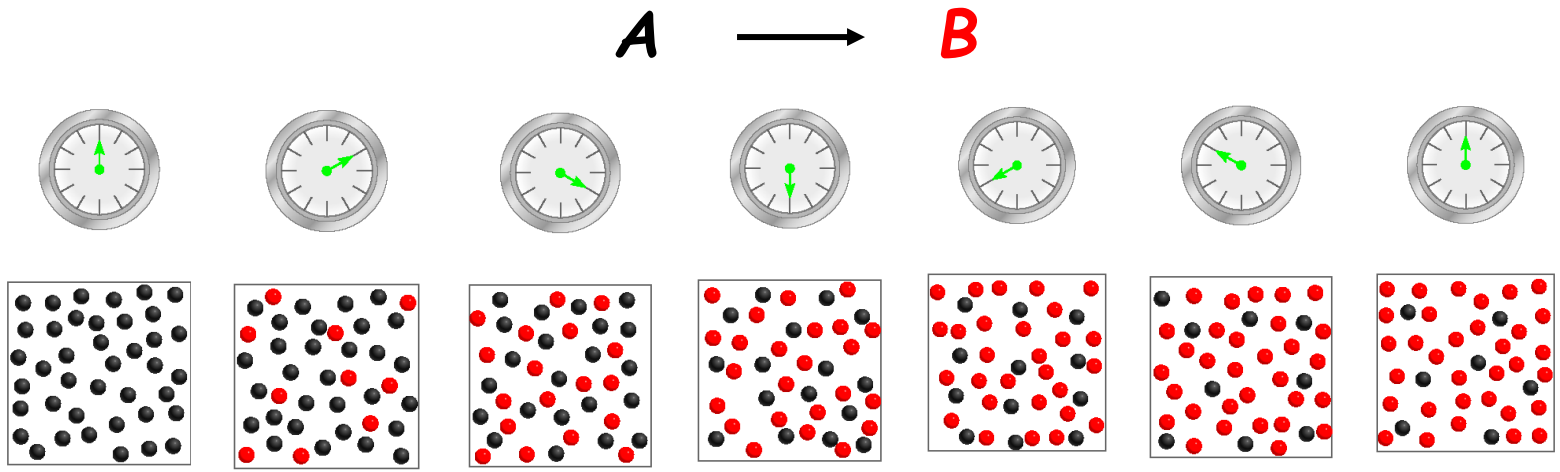
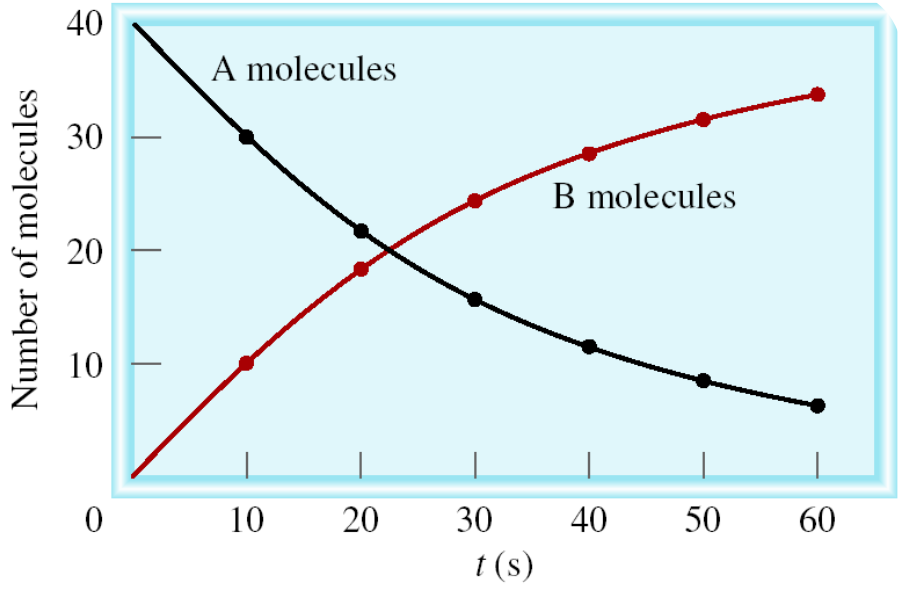
Rate of a Reaction





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

$$rate = \frac{\Delta[B]}{\Delta t}$$

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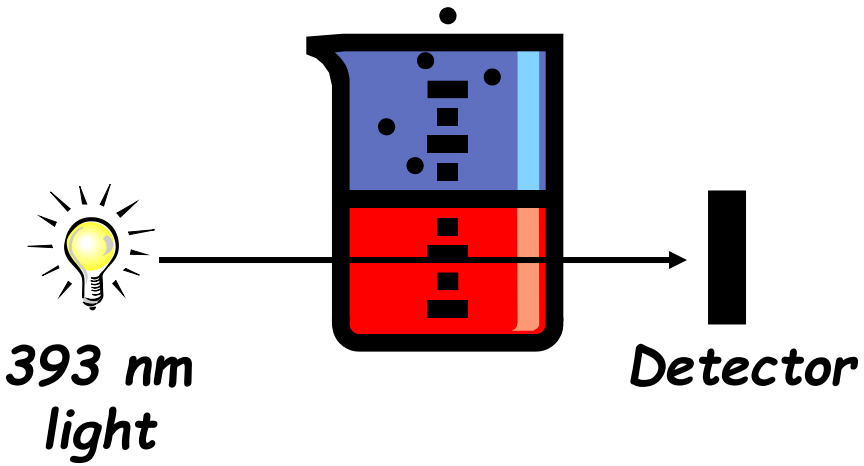
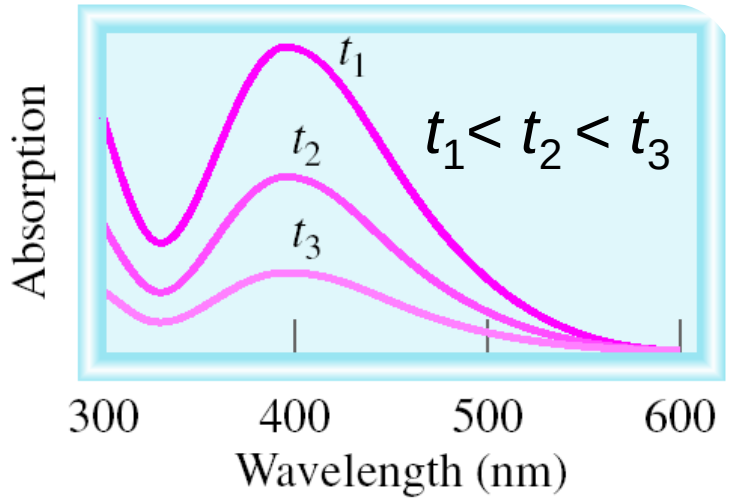
Rate of a Reaction

$Br_2(aq) + HCOOH(aq) \longrightarrow 2Br^-(aq) + 2H^+(aq) + CO_2(g)$

red-brown

time





$\Delta[Br_2]$ a ΔAbsorption

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Reaction Rates and Stoichiometry

$$2A \longrightarrow B$$

Two moles of A disappear for each mole of B that is formed.

$$\text{rate} = -\frac{1}{2}\frac{\Delta[A]}{\Delta t} \qquad \text{rate} = \frac{\Delta[B]}{\Delta t}$$

$$aA + bB \longrightarrow cC + dD$$

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

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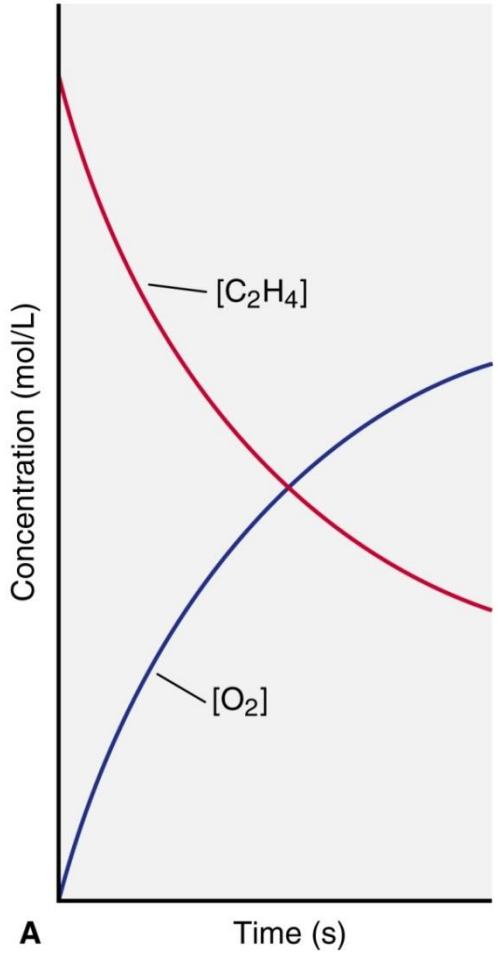
Reaction Rates and Stoichiometry

$$C_2H_4 + O_3 \rightarrow C_2H_4O + O_2$$

$[O_2]$ increases just as fast as $[C_2H_4]$ decreases.

$$\text{Rate} = -\frac{\Delta[C_2H_4]}{\Delta t} = -\frac{\Delta[O_3]}{\Delta t} = \frac{\Delta[C_2H_4O]}{\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$



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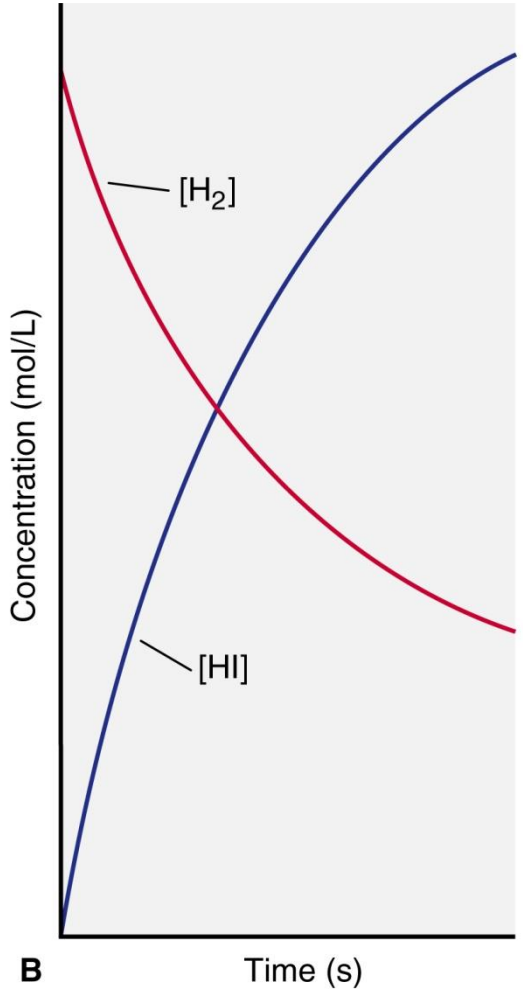
Reaction Rates and Stoichiometry

$$H_2 + I_2 \rightarrow 2HI$$

[HI] increases twice as fast as [H_2] decreases.

$$\text{Rate} = -\frac{\Delta[H_2]}{\Delta t} = -\frac{\Delta[I_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[HI]}{\Delta t}$$



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Reaction Rates and Stoichiometry

- ***Write the rate expression for the following reaction:***

$$CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(g)$$

$$\text{rate} = -\frac{\Delta[CH_4]}{\Delta t} = -\frac{1}{2}\frac{\Delta[O_2]}{\Delta t} = \frac{\Delta[CO_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[H_2O]}{\Delta t}$$

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The Rate Law

- The rate law expresses the relationship of the rate of a reaction to the rate constant and the concentrations of the reactants raised to some powers.

$$aA + bB \longrightarrow cC + dD$$

$$\text{Rate} = k\,[A]^m\,[B]^n$$

  - The term k is the rate constant, which is specific for a given reaction at a given temperature.
  - The exponents m and n are reaction orders and are determined by experiment.
    - The values of m and n are not necessarily related in any way to the coefficients a and b.

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The Order of a Reaction

- *The order of a reaction is defined as the sum of the powers of concentrations in the rate law.*

$$aA + bB \longrightarrow cC + dD$$

$$\text{Rate} = k[A]^m[B]^n$$

- *The order of such a reaction is (m + n)*

Examples of reaction order :

RATE LAW	REACTION ORDER
$\text{rate} = k[N_2O_5]$	1
$\text{rate} = k[H_2][I_2]$	$1+1=2$
$\text{rate} = k[NO_2]^2$	2
$\text{rate} = k[[H_2][NO]^2$	$1+2=3$
$\text{rate} = k[CHCl_3][Cl_2]^{1/2}$	$1+\frac{1}{2}=1\frac{1}{2}$

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Features of The Order of a Reaction

- *Reaction orders are always determined experimentally.*
- *Order is always defined in terms of reactant (not product) concentrations.*
- *The order of a reactant is not related to the stoichiometric coefficient of the reactant in the balanced chemical equation.*
- *A reaction has an individual order "with respect to" or "in" each reactant.*

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The Order of a Reaction - Problems

Problem Many gaseous reactions occur in car engines and exhaust systems. One of these is

$$NO_2(g) + CO(g) \longrightarrow NO(g) + CO_2(g) \qquad \text{Rate} = k[NO_2]^m[CO]^n$$

(a) Use the following data to determine the individual and overall reaction orders.
(b) Calculate k using the data from experiment 1.

Experiment	Initial Rate (mol/L·s)	Initial $[NO_2]$ (mol/L)	Initial [CO] (mol/L)
1	0.0050	0.10	0.10
2	0.080	0.40	0.10
3	0.0050	0.10	0.20

Solution (a) Calculating m in $[NO_2]^m$: We take the ratio of the rate laws for experiments 1 and 2, in which $[NO_2]$ varies but [CO] is constant:

$$\frac{\text{Rate 2}}{\text{Rate 1}} = \frac{k[NO_2]_2^m[CO]_2^n}{k[NO_2]_1^m[CO]_1^n} = \left(\frac{[NO_2]_2}{[NO_2]_1}\right)^m \quad \text{or} \quad \frac{0.080 \text{ mol/L·s}}{0.0050 \text{ mol/L·s}} = \left(\frac{0.40 \text{ mol/L}}{0.10 \text{ mol/L}}\right)^m$$

This gives $16 = (4.0)^m$, so we have $m = \log 16/\log 4.0 = 2.0$. The reaction is second order in NO_2.

Calculating n in $[CO]^n$: We take the ratio of the rate laws for experiments 1 and 3, in which [CO] varies but $[NO_2]$ is constant:

$$\frac{\text{Rate 3}}{\text{Rate 1}} = \frac{k[NO_2]_3^m[CO]_3^n}{k[NO_2]_1^m[CO]_1^n} = \left(\frac{[CO]_3}{[CO]_1}\right)^n \quad \text{or} \quad \frac{0.0050 \text{ mol/L·s}}{0.0050 \text{ mol/L·s}} = \left(\frac{0.20 \text{ mol/L}}{0.10 \text{ mol/L}}\right)^n$$

We have $1.0 = (2.0)^n$, so $n = 0$. The rate does not change when [CO] varies, so the reaction is zero order in CO.

Therefore, the rate law is

$$\text{Rate} = k[NO_2]^2[CO]^0 = k[NO_2]^2(1) = k[NO_2]^2$$

The reaction is second order overall.

(b) From experiment 1, rate = 0.0050 mol/L·s and $[NO_2]$ = 0.10 mol/L. Calculating k:

$$\text{Rate} = k[NO_2]^2$$

$$k = \frac{\text{rate}}{[NO_2]^2} = \frac{0.0050 \text{ mol/L·s}}{(0.10 \text{ mol/L})^2}$$

$$= 0.50 \text{ L/mol·s}$$

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The Order of a Reaction - Problems

Practice Exercise The reaction of peroxydisulfate ion ($S_2O_8^{2-}$) with iodide ion (I^-) is

$$S_2O_8^{2-}(aq) + 3I^-(aq) \longrightarrow 2SO_4^{2-}(aq) + I_3^-(aq)$$

From the following data collected at a certain temperature, determine the rate law and calculate the rate constant.

Experiment	$[S_2O_8^{2-}]$ (M)	$[I^-]$ (M)	Initial Rate (M/s)
1	0.080	0.034	2.2×10^{-4}
2	0.080	0.017	1.1×10^{-4}
3	0.16	0.017	2.2×10^{-4}

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Molecularity of a Reaction

- *Chemical reactions are classified into two types-*
  - ✓ *Elementary reactions - occurs in a single step*
  - ✓ *Complex reactions - occurs in two or more steps*

- *For elementary reactions, molecularity is defined as the number of molecules involved in the reaction.*
- *Examples:*

Br_2	$\longrightarrow$	$2Br$	Unimolecular
$H_2 + I_2$	$\longrightarrow$	$2HI$	Bimolecular
$2NO + O_2$	$\longrightarrow$	$2NO_2$	Termolecular

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Molecularity of a Complex Reaction

- *Most chemical reactions are complex reactions. These occur in a series of steps.*
- *Each step is an elementary reaction.*
- *Some of the steps will be fast, others will be slow. A reaction can proceed no faster than its slowest step.*
- *Thus the slowest step is the rate determining step of the reaction.*
- *Molecularity of a complex reaction is defined as the number of molecules or atoms taking part in the rate-determining step.*

$$2N_2O_5 \longrightarrow 4NO_2 + O_2$$

Step 1: $2N_2O_5 \longrightarrow 2NO_2 + 2NO_3$; Fast

Step 2: $NO_2 + NO_3 \longrightarrow NO_2 + NO + O_2$; Slow

Step 3: $NO + NO_3 \longrightarrow 2NO_2$; Fast

Thus step 2 in the above mechanism is rate-determining and has molecularity '2' which could be considered as the molecularity of the decomposition reaction of N_2O_5.

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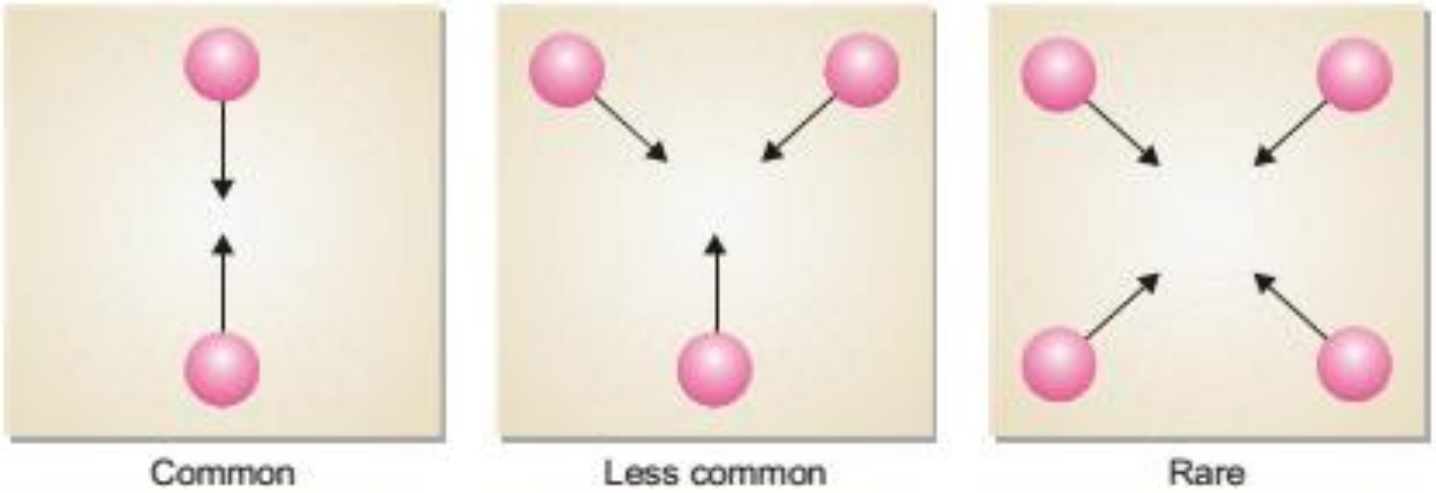
Why High Molecularity Reactions are Rare?

- *The rate of a chemical reaction is proportional to the number of collisions taking place between the reacting molecules.*
- *The chances of simultaneous collision of reacting molecules will go on decreasing with increase in number of molecules.*



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Order and Molecularity of Reaction

Order	Molecularity
Sum of powers of the concentration terms in the rate law expression.	Number of reactant molecular involved in a reaction
It is experimentally determined value.	It is a theoretical concept.
It can have a fractional value.	It is always a whole number.
Order of reaction can be zero.	Molecularity can never be zero.
Order of reaction can change with conditions like temperature, pressure, concentration etc.	For a chemical reaction, molecularity is always same.

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Zero Order Reaction

- A zero order reaction is one whose rate is independent of concentration.
- A reactant whose concentration does not affect the reaction rate is not included in the rate law.
- The concentration of such a reactant has the power 0. Thus $[A]^0 = 1$.
- Let's consider the following zero order reaction,

	$A \longrightarrow$	product
Initial conc.	a	0
Final conc.	$(a-x)$	x

Rate of reaction, $dx/dt = k_0 [A]^0 = k_0 (a-x)^0 = k_0$

$$\text{or, } dx = k_0\, dt$$

$$\text{or, } x = k_0\, t \qquad \text{(by integrating)}$$

$$\text{or, } k_0 = x/t$$

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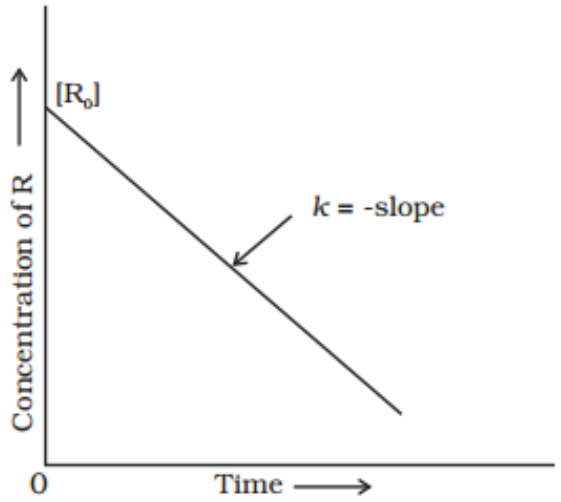
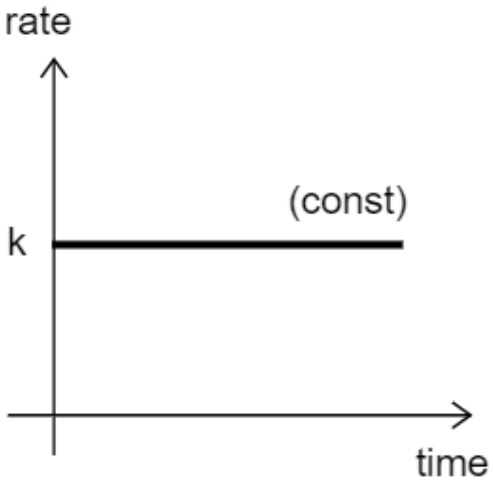
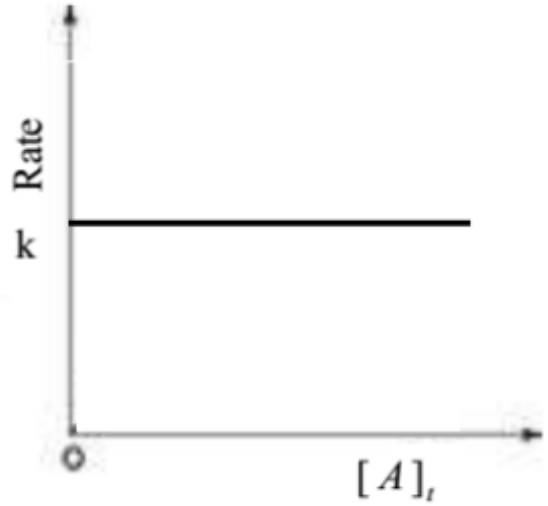
Zero Order Reaction - Graphs







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First Order Reaction

- Let's consider a first order reaction,

$$A \longrightarrow \text{product}$$

Initial conc.	a	0
Final conc.	(a-x)	x

Rate of reaction, $dx/dt = k[A] = k(a-x)$

By integration and mathematical simplification,

$$kt = \ln \frac{[A]_o}{[A]} \qquad kt = \ln \frac{a}{a-x}$$

$$k = \frac{2.303}{t} \log \frac{[A]_o}{[A]} \qquad k = \frac{2.303}{t} \log \frac{a}{(a-x)}$$

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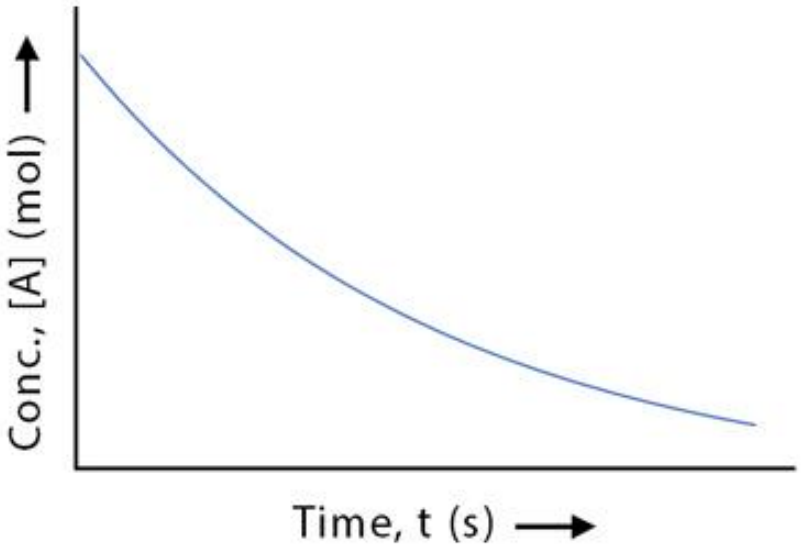
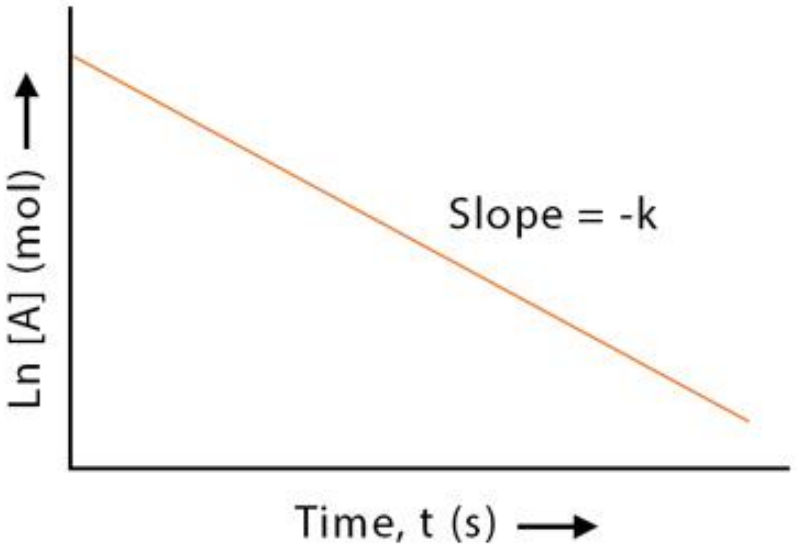
First Order Reaction - Graphs



$$[A] = [A]_o \exp(-kt)$$



$$\ln [A] = \ln [A]_o - kt$$

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First Order Reaction – Half Life

$$t = \frac{2.303}{k} \log \frac{[A]_o}{[A]} \qquad t = \frac{2.303}{k} \log \frac{a}{(a-x)}$$

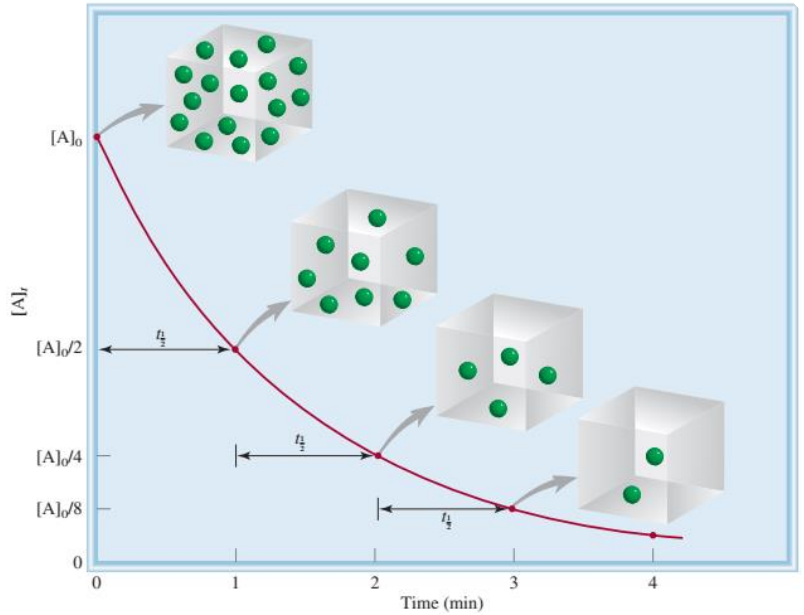
When $t = t_{1/2}$, $[A] = [A]_o/2$ or $x = a/2$; then we can write,

$$t_{1/2} = \frac{2.303}{k} \log \frac{[A]_o}{[A]_o/2} \qquad t_{1/2} = \frac{2.303}{k} \log \frac{a}{(a-a/2)}$$

$$\text{or, } t_{1/2} = \frac{2.303}{k} \log 2 \qquad t_{1/2} = \frac{2.303}{k} \log 2$$

$$\text{or, } t_{1/2} = \frac{0.693}{k}$$

Half life of a first order reaction does not depend on the initial concentration of the reactant



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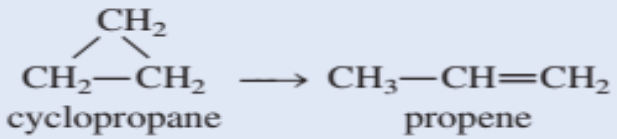
First Order Reaction – Problems

The conversion of cyclopropane to propene in the gas phase is a first-order reaction with a rate constant of $6.7 \times 10^{-4}\ s^{-1}$ at 500°C.



(a) If the initial concentration of cyclopropane was 0.25 *M*, what is the concentration after 8.8 min? (b) How long (in minutes) will it take for the concentration of cyclopropane to decrease from 0.25 *M* to 0.15 *M*? (c) How long (in minutes) will it take to convert 74 percent of the starting material?

Solution (a) In applying Equation (13.4), we note that because k is given in units of s^{-1}, we must first convert 8.8 min to seconds:

$$8.8\ \text{min} \times \frac{60\ \text{s}}{1\ \text{min}} = 528\ \text{s}$$

We write

$$\begin{aligned} \ln [\text{A}]_t &= -kt + \ln [\text{A}]_0 \\ &= -(6.7 \times 10^{-4}\ \text{s}^{-1})(528\ \text{s}) + \ln (0.25) \\ &= -1.74 \end{aligned}$$

Hence, $[\text{A}]_t = e^{-1.74} = 0.18\ M$

(b) Using Equation (13.3),

$$\ln \frac{0.15\ M}{0.25\ M} = -(6.7 \times 10^{-4}\ \text{s}^{-1})t$$

$$t = 7.6 \times 10^2\ \text{s} \times \frac{1\ \text{min}}{60\ \text{s}}$$

$$= 13\ \text{min}$$

(c) From Equation (13.3),

$$\ln \frac{0.26}{1.00} = -(6.7 \times 10^{-4}\ \text{s}^{-1})t$$

$$t = 2.0 \times 10^3\ \text{s} \times \frac{1\ \text{min}}{60\ \text{s}} = 33\ \text{min}$$

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First Order Reaction – Problems

Practice Exercise The reaction $2A \longrightarrow B$ is first order in A with a rate constant of $2.8 \times 10^{-2}\ s^{-1}$ at 80°C. How long (in seconds) will it take for A to decrease from 0.88 *M* to 0.14 *M*?

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First Order Reaction – Problems

The decomposition of ethane (C_2H_6) to methyl radicals is a first-order reaction with a rate constant of $5.36 \times 10^{-4}\ s^{-1}$ at 700°C:

$$C_2H_6(g) \longrightarrow 2CH_3(g)$$

Calculate the half-life of the reaction in minutes.

Solution For a first-order reaction, we only need the rate constant to calculate the half-life of the reaction. From Equation (13.6)

$$\begin{aligned} t_{\frac{1}{2}} &= \frac{0.693}{k} \\ &= \frac{0.693}{5.36 \times 10^{-4}\ s^{-1}} \\ &= 1.29 \times 10^{3}\ s \times \frac{1\ min}{60\ s} \\ &= 21.5\ min \end{aligned}$$

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First Order Reaction – Problems

SOLVED PROBLEM. From the following data for the decomposition of N_2O_5 in CCl_4 solution at 48°C, show that the reaction is of the first order

t (mts)	10	15	20	∞
Vol of O_2 evolved	6.30	8.95	11.40	34.75

SOLUTION

For a first order reaction the integrated rate equation is

$$\frac{1}{t}\log\frac{V_\infty}{V_\infty - V_t} = k$$

In this example, $V_\infty = 34.75$

t	$V_\infty - Vt$	$\frac{1}{t}\log\frac{V_\infty}{V_\infty - V_t}$	=	k
10	28.45	$\frac{1}{10}\log\frac{34.75}{28.45}$	=	0.00868
15	25.80	$\frac{1}{15}\log\frac{34.75}{25.80}$	=	0.00862
20	23.35	$\frac{1}{20}\log\frac{34.75}{23.35}$	=	0.00863

Since the value of k is fairly constant, it is a **first order reaction**.

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First Order Reaction – Problems

SOLVED PROBLEM 5. For a certain first order reaction $t_{0.5}$ is 100 sec. How long will it take for the reaction to be completed 75% ?

SOLUTION

Calculation of k

For a first order reaction

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

or $$100 = \frac{0.693}{k}$$

$$\therefore \quad k = \frac{0.693}{100} = 0.00693 \text{ sec}^{-1}$$

Calculation of time for 75% completion of reaction

The integrated rate equation for a first order reaction is

$$k = \frac{2.303}{t} \log \frac{[A]_0}{[A]}$$

or $$t = \frac{2.303}{k} \log \frac{[A]_0}{[A]}$$

When $\frac{3}{4}$ initial concentration has reacted, it is reduced to $\frac{1}{4}$

Substituting values in the rate equation

$$t_{3/4} = \frac{2.303}{0.00693} \log \frac{[A]_0}{\frac{1}{4}[A]_0}$$

$$= \frac{2.303}{0.00693} \log 4 = \mathbf{200\,sec}$$

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First Order Reaction – Problems

SOLVED PROBLEM 6. A first order reaction is one-fifth completed in 40 minutes. Calculate the time required for its 100% completion.

SOLUTION

Calculation of k

For a first order reaction

$$k = \frac{2.303}{t} \log \frac{[A]_0}{[A]}$$

After 40 mts, the initial concentration is reduced to $\frac{4}{5}$ That is,

$$[A] = \frac{4}{5}[A]_0$$

Substituting values in the equation above

$$k = \frac{2.303}{40} \log \frac{[A]_0}{\frac{4}{5}[A]_0}$$

or

$$k = \frac{2.303}{40} \log 5 - \log 4 = 0.00558\,\text{mt}^{-1}$$

Calculation of time required for 100% completion

We know that for first order reaction

$$k = \frac{2.303}{t} \log \frac{[A]_0}{[A]}$$

$$t = \frac{2.303}{k} \log \frac{[A]_0}{[A]}$$

If reaction is 100% complete in, say, t_1 time, we have, $[A] = 0$. Thus,

$$t_1 = \frac{2.303}{0.00558} \log \frac{[A]_0}{0} = \infty$$

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Second Order Reaction

- Let's consider a second order reaction,

$$2A \longrightarrow \text{product}$$

Rate of reaction, $dx/dt = k\,(a-x)^2$

$$\text{or, } \int \frac{dx}{(a-x)2} = \int k\,dt \text{ or, } \frac{1}{a-x} = kt + I \text{ -------- (1)}$$

When $t = 0$, $x = 0$; therefore, $I = 1/a$

Then, equation (1) will be, $\frac{1}{a-x} = kt + \frac{1}{a}$

$$\text{or, } k = \frac{1}{t}\,\frac{x}{a(a-x)} \text{ ---------- (2)}$$

When $t = t_{1/2}$, $x = a/2$; then equation (2) becomes –

$$t_{1/2} = \frac{1}{k\,a}$$

While half-life of a first order reaction is independent of initial concentration, half-life of a second order reaction depends on initial concentration.

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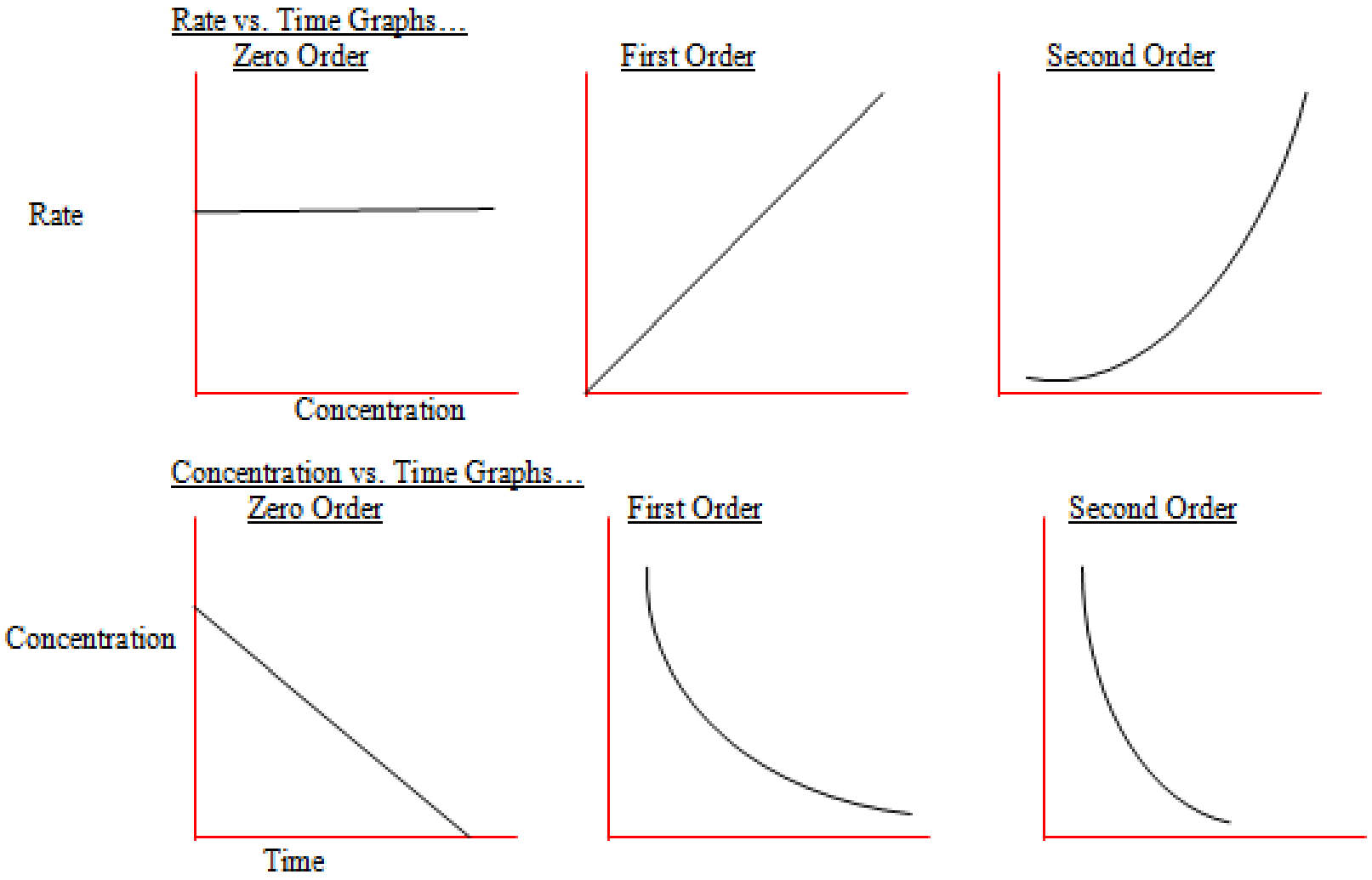
Second Order Reaction



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Pseudo-Order Reaction

- *A reaction in which one of the reactants is present in a large excess shows an order different from the actual order.*
- *The experimental order which is not the actual one is referred to as the pseudo-order.*
- *Let's consider a reaction,* $A + B \longrightarrow$ *product*

Then, rate $= k[A][B]$

If B is present in large extent, its concentration practically remains unchanged over the course of the reaction. Therefore,

rate $= k'[A]$ *where,* $k' = k[B]$

Thus, the actual order of the reaction was second-order, but in practice it is first order reaction. Hence it is said to be pseudo first order reaction.

Example: $CH_3COOC_2H_5 + H_2O \longrightarrow CH_3COOH + C_2H_5OH$

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Units of rate constant

- *Zero order reaction*

$$k = dx/dt$$

Unit: mole L^{-1} $time^{-1}$

1st order reaction

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

Unit: $time^{-1}$

2nd order reaction

$$k = \frac{1}{t} \frac{x}{a(a-x)}$$

Unit: L $mole^{-1}$ $time^{-1}$

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Determination of Order of a Reaction

- *Using integrated rate equation*
- *Graphical method*
- *Using half-life*
- *Differential method*
- *Ostwald's isolation method*

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Determination of Order of a Reaction

❑ *Using integrated rate equation*

- *For a particular reaction, A ⟶ product, 'a' is the initial concentration, (a-x) is the concentration at time t.*
- *The experimental value of a, (a-x) and t for different time interval are then substituted in the integrated rate equation for zero, 1st, 2nd or 3rd order reactions.*
- *The rate equation which yields a constant value of k corresponds to the correct order of the reaction.*
- *This method of ascertaining the order of a reaction is essentially a method of hit-and-trial but still used extensively to find the order of simple reactions.*

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Determination of Order of a Reaction

❑ Graphical Method

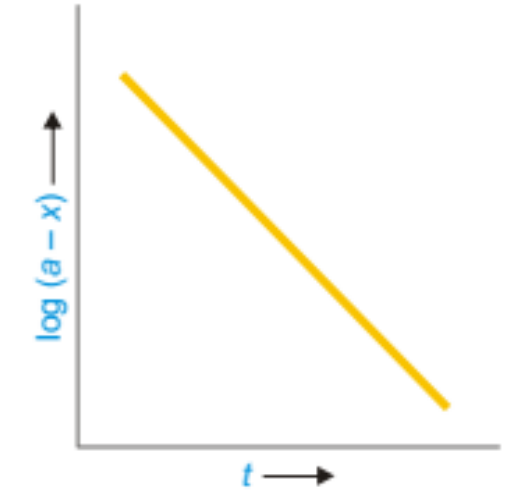
- For 1st order reaction, the rate equation is –

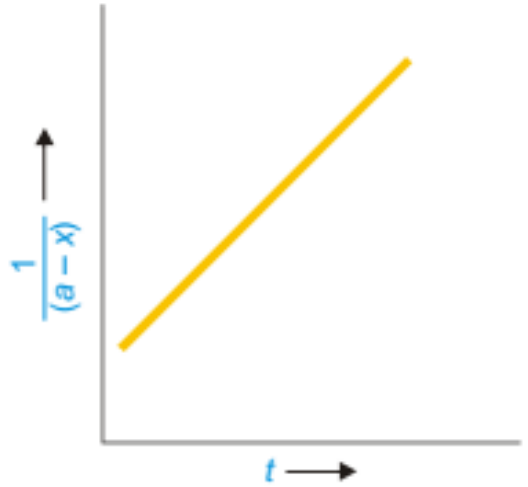
$$\ln a/(a-x) = kt$$

$$\text{or, } \ln (a-x) = -kt + \ln a$$

For 2nd order reaction,

$$1/(a-x) = kt + 1/a$$





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Determination of Order of a Reaction

❑ Using Half-life

$$\text{Half life} \propto \frac{1}{[A]} \text{ for 2nd order reaction}$$

Similarly, Half life $\propto \frac{1}{[A]^2}$ for 3rd order; Half life $\propto \frac{1}{[A]^{n-1}}$ for nth order

Let the initial concentrations in the two experiments be $[A_1]$ and $[A_2]$, while times for completion of half change are t_1 and t_2 respectively.

Then, $t_1 \propto \frac{1}{[A_1]^{n-1}}$ and $t_2 \propto \frac{1}{[A_2]^{n-1}}$

$\frac{t_2}{t_1} = [\frac{A_1}{A_2}]^{n-1}$; or, $(n-1) \log [\frac{A_1}{A_2}] = \frac{t_1}{t_2}$; or, $n = 1 + \frac{\log (\frac{t_1}{t_2})}{\log [\frac{A_1}{A_2}]}$

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Determination of Order of a Reaction

❑ The Differential Method

The rate of a reaction of nth order is, $-\frac{dc}{dt} = k C^n$

Where, C is the concentration at any instant. For two reactions performed with two initial concentrations, we can write,

$$-\frac{dc_1}{dt} = k C_1^n \text{ and } -\frac{dc_2}{dt} = k C_2^n$$

Taking logs for both cases, $\log\left(-\frac{dc_1}{dt}\right) = \log k + n \log C_1$ ---------- (1)

and $\log\left(-\frac{dc_2}{dt}\right) = \log k + n \log C_2$ ----------(2)

Subtracting (2) from (1), we have,

$$n = \frac{\log\left(-\frac{dc_1}{dt}\right) - \log\left(-\frac{dc_2}{dt}\right)}{\log C_1 - \log C_2}$$

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Determination of Order of a Reaction

❑ *Ostwald's Isolation Method*

- *This method is employed in determining the order of complicated reactions by 'isolating' one of the reactants so far as its influence on the rate of reaction is concerned.*
- *Let's consider a reaction ,* $A + B + C \rightarrow$ *product*
- *The order of the reaction with respect to A, B and C is determined.*
- *For the determination of the order of reaction with respect to A, B and C are taken in a large excess so that their concentrations are not affected during the reaction.*
- *The order of the reaction is then determined by using any of the methods described earlier.*
- *Likewise, the order of the reaction with respect to B and C is determined.*
- *Therefore,* $n = n_A + n_B + n_C$

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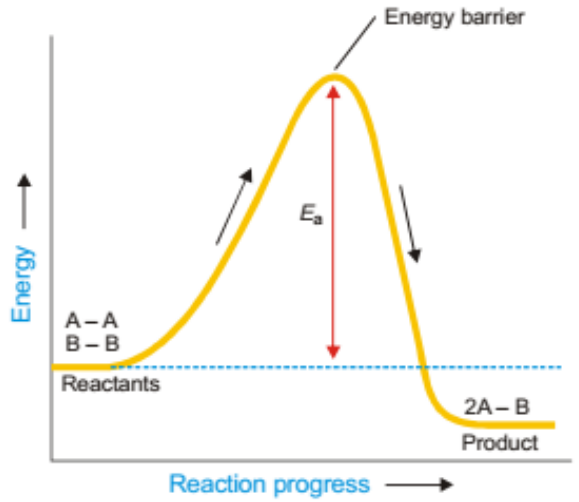
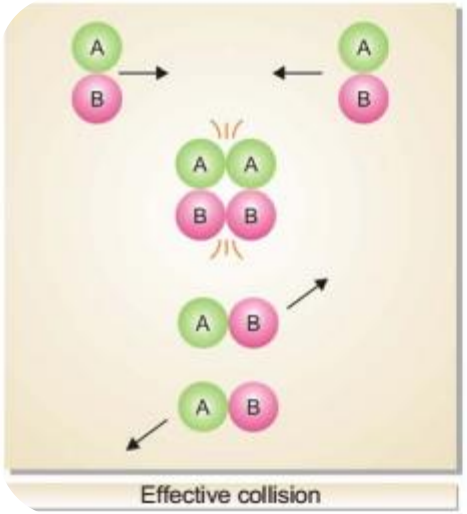
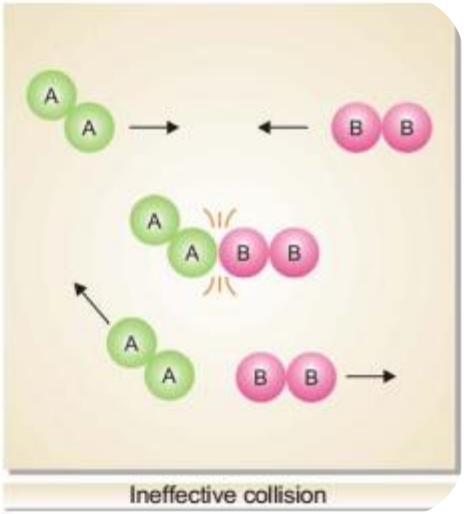
Collision Theory of Reaction Rates

- *A chemical reaction takes place only by collisions between the reacting molecules.*
- *The two main conditions for a collision between the reacting molecules to be productive are –*
  - *The colliding molecules must posses sufficient kinetic energy to cause a reaction.*
  - *The reacting molecules must collide with proper orientation.*







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Temperature and Rate Constant

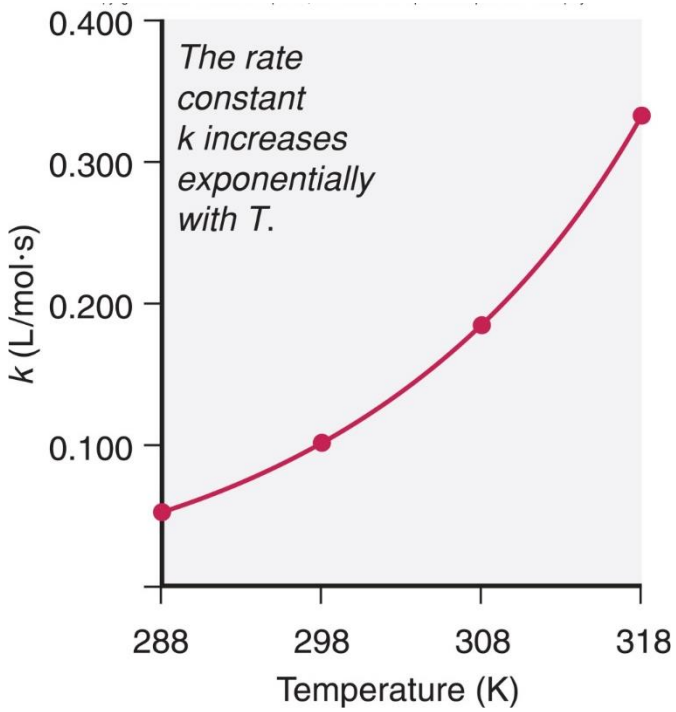
- *Temperature has a dramatic effect on reaction rate. For many reactions, an increase of 10°C will double or triple the rate.*
- *Experimental data shows that k increases exponentially as T increases. This is expressed in the Arrhenius equation:*

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

k = rate constant, A = constant, E_a = activation energy.

Higher *T* ⟶ larger *k* ⟶ increased rate

Reaction rate and k increase exponentially as T increases.



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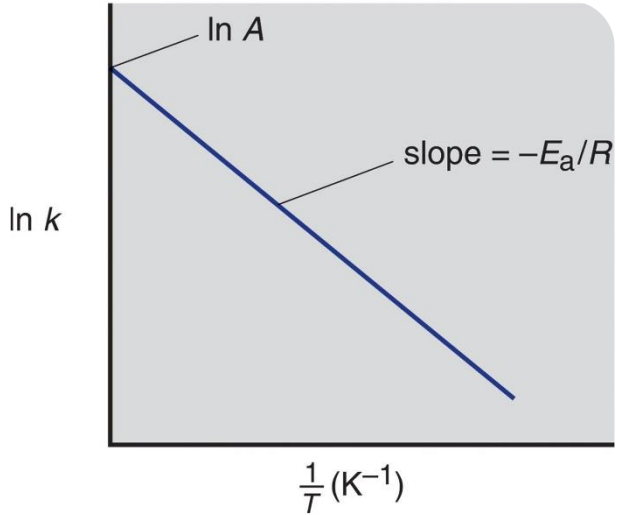
Temperature and Rate Constant

- E_a can be calculated from the Arrhenius equation.

$$k = Ae^{-E_a/RT} \qquad \ln k = \ln A \frac{-E_a}{R}\left(\frac{1}{T}\right)$$

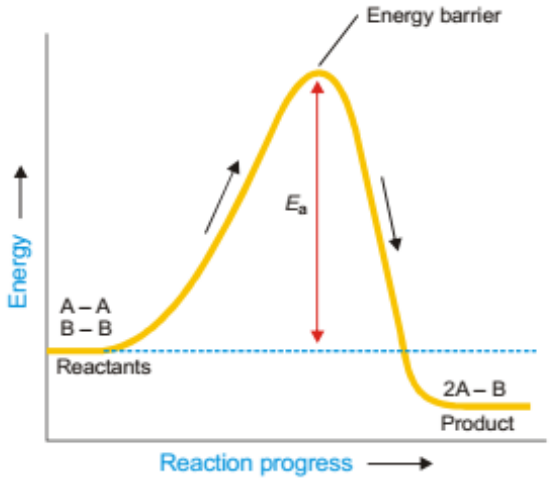
straight-line form



- If k_1 and k_2 are rate constants at temperature T_1 and T_2, then the above equation will be -

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

- Activation energy (E_a) is defined as the minimum amount of energy required by a reacting molecule to get converted into product.



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Temperature and Rate Constant - Problem

- The decomposition of hydrogen iodide, $2HI(g) \rightarrow H_2(g) + I_2(g)$, has rate constants of 9.51×10^{-9} L/mol·s at 500 K and 1.10×10^{-5} L/mol·s at 600 K. Find E_a.
- Solution

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

Ans: $= 1.76 \times 10^5$ J/mol $= 1.76 \times 10^2$ kJ/mol

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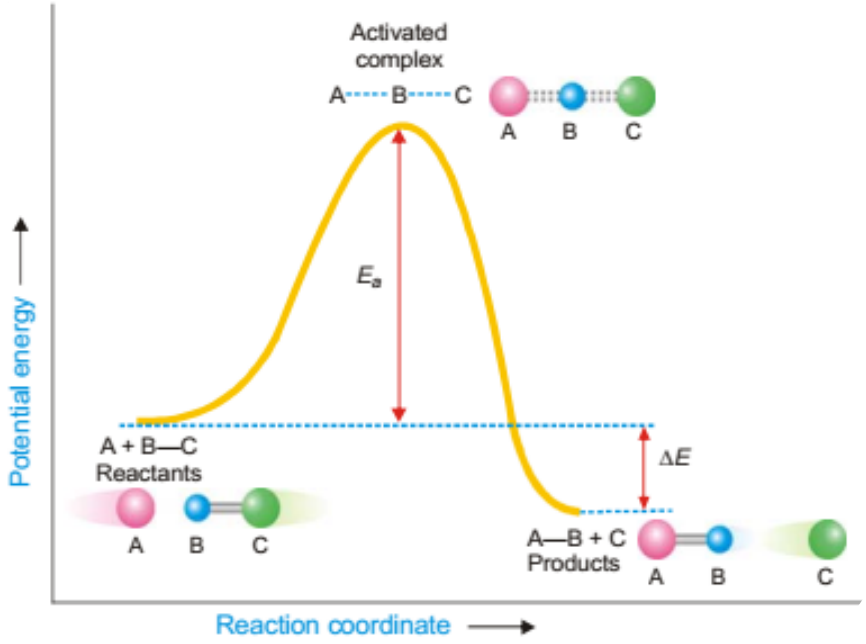
Transition State Theory

- *An effective collision between particles leads to the formation of a transition state or activated complex.*
- *During the collision, the reactant molecules form a transition state or activated complex which decomposes to give the products.*
- *The transition state is an unstable species that contains partial bonds.*
- *It's a transitional pathway between reactants and products.*
- *Transition states cannot be isolated.*



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Thank You

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