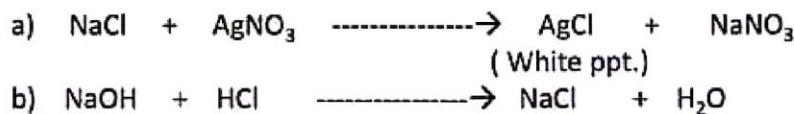


Chemical Kinetics

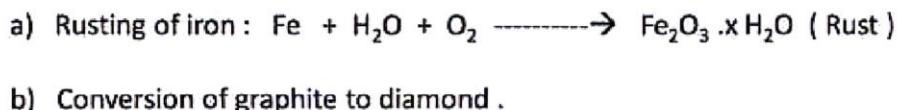
Chemical Kinetics :

Some reactions take place very fast and some reactions take place very slow. While some reactions take place with moderate rate (speed). So different reactions occur with different rate . Let us consider following reactions i.e.

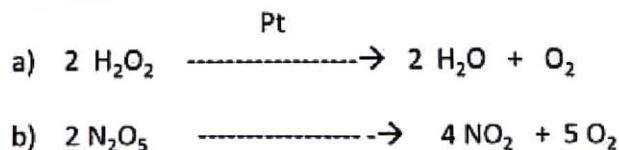
1) Very fast reaction :



2) Very slow reaction :



3) Moderate reaction :



From the above reaction, it is clear that different reactions occur with different rate.

The branch of physical chemistry that deals with the study of rate of chemical reaction and factors that affect the rate of chemical reaction is called chemical kinetics. Sometimes , it also deals about mechanism of the reaction .

Rate of chemical reaction: [2 Marks]

It is defined as the decrease in concentration of reactant per unit time or increase in concentration of product per unit time or change in concentration of either reactant or product per unit time.

Mathematically ,

$$\text{Rate of reaction} = \frac{\text{Change in concentration of reactant or product}}{\text{Change in time}}$$

Let us consider the following reaction i. e. $\text{A} \longrightarrow \text{B}$

Let the concentration of reactant and product be $[\text{A}_1]$ and $[\text{B}_1]$ respectively at time t_1 , and the concentration be $[\text{A}_2]$ and $[\text{B}_2]$ respectively at time t_2 , then the rate of reaction in terms of reactant can be written as :

$$\text{Rate of reaction} = \frac{\text{Change in concentration of reactant}}{\text{change in time}}$$

$$= [A_2] - [A_1]/t_2 - t_1$$

$$= \Delta [A]/\Delta t$$

The expression $[A_2] - [A_1]$ is a negative quantity but the rate of reaction can't be negative. So to make the rate of reaction +ve, the expression $[A_2] - [A_1]$ must be multiplied with -ve sign. Hence, the rate of reaction in terms of reactant must be written as follows:

$$\text{Rate of reaction} = - \Delta [A]/\Delta t$$

Where -ve sign indicates the decrease in concentration of reactant.

The rate of reaction in terms of product can be written as follows :

$$\text{Rate of reaction} = \text{Change in concentration of product} / \text{Change in time}$$

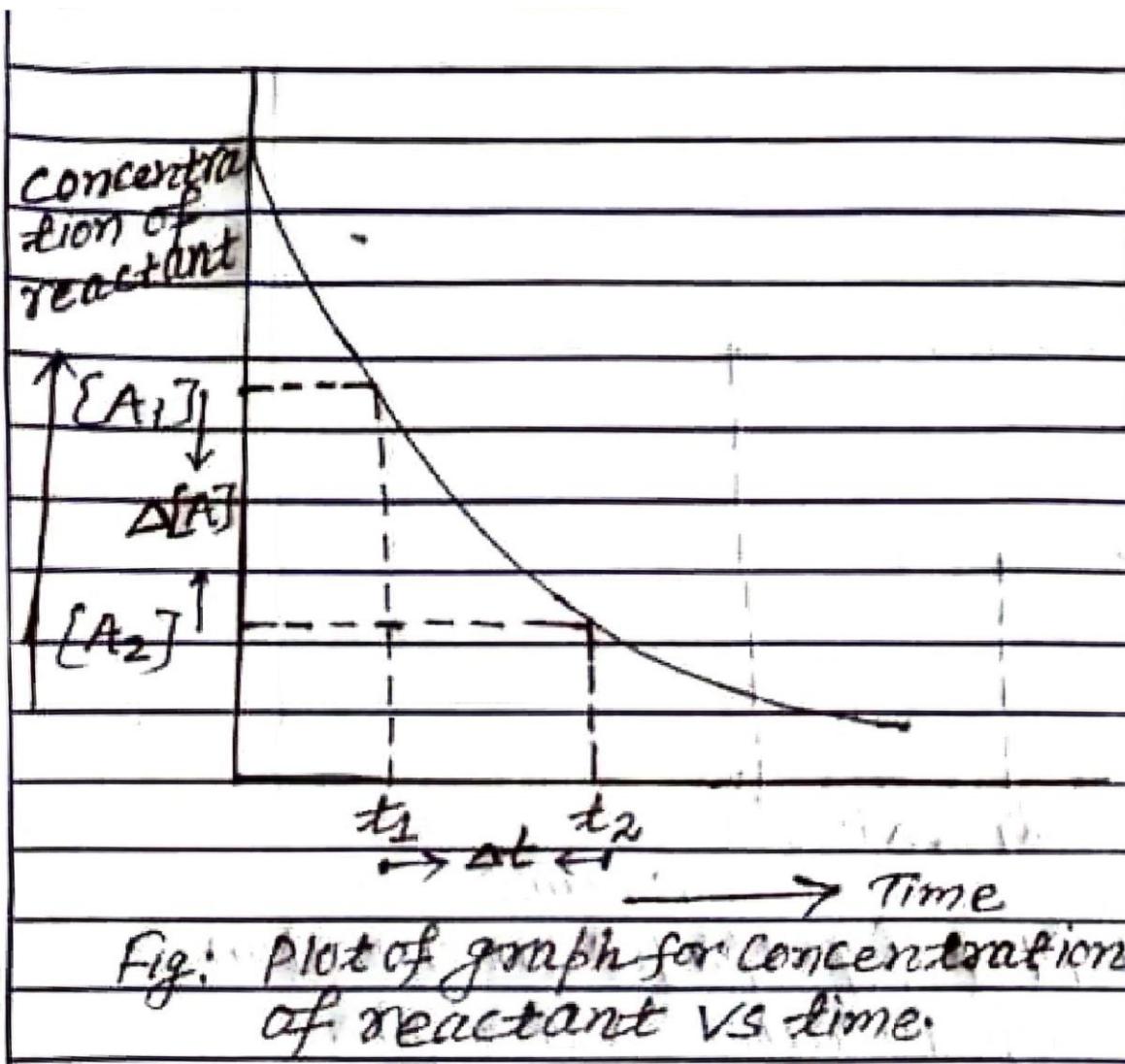
$$= [B_2] - [B_1]/t_2 - t_1$$

$$= \Delta [B]/\Delta t$$

Where +ve sign indicates the increase in concentration of product .

$$\text{Unit of rate of reaction} = \text{Unit of concentration} / \text{Unit of time}$$

$$= \text{mol L}^{-1} \text{time}^{-1} = \text{mol L}^{-1} \text{s}^{-1} \text{ or mol L}^{-1} \text{min}^{-1} \text{ or mol L}^{-1} \text{hr}^{-1}$$



The plot of graph for concentration of product vs time is given below :

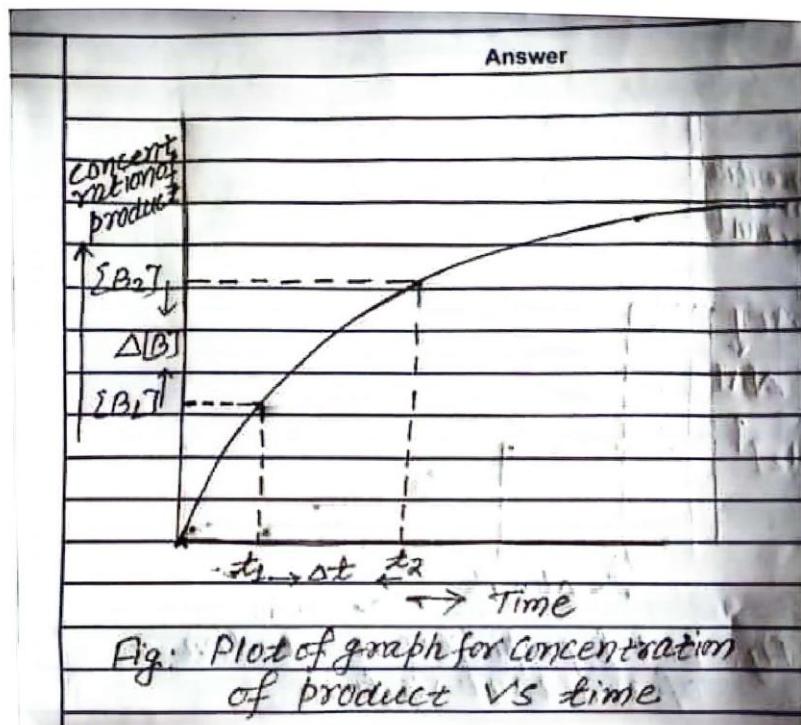


Fig: Plot of graph for concentration of product VS time

Imp. The rate of reaction can be classified into following types : [2 Marks]

1) Average rate of reaction

2) Instantaneous rate of reaction

Average rate of reaction :

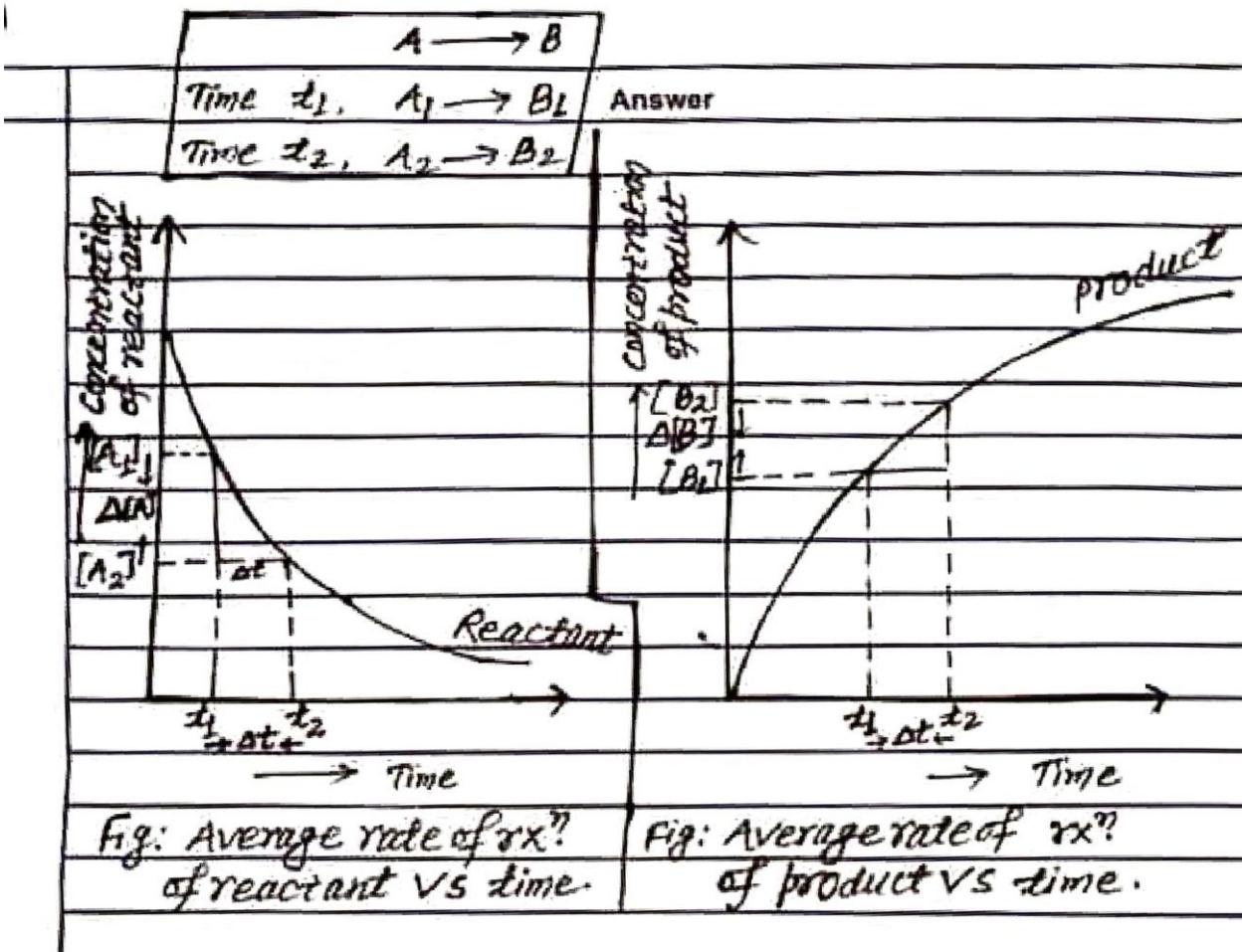
The rate of reaction during certain interval of time is called average rate of reaction .

[Or It can be defined as the rate of change of concentration of either reactant or product at a particular interval of time .] . During average rate of reaction, change in time (Δt) is finite and measurable.

The expression for average rate of reaction can be written as follows :

$$\text{Average rate of reaction} = - \Delta[A] / \Delta t = + \Delta B / \Delta t$$

The plot of average rate of reaction of reactant vs time and the average rate of reaction of product vs time is given below :



Instantaneous rate of reaction :

The rate of reaction during particular instant of time is called instantaneous rate of reaction . [Or It can be defined as the rate of change of concentration of either reactant or product at a particular instant of time] . During instantaneous rate of reaction , change in time (Δt) is so small that it tends to zero (0) . i.e. It is non finite and non measurable. Expression for instantaneous rate of reaction can be written as follows:

$$\text{i.e. Instantaneous rate of reaction} = \lim_{\Delta t \rightarrow 0} - \frac{\Delta [A]}{\Delta t} = - \frac{d[A]}{dt} \text{ [for reactant]}$$

$$\Delta t \rightarrow 0$$

$$= \lim_{\Delta t \rightarrow 0} \frac{\Delta [B]}{\Delta t} = \frac{d[B]}{dt} \text{ [for product]}$$

Also ,

$$\text{Instantaneous rate of reaction} = [\text{Average rate}] \Delta t \rightarrow 0$$

Note : Instantaneous rate of reaction is preferred over average rate of reaction because it gives the rate of reaction at a particular instant of time not at the particular interval of time .

Instantaneous rate of reaction can be calculated graphically from the slope of a tangent drawn at the concentration curve for the particular time as shown in following graph:

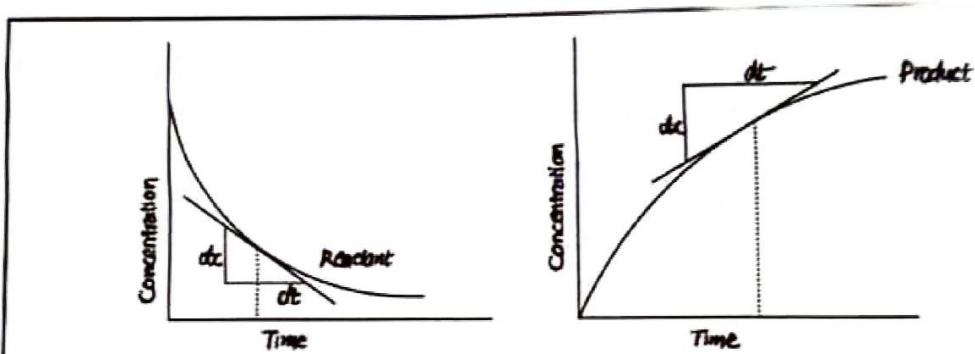
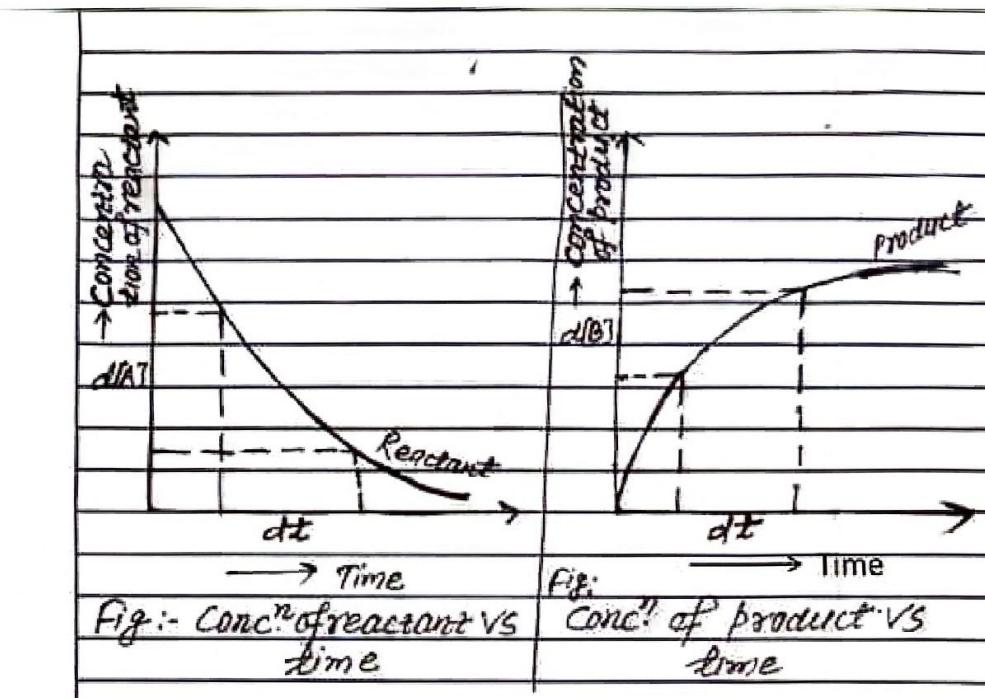


Fig. ■■: Diagram to show the instantaneous rate of reaction
The plot of graph for instantaneous rate of reaction is shown below :



Rate of reaction and stoichiometric coefficient:

Let us consider a chemical reaction , i.e. $H_2 + I_2 \rightarrow 2 HI$
 Here, 1 mole of H_2 and 1 mole of I_2 are consumed to form 2 moles of HI . The consumption and production takes same(equal) time. For this reaction , three different rate expressions are possible which are given below:

$$\text{Rate of disappearance of } H_2 = - \Delta [H_2] / \Delta t = - d[H_2] / dt$$

$$\text{Rate of disappearance of } I_2 = - \Delta [I_2] / \Delta t = - d[I_2] / dt$$

$$\text{Rate of formation of } HI = + \Delta [HI] / \Delta t = + d[HI] / dt$$

While writing the rate expression, the name of species must be necessarily specified. But the rate expression can also be written without specifying the name of species. To do so, the rate expression must be divided by the stoichiometric coefficient of the species present in balanced chemical equation. Such type of rate expression is called equivalent rate expression . (or reaction velocity).

The equivalent rate expression for above reaction can be written as follows:

$$\begin{aligned} \text{i. e. Equivalent rate} &= - \Delta [H_2] / \Delta t = - \Delta [I_2] / \Delta t = + \Delta [HI] / 2 \Delta t \\ &= - d[H_2] / dt = - d[I_2] / dt = + d[HI] / 2 dt \end{aligned}$$

If the reaction is ,



Then , the possible rate expression and equivalent rate expression can be written as follows :

$$\text{Rate of disappearance of } A = -d[A]/dt$$

$$\text{Rate of disappearance of } B = -d[B]/dt$$

$$\text{Rate of formation of } C = +d[C]/dt$$

$$\text{Rate of formation of } D = +d[D]/dt$$

Also,

$$\text{Equivalent rate} = -d[A]/a dt = -d[B]/b dt = +d[C]/c dt = +d[D]/d dt$$

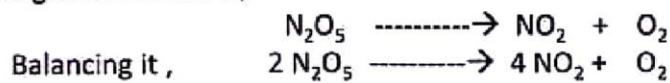
Imp. If the reaction is,



Write the possible rate expression. Also write the equivalent rate expression .

Solⁿ:

Here, the given reaction is,



The possible rate of expression of above reaction can be written as follows:

$$\text{Rate of disappearance of } N_2O_5 = -d[N_2O_5]/dt$$

$$\text{Rate of formation of } NO_2 = +d[NO_2]/dt$$

$$\text{Rate of formation of } O_2 = +d[O_2]/dt$$

Also,

$$\text{Equivalent rate} = -d[N_2O_5]/2 dt = +d[NO_2]/4 dt = +d[O_2]/dt$$

Write the rate of reaction for the given equation and also write the rate of disappearance and rate of formation in respective species.



$$\text{Ans , rate of reaction} = -d[N_2]/dt = -d[H_2]/3 dt = +d[NH_3]/2 dt$$

Now ,

$$\text{Rate of disappearance of } \text{N}_2 = - d[\text{N}_2] / dt$$

$$\text{Rate of disappearance of } \text{H}_2 = - d[\text{H}_2] / dt$$

$$\text{Rate of formation of } \text{NH}_3 = + d[\text{NH}_3] / dt$$



If the rate formation of I_2 is $9.1 * 10^{-6}$ mol lit $^{-1}$ sec $^{-1}$. What will be the rate of disappearance of HI ?

Soln. The given reaction is, $2 \text{HI} \longrightarrow \text{H}_2 + \text{I}_2$

$$\text{Rate of formation of } \text{I}_2 = + d[\text{I}_2] / dt = 9.1 * 10^{-6} \text{ mol lit}^{-1} \text{ sec}^{-1}$$

$$\text{Rate of disappearance of HI} = - d[\text{HI}] / dt = ?$$

$$\text{We have, rate of reaction} = - d[\text{HI}] / 2 dt = + d[\text{H}_2] / dt = + d[\text{I}_2] / dt$$

$$\text{Taking 1}^{\text{st}} \text{ and 3}^{\text{rd}} \text{ ratios, } - d[\text{HI}] / 2 dt = + d[\text{I}_2] / dt$$

$$\text{Or } - d[\text{HI}] / dt = 2 * 9.1 * 10^{-6} \text{ mol lit}^{-1} \text{ sec}^{-1} = 1.82 * 10^{-5} \text{ mol lit}^{-1} \text{ sec}^{-1}$$

$$\text{Therefore, rate of disappearance of HI} = 1.82 * 10^{-5} \text{ mol lit}^{-1} \text{ sec}^{-1}$$

Energy profile diagram :

It is a graphical plot between energy and progress of reaction . The slower or faster reaction rate can be explained by energy profile diagram .

According to this diagram , reactant changes into product only if it can cross energy barrier or hill which is called as activation energy. If the magnitude of activation energy is small , the reaction becomes fast and vice – versa . Energy profile diagram is plotted below :

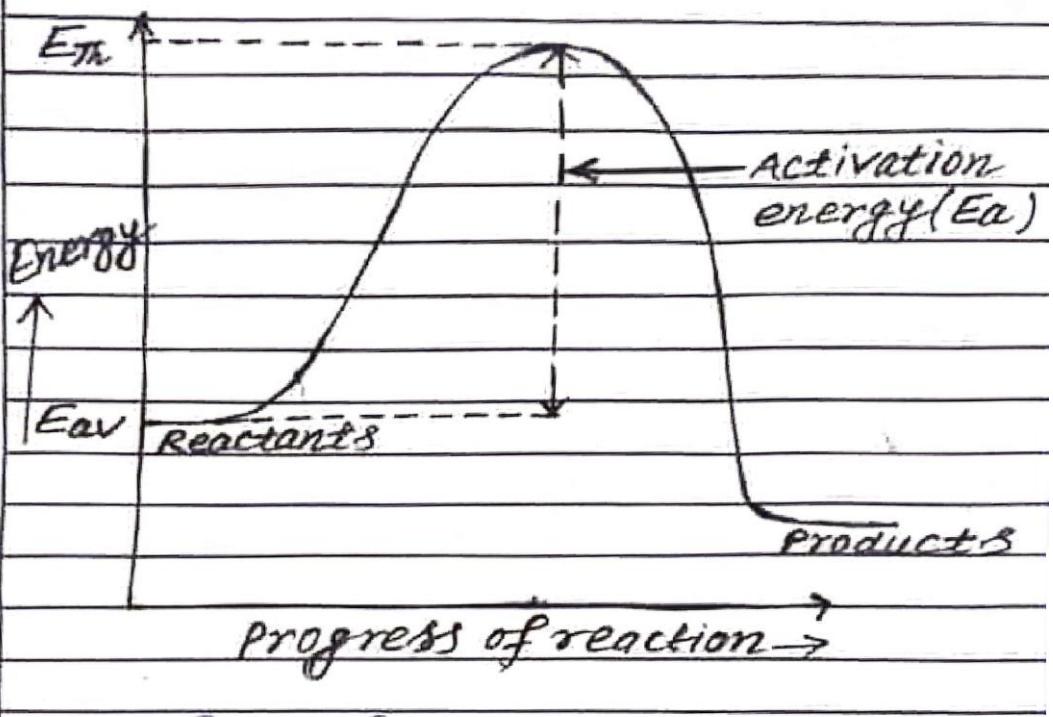
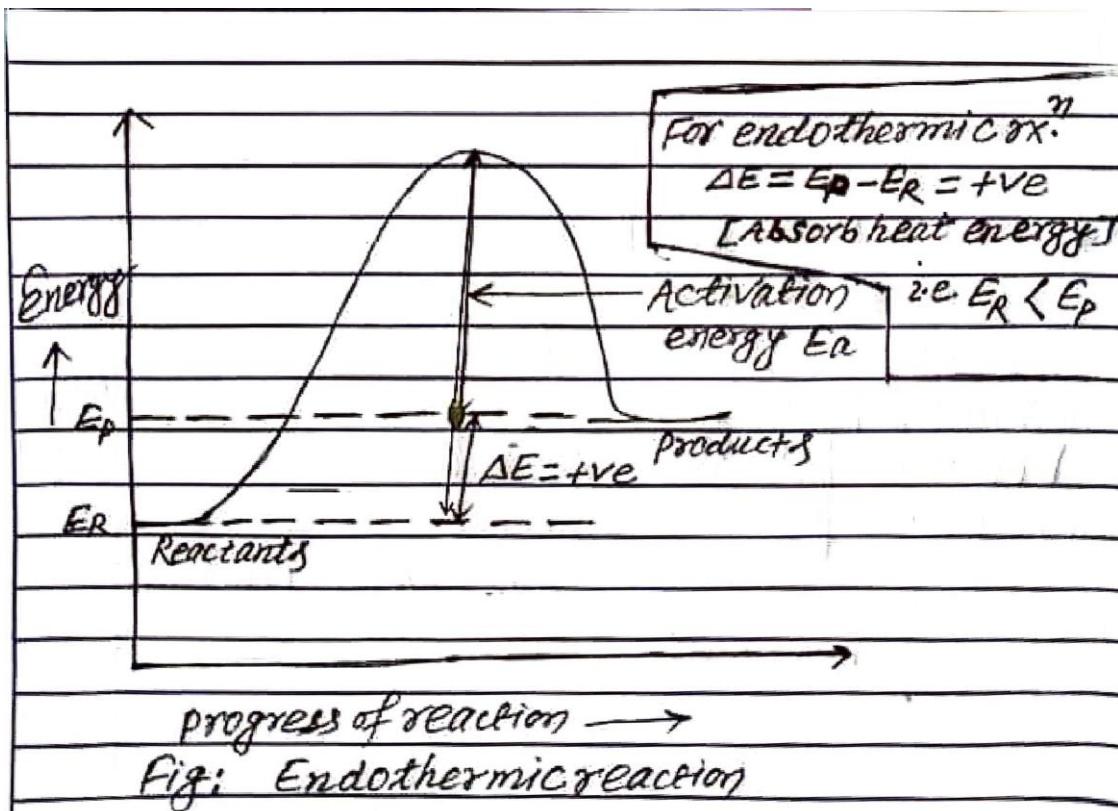
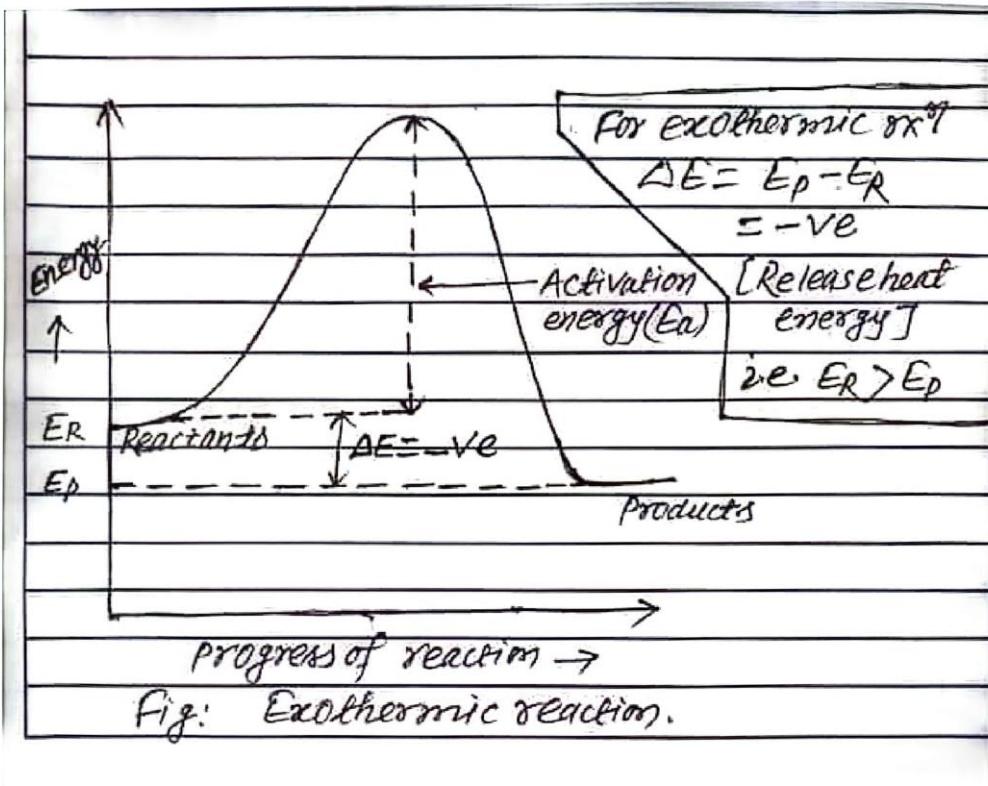


Fig. Energy profile diagram

- a) **Activation energy :** It is the minimum additional amount of energy required by the reactant to cross the energy barrier in order to form the product. Greater the magnitude of activation energy, slower will be the reaction rate and vice – versa. It is denoted by E_a or E_{act} .
- b) **Average energy :** It is the energy stored in a substance due to its chemical nature. It is denoted by E_{av} . Greater the magnitude of average energy, lesser will be the stability and vice – versa.
- c) **Threshold energy :** The minimum amount of energy that reactant molecule need to have in order for reaction to take place is called threshold energy. [Or It is the energy required by the substance to undergo chemical reaction]. It is the sum of average energy and activation energy . It is denoted by E_{th} .

$$\text{i.e. } E_{th} = E_{av} + E_a$$

Energy profile diagram for exothermic and Endothermic reaction :



V imp.

Factors affecting reaction rate :

The various factors that effect reaction rate are discussed below :

1) Concentration of reactants :

Generally , the rate of reaction increases with increase in concentration of reactant . When concentration of reactant is increased , the no. of reactant molecules present per unit volume increases . This causes the increase in collision frequency and hence, the rate of effective collision increases . So , the rate of formation of product increases .

This fact can also be supported from law of mass action , which states that " the rate of forward reaction is directly proportional to the active mass of reactant."

2) Temperature of system :

When the temperature of system is increased , the rate of reaction generally increased . When temperature is increased , kinetic energy (K.E.) of reactant increased which causes the increases in rate of vibration . This causes the increase in collision frequency and hence the rate of effective collision increases. So, the rate of formation of product increases . It is estimated that the reaction rate increases by 2-3 times for every 10°C rise in temperature .

This fact can also be supported from Maxwell's distribution of energy curve which is plotted below :

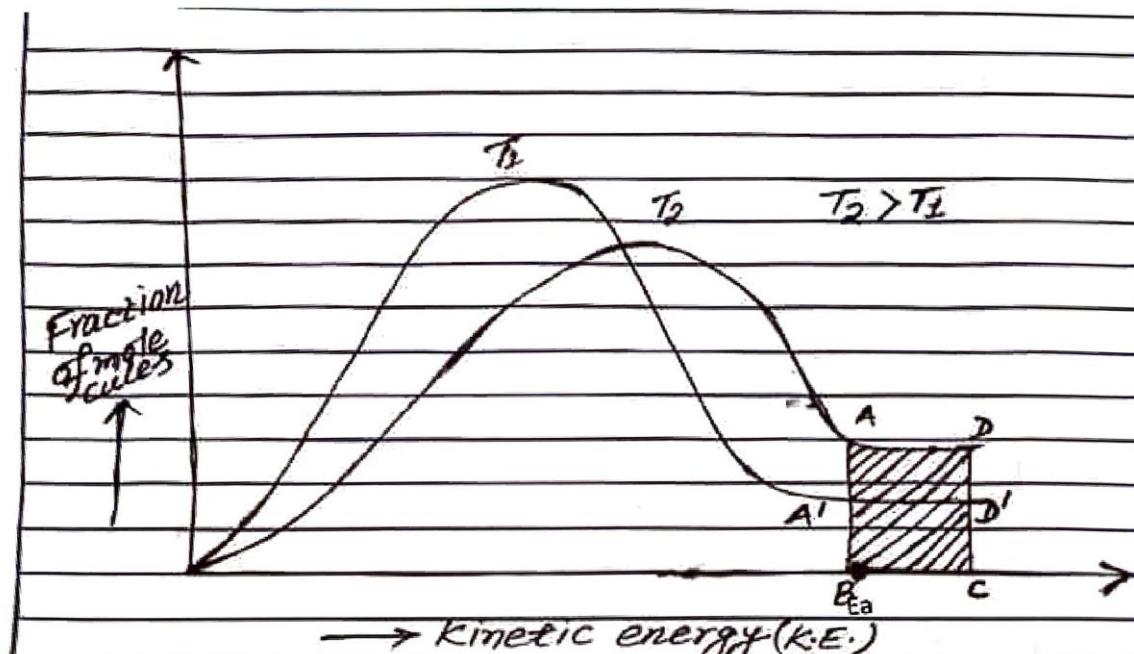


Fig: Maxwell distribution of energy curve

At lower temperature ' T_1 ' , fraction of molecules that can cross energy barrier is denoted by shaded area $A' B C D'$. At higher temperature ' T_2 ' , fraction of molecules that can cross energy barrier is denoted by shaded area $A B C D$. Since area of $A B C D$ is greater than that of $A' B C D'$, greater fraction of reactant molecules can cross energy barrier at higher temperature . It can also be explained according to Arrhenius equation which is discussed later.

3) Catalyst :

It is a third substance used in chemical reaction which alters the rate of reaction without itself being consumed by the end of reaction . Catalyst provides an alternate path for the reaction to occur in which magnitude of activation energy differs as compared to original path . So the rate of reaction changes . There are two types of catalyst i. e. positive catalyst and negative catalyst .

Positive catalyst provides an alternate path for the reaction to occur in which magnitude of activation energy becomes lower as compared to original path. So that more reactants can cross the energy barrier easily and it increases the rate of reaction .

On the other hand , a negative catalyst provides an alternate path for the reaction to occur in which magnitude of activation energy becomes higher as compared to original path. So that less reactants can cross the energy barrier and it decreases the rate of reaction .

The energy profile diagram for catalyzed and uncatalysed reaction is plotted below :

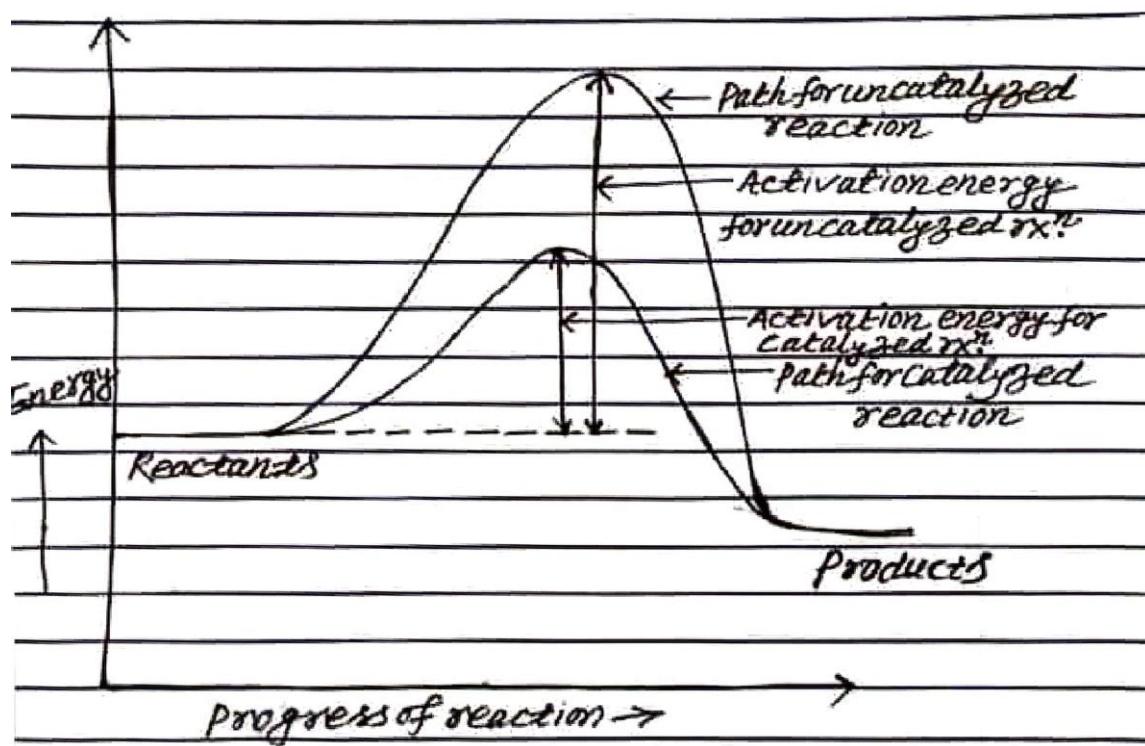


Fig: Energy profile diagram for catalyzed and uncatalyzed reaction. [Use +ve catalyst]

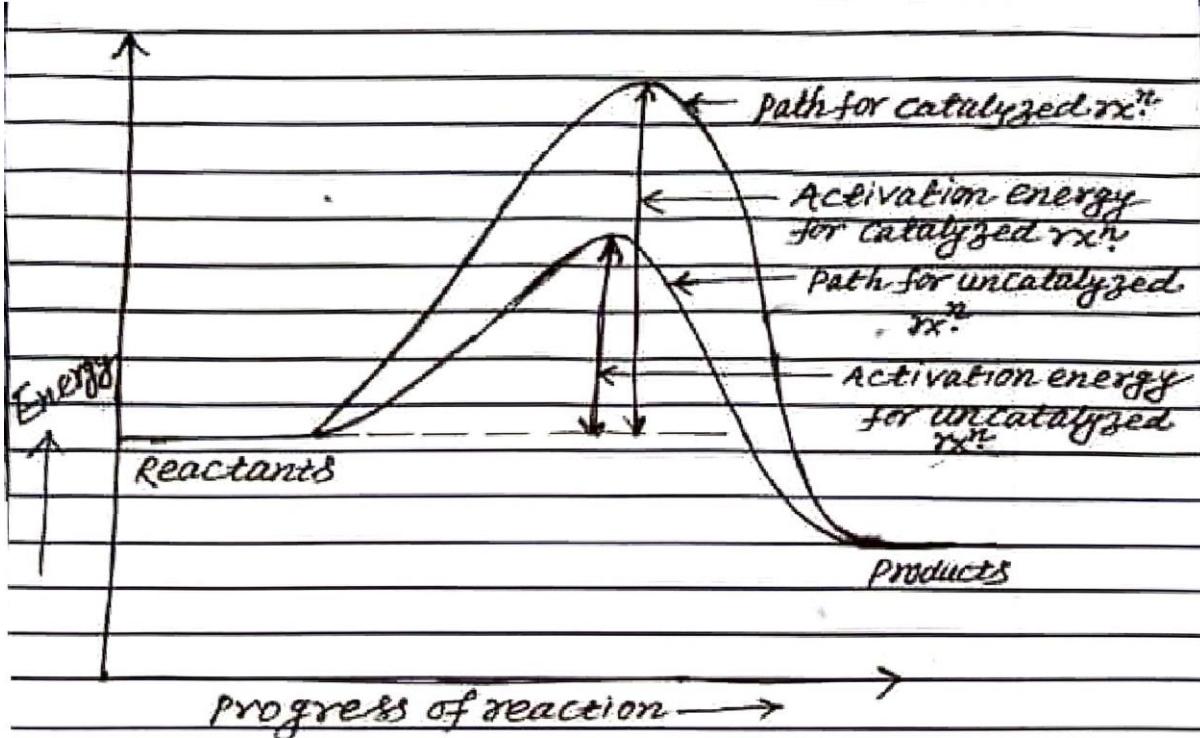


Fig: Energy profile diagram for catalyzed and uncatalyzed rxn [use -ve Catalyst]

4) Light *:

Light has effect only in photochemical reaction. The reaction that proceeds due to the absorption of photon from light is called photochemical reaction. The rate of photochemical reaction is directly proportional to the intensity of light. Light contains packet of energy called as photon which is absorbed by reactant . Thus K.E. of reactant increases and hence the rate of vibration increases . This causes the increase in collision frequency and hence the rate of effective collision increases . So the rate of formation of product increases.

5) Surface area of reactant* :

The rate of reaction is fastest in gaseous state and the slowest in solid state . In case of solid , if the reactant is taken in lump , surface area of exposed reactant molecules decreases which causes the decrease in collision frequency . So the rate of effective collision decreases and hence the rate of formation of product decreases.

On the other hand , if reactant is taken in powdered form then surface area of exposed reactant molecules increases which causes the increase in collision frequency. So, the rate of effective collision increases and hence the rate of formation of product increases .

6) Nature of reactant* : Rate of reactions are certainly governed by the nature of reactants. Different reactants have different composition , different shape and size , different chemical bond etc. these properties of substance greatly influence the rate of chemical reaction .

For example: $\text{Na} + \text{HCl} \longrightarrow$ reacts violently and $\text{Sn} + \text{HCl} \longrightarrow$ reacts very slowly. The change in rate of reaction of above chemical reaction is due to different nature of Na and Sn .

Order of reaction :

It is defined as the sum of powers raised in each concentration term of the reactant in rate law or rate law expression which can express the observed rate of reaction. It can be determined experimentally and it may be small integers like 1, 2, 3 etc. or zero or fractional number both positive and negative.

Let us consider the following reaction , i.e. $a A + b B \longrightarrow \text{Product}$

From the law of mass action, the theoretical rate express can be written as follows :

i.e. $\text{Rate} \propto [A]^a * [B]^b$

The experimental rate expression for this reaction becomes ,

$$\text{Rate} \propto [A]^m * [B]^n$$

Where 'm' is the order of reaction with respect to reactant 'A' and 'n' is the order of reaction with respect to reactant 'B'. The over all order is (m + n).

For elementary reaction (reaction that occurs in single step), the order of reaction will be stoichiometric coefficient of the reactant present in balanced chemical equation. In this case $a = m$ and $b = n$. For complex reaction (reaction that occurs several steps), the order of reac-

tion will not be stoichiometric coefficient of the reactant present in balanced chemical equation. It must be determined experimentally. In this case, $m \neq a$ and $n \neq b$.

Reactions of higher order are very rare because collision between greater number of reactant molecules at a time is very less possible .

Molecularity of reaction :

It is defined as the total number of reactant molecules which can collide simultaneously to form the product in rate determining step. It is determined theoretically and it is only small integers like 1, 2, 3 etc. but never zero or fractional number.

For elementary reaction (reaction that occurs in single step), the order and molecularity will be same. In this case, molecularity of reaction will be the stoichiometric coefficient of the reactant present in balanced chemical equation. For complex reaction (reaction that occurs several steps), the order and molecularity of reaction will be different. In this case, the step that proceeds with slowest rate is determined which is called rate determining step [RDS]. The molecularity for this complex reaction will be the no. of reactant molecules involved in rate determining step . Over all molecularity for complex reaction has no significance.

Reactions of higher molecularity are very rare because collision between greater number of reactant molecules at a time is very less possible .

[Alternative] Molecularity of reaction :

To define molecularity of reaction , chemical reaction can be classified into two types ,

i.e. 1) Simple or Elementary reaction

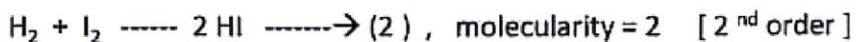
2) Complex reaction

1) Simple or Elementary reaction : The reaction that completes in a single step is known as simple or elementary reaction . Molecularity of simple reaction may be defined as the no. of reacting species that undergo simultaneous collision to give product.

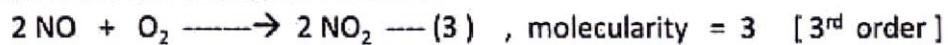
If there are only one reacting species , then the molecularity of reaction is one (1) and the reaction is known as Uni – molecular reaction .



If there are two reacting species , molecularity of reaction is two and reaction is known as Bi – molecular reaction .



If there are three reacting species , molecularity of reaction is 3, and the reaction is known as ter molecular reaction .



Note: Molecularity and order of elementary reaction are equal. Therefore, the 1st reaction is 1st order , 2nd reaction is 2nd order and 3rd reaction is 3rd order.

The rate of reaction of (1) is ## The rate of reaction of (2) is

$$\text{rate} = K [\text{SO}_2\text{Cl}_2]$$

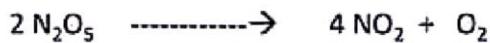
$$\text{rate} = K [\text{H}_2] [\text{I}_2]$$

The rate of reaction (3) is , $\text{rate} = K [\text{NO}]^2 [\text{O}_2]$

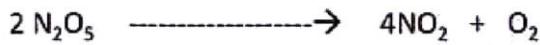
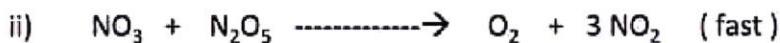
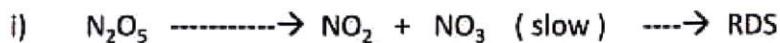
Note : Molecularity more than 3 is rare because as the no. of reacting species increases, the probability of simultaneous collision with proper orientation decreases significantly making molecularity more than 3 rare .

2) Complex reaction : The reaction that completes in more than one step (multiple steps) is known as complex reaction . A complex reaction can be written in more than one steps. These steps are elementary steps and each step has its own speed . The step having the least speed is known as rate determining step (RDS) because it determines the overall speed of the reaction. Then , molecularity of complex reaction is defined as the no. of reacting species involved in rate determining step.

For example ,



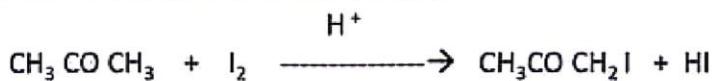
The molecularity of above reaction is supposed to be 2 if the reaction is elementary Or simple but the reaction is complex and it involves following steps which are also Known as mechanism of reaction .



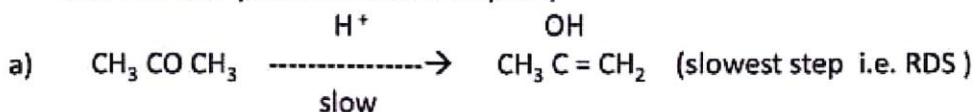
Here , 1st step is slow , there fore it is rate determining step and its rate is given by
rate = $K[\text{N}_2\text{O}_5]$

The RDS involves only 1 reacting species , so its molecularity is 1.

Iodination of acetone in acidic medium :



This reaction proceeds in two steps as,



Molecularity of this reaction seems to be 2 but it's molecularity is 1 as seen from the RDS .

Why high molecularity of reaction are rare ?

Since, the chances of simultaneous collision of reacting molecules goes on decreasing with increasing no. of molecules. So, high molecularity of reactions are rare.

V. Imp. Difference between order and molecularity

Order of reaction

- 1) It is the sum of powers raised in each concentration term of the reactant in rate law expression which can express the observed rate of reaction.
- 2) It is determined experimentally.
- 3) It may be small integers like 1,2,3 etc. or zero or fractional number both positive and negative.
- 4) Over all order for complex reaction has significance.
- 5) It is dependent upon the parameters like temperature, pressure, concentration etc.

Molecularity of reaction

- 1) It is the total no. of reactant molecules that can collide simultaneously to form the product in rate determining step.
- 2) It is determined theoretically.
- 3) It is only small integers like 1,2,3 etc. but never zero or fractional number.
- 4) Over all molecularity for complex reaction has no significance.
- 5) It is independent upon the parameters like temperature, pressure, concentration etc.

Rate law or rate law expression or rate law equation :

It is the mathematical relationship that relates rate of reaction with concentration of reactant . let us consider the following reaction, i.e. $a A + b B \rightarrow \text{product}$

The experimental rate expression for this reaction can be written as follows :

$$\text{i.e.} \quad \text{Rate} \propto [A]^m [B]^n$$

$$\text{Or} \quad \text{Rate} = k [A]^m [B]^n \quad \dots \quad (\text{i})$$

This eqⁿ. (i) is the rate law or rate law expression . Where 'm' and 'n' are the order of reaction with respect to reactant 'A' and 'B' respectively . K is the proportionality constant called as rate constant or specific rate constant or velocity constant .

Rate constant (K) :

let us consider the following reaction, i.e. $a A + b B \rightarrow \text{product}$

The rate law of the reaction is ,

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

Where K is rate constant of reaction .

If the concentration of both A and B is 1 mol L^{-1} . Then , the rate law becomes ,

$$\begin{aligned}\text{Rate} &= K * 1 * 1 \\ \text{Or} \quad \text{Rate} &= K\end{aligned}$$

Thus , Rate constant of a reaction may be defined as the rate of reaction when the concentration of all reactants is one mole per liter.

Unit for rate constant :

let us consider the following reaction, i.e. $A \xrightarrow{\hspace{1cm}}$ product

The rate law for this reaction becomes ,

$$\begin{aligned}\text{Rate} &= K [A]^m \\ \text{where } m &= \text{Order of reaction with respect to reactant A .}\end{aligned}$$

a) For zero order reaction ,

$$\begin{aligned}\text{i.e. } m &= 0, \quad \text{So , above rate law becomes , } \text{Rate} = K_0 [A]^0 \\ &\quad \text{Or } K_0 = \text{Rate}\end{aligned}$$

$$= \text{mol L}^{-1} \text{S}^{-1} \text{ or } \text{mol L}^{-1} \text{min}^{-1} \text{ or } \text{mol L}^{-1} \text{hr}^{-1}$$

[Note: Unit of rate of reaction is $\text{mol L}^{-1} \text{S}^{-1}$ and unit of concentration is mol L^{-1}]

b) For first order reaction ,

$$m = 1, \text{ So, above rate law becomes , } \text{Rate} = K_1 [A]^1$$

$$\begin{aligned}\text{Or } K_1 &= \text{Rate} / [A]^1 \\ &= \text{mol L}^{-1} \text{S}^{-1} / \text{mol L}^{-1} = \text{S}^{-1} \text{ or } \text{min}^{-1} \text{ or } \text{hr}^{-1}\end{aligned}$$

c) For second order reaction ,

$$m = 2, \text{ So, above rate law becomes , } \text{Rate} = K_2 [A]^2 \quad \text{Or } K_2 = \text{rate} / [A]^2$$

$$= \text{mol L}^{-1} \text{S}^{-1} / (\text{mol L}^{-1})^2 = \text{L mol}^{-1} \text{S}^{-1} \text{ or } \text{L mol}^{-1} \text{min}^{-1} \text{ or } \text{L mol}^{-1} \text{hr}^{-1}$$

d) For third order reaction ,

$$m = 3, \text{ So , above rate law becomes , } \text{Rate} = K_3 [A]^3$$

$$\text{Or } K_3 = \text{Rate} / [A]^3$$

$$= \text{mol L}^{-1} \text{S}^{-1} / (\text{mol L}^{-1})^3 = \text{L}^2 \text{ mol}^{-2} \text{S}^{-1} \text{ or } \text{L}^2 \text{ mol}^{-2} \text{min}^{-1} \text{ or } \text{L}^2 \text{ mol}^{-2} \text{hr}^{-1}$$

e) For m order reaction ,

So above reaction becomes , Rate = $K_m [A]^m$, Or $K_m = \text{Rate} / [A]^m$

$$\text{Or } K_m = (\text{Concentration} / \text{time}) / (\text{Concentration})^m$$

$$= (\text{Conc}^n.)^{1-m} * \text{time}^{-1}$$

$$= (\text{mol L}^{-1})^{1-m} * \text{time}^{-1}$$

Which is general form of unit of rate constant of reaction .

Note : For 2.5 order reaction ,

$$\text{We have, } K_{2.5} = (\text{mol L}^{-1})^{1-2.5} * \text{time}^{-1}$$

$$= \text{L}^{1.5} \text{ mol}^{-1.5} \text{ time}^{-1}$$

Characteristics / Features / Significance of rate constant

1) The value of rate constant identifies the rate of reaction . Greater the value of rate constant , faster will be the reaction rate and vice- versa.

2) The value of rate constant is always constant for a given reaction at the given set of conditions. So, it is also called as specific rate constant.

3) Different reactions have different values of rate constant .

4) The value of rate constant is dependent upon the temperature at which reaction is carried out.

5) Different order reactions have different unit for rate constant .

6) Rate constant is independent to the initial concentration of reactant for most of the reactions.

Imp. If the reaction is , $a A + b B \rightarrow$ products is second order write the possible rate expression. (differential rate expression).

i.e. $\frac{dx}{dt} = K_2 [A]^2 [B]^0$ [The rxⁿ. is 2nd order with respect to 'A' and zero order w.r.t.'B'].

$\frac{dx}{dt} = K_2 [A]^0 [B]^2$ [The rxⁿ. is 2nd order with respect to 'B' and zero order w.r.t.'A'].

$\frac{dx}{dt} = K_2 [A]^1 [B]^1$ [The rxⁿ. is 1st order with respect to both 'A' and 'B'].

If the reaction is , $a A + b B \rightarrow$ products is third order write the possible rate expression .

i.e. $\frac{dx}{dt} = K_3 [A]^3 [B]^0$

$$\frac{dx}{dt} = K_3 [A]^0 [B]^3$$

$$\frac{dx}{dt} = K_3 [A]^2 [B]^1$$

$$\frac{dx}{dt} = K_3 [A]^1 [B]^2$$

Imp. If the reaction $a A + b B + c C \rightarrow$ products is third order write the possible rate expression .

The possible rate expression are ,

1) $\frac{dx}{dt} = K_3 [A]^3 [B]^0 [C]^0$ 2) $\frac{dx}{dt} = K_3 [A]^0 [B]^3 [C]^0$

3) $\frac{dx}{dt} = K_3 [A]^0 [B]^0 [C]^3$ 4) $\frac{dx}{dt} = K_3 [A]^1 [B]^1 [C]^1$

5) $\frac{dx}{dt} = K_3 [A]^1 [B]^2 [C]^0$ 6) $\frac{dx}{dt} = K_3 [A]^1 [B]^0 [C]^2$

7) $\frac{dx}{dt} = K_3 [A]^0 [B]^1 [C]^2$ 8) $\frac{dx}{dt} = K_3 [A]^2 [B]^1 [C]^0$

9) $\frac{dx}{dt} = K_3 [A]^2 [B]^0 [C]^1$ 10) $\frac{dx}{dt} = K_3 [A]^0 [B]^2 [C]^1$

Identify the order of reactions whose rate constant are ,

$$1) \ 7.6 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1} \quad 2) \ \text{mol}^{-2} \text{ L}^2 \text{ s}^{-1}$$

Solution :

The general unit of rate constant is ,

$$(\text{mol L}^{-1})^{1-m} \text{ time}^{-1}$$

If time is in second ,

$$= (\text{mol L}^{-1})^{1-m} \text{ s}^{-1}$$

Comparing the unit with given unit ,

$$1) (\text{mol L}^{-1})^{1-m} \text{ s}^{-1} = \text{mol L}^{-1} \text{ s}^{-1} = (\text{mol L}^{-1})^1 \text{ s}^{-1}$$

Comparing ,

$$1-m = 1 \quad \text{Or} \quad m = 0$$

Hence, order of reaction is zero .

$$2) (\text{mol L}^{-1})^{1-m} \text{ s}^{-1} = \text{mol}^{-2} \text{ L}^2 \text{ s}^{-1} = (\text{mol L}^{-1})^{-2} \text{ s}^{-1}$$

Comparing ,

$$1-m = -2 \quad \text{Or} \quad m = 3$$

Hence, order of reaction is third .

For a reaction $P + Q \longrightarrow$ product is third order. Write the possible rate law expressions for the above reaction .

Solⁿ. The reaction is third order and P and Q are the reactants. So , the possible rate law expressions are:

$$\text{Rate} = K [P]^3 [Q]^0$$

$$\text{Rate} = K [P]^0 [Q]^3$$

$$\text{Rate} = K [P]^1 [Q]^2$$

$$\text{Rate} = K [P]^2 [Q]^1$$

The reaction $X + Y \rightarrow$ product is a second order reaction . Write three different rate law expressions which may be true to the above reaction .

Solⁿ . The reaction is second order and X and Y are the reactants. So , the possible rate law expressions are :

$$\text{Rate} = K[X]^2[Y]^0$$

$$\text{Rate} = K[X]^0[Y]^2$$

$$\text{Rate} = K[X]^1[Y]^1$$

For a given reaction, $A + B \rightarrow$ product , the rate law is found to be ,
rate = $K[A]^1[B]^2$. What happens to the rate of reaction when :

- a) Concentration of both A and B are doubled.
- b) Concentration of A is doubled and that of B remains constant.
- c) Concentration of A is halved and that of B is doubled.

Solⁿ . The rate law is, Rate = $K[A]^1[B]^2$

a) When concentration of both A and B are doubled , the rate law becomes,

$$\text{Rate}' = K[2A]^1[2B]^2$$

$$\text{Rate}' = 8K[A]^1[B]^2$$

i.e. , The reaction rate increases by 8 times .

b) When concentration of A is doubled and that of B remains constant, the rate law becomes,

$$\text{Rate}'' = K[2A]^1[B]^2$$

$$\text{Rate}'' = 2K[A]^1[B]^2$$

i.e. , The reaction rate increases by 2 times .

c) When concⁿ. of A is halved and that of B is doubled , the rate law becomes,

$$\text{Rate}''' = K[A/2]^1[2B]^2$$

$$\text{Rate}''' = 2K[A]^1[B]^2$$

i.e. , The reaction rate becomes double .

Identify reaction orders if the units of rate constant are :

- a) min^{-1}
- b) $\text{mol L}^{-1} \text{min}^{-1}$

Solution :

The general unit of rate constant is ,

$$(\text{mol L}^{-1})^{1-m} \text{ time}^{-1}$$

If time is in min ,

$$= (\text{mol L}^{-1})^{1-m} \text{ min}^{-1}$$

Comparing the unit with given unit,

a) $(\text{mol L}^{-1})^{1-m} \text{ min}^{-1} = \text{min}^{-1} = (\text{mol L}^{-1})^0 \text{ min}^{-1}$

Comparing ,

$$1-m = 0 \quad \text{Or} \quad m=1$$

Hence, order of reaction is 1st.

b) $(\text{mol L}^{-1})^{1-m} \text{ min}^{-1} = \text{mol L}^{-1} \text{ min}^{-1} = (\text{mol L}^{-1})^1 \text{ min}^{-1}$

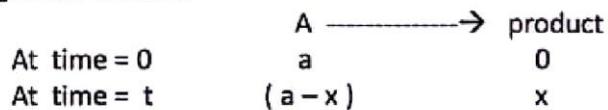
Comparing ,

$$1-m = 1 \quad \text{Or} \quad m=0$$

Hence, order of reaction is zero .

Differential and Integrated rate equation of zero order reaction :

Let us consider a general reaction



Let , the initial concentration of reactant be 'a' mol L⁻¹ and 'x' mol L⁻¹ be the concentration of product after time 't' , then remaining concentration of reactant after time 't' becomes (a - x) mole L⁻¹ .

The rate law equation of reaction is ,

$$\text{rate} = K [A]^m$$

For zero order reaction ,

$m = 0$, So above rate law becomes

$$\frac{dx}{dt} = K_0 [A]^0, \quad \text{Or} \quad \frac{dx}{dt} = K_0$$

$$= K_0 (a - x)^0 \quad \text{at time } t$$

$$\frac{dx}{dt} = K_0$$

This is the differential rate equation for zero order reaction .

To get integrated rate equation , the differential rate equation must be integrated.

Now, $\frac{dx}{dt} = K_0$

Or, $dx = K_0 dt$

Integrating both sides,

$$\int dx = \int K_0 dt$$

$$\text{Or, } x = K_0 t + c \quad \text{--- (i)}$$

Where c is the integration constant .

Now , to get the value of c ,

Applying the initial condition i.e. $t = 0$ and $x = 0$, then eqⁿ. (i) becomes,

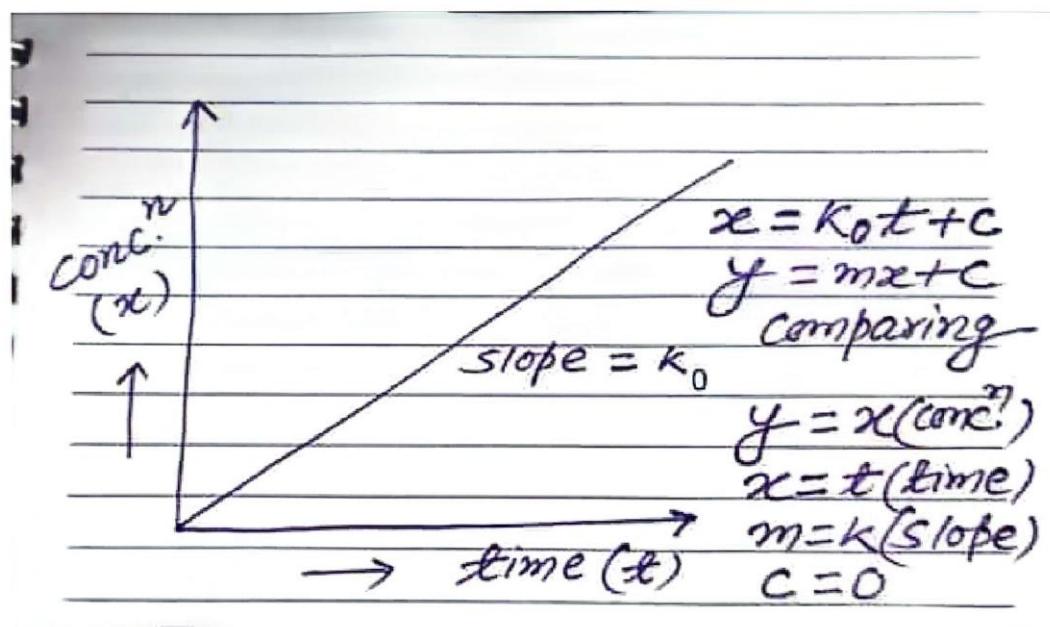
$$0 = K_0 * 0 + c, \quad \text{Or} \quad c = 0$$

Substituting the value of c in equation (i) . We get,

$$x = K_0 t \quad \text{Or, } K_0 = x/t$$

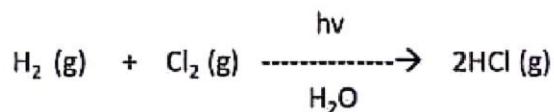
This is the integrated rate equation for zero order reaction.

Graphical representation :



Imp. Zero order reaction : The reaction in which reaction rate is independent to the concentration of reactant is called zero order reaction .

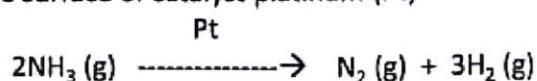
For e.g., 1) Formation of HCl gas from H₂ and Cl₂ gas in presence of sunlight on the surface of water.



2) Iodination of acetone :



3) Catalytic decomposition of ammonia gas : When decomposition of ammonia is carried out on the surface of catalyst platinum (Pt)

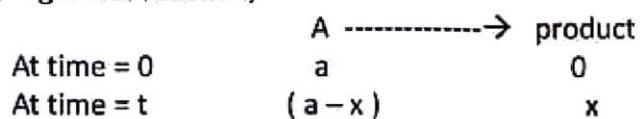


Characteristics of zero order reaction :

- i) Rate of the reaction is always constant i.e. it is independent of concentration of reactant.
- ii) Unit of rate constant is mol L⁻¹ time⁻¹.
- iii) Half life of zero order reaction is dependent on initial concentration of reactant.

Differential and Integrated rate equation of first order reaction :

Let us consider a general reaction,



Let , the initial concentration of reactant be 'a' mol L⁻¹ and 'x' mol L⁻¹ be the concentration of product after time 't' , then remaining concentration of reactant after time 't' becomes (a - x) mole L⁻¹ .

The rate law equation of reaction is ,

$$\text{rate} = K [A]^m$$

For first order reaction ,

$m = 1$, So above rate law becomes

$$\begin{aligned} \frac{dx}{dt} &= K_1 [A]^1, \quad \text{Or} \quad \frac{dx}{dt} = K_1 [A] \\ &= K_1 (a-x)^1 = K_1 (a-x) \quad \text{at time } t \end{aligned}$$

This is the differential rate equation for 1st order reaction .

To get integrated rate equation , the differential rate equation must be integrated.

Now, $\frac{dx}{dt} = K_1 (a-x)$

Or, $\frac{dx}{(a-x)} = K_1 dt$

Integrating both sides,

$$\int dx/(a-x) = K_1 \int dt$$

Or, $-\ln(a-x) = K_1 t + c \quad \text{----- (i)}$

Where c is the integration constant .

Now , to get the value of c , applying the initial condition i.e. $t = 0$ and $x = 0$, then eqⁿ . (i) becomes,

Or, $-\ln(a-0) = K_1 * 0 + c, \quad \text{Or}, \quad c = -\ln a$

Substituting the value of c in eqⁿ .(i)

Or, $-\ln(a-x) = K_1 * t - \ln a$

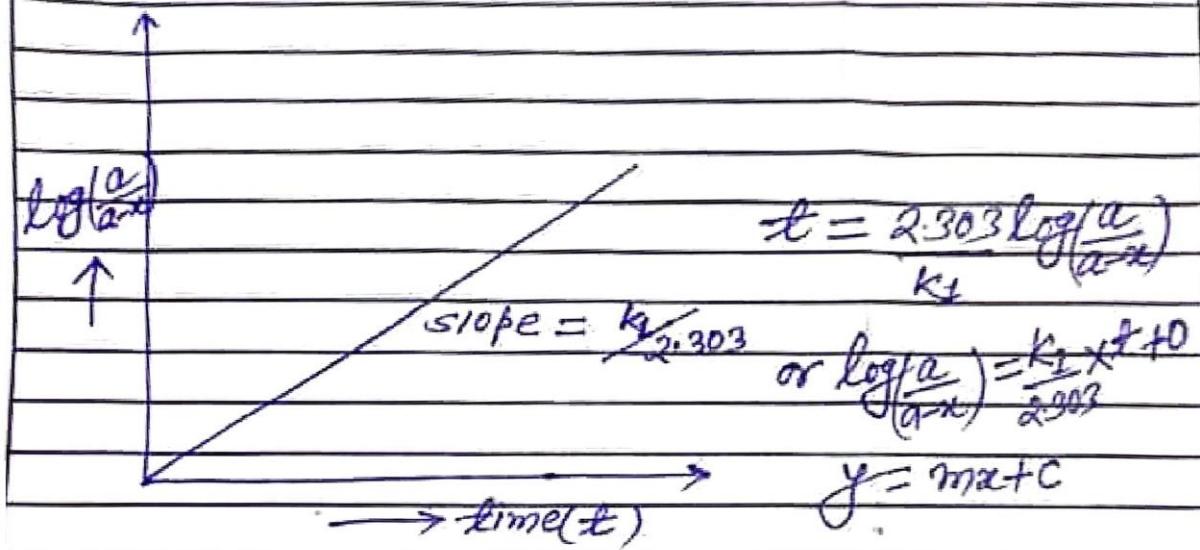
Or, $K_1 * t = \ln a - \ln(a-x)$

Or, $K_1 * t = \ln \left(\frac{a}{a-x} \right)$

Or, $K_1 * t = 2.303 \log(a/a-x) \quad [\ln = 2.303 \log]$

Or, $t = \frac{2.303}{K_1} \log(a/a-x) \quad \text{-----(ii)}$

This eqⁿ . is known as integrated rate equation for 1st order reaction .



Experimental determination of K_1 [Important only for Entrance] *:

The differential rate equation for first order reaction can be written as ,

$$dx/dt = K_1 [A]$$

$$\text{Or, } dx/dt = K_1 (a - x)$$

$$\text{Or, } dx/(a - x) = K_1 dt$$

Integrating both sides,

$$\int dx/(a - x) = K_1 \int dt$$

$$\text{Or, } -\ln(a - x) = K_1 t + c \quad \dots \quad (i)$$

Where c is the integration

Now , to get the value of c , applying the initial condition i.e. $t = 0$ and $x = 0$, then eqⁿ. (i) becomes ,

$$\text{Or, } -\ln(a - 0) = K_1 * 0 + c \quad \text{Or, } c = -\ln a$$

Substituting the value of c in eqⁿ.(i)

$$\text{Or, } -\ln(a - x) = K_1 * t - \ln a$$

$$\text{Or, } -2.303 \log(a-x) = K_1 * t - 2.303 \log a \quad [\ln = 2.303 \log]$$

Dividing this equation by - 2.303 on both sides, we get,

$$\text{Or, } \log(a-x) = K_1 * t / -2.303 + \log a$$

$$\text{Or, } \log(a-x) = -K_1 / 2.303 * t + \log a$$

Comparing above equation with $y = mx + c$, we get,

$$y = \log(a-x), m = -K_1 / 2.303, x = t \text{ and } c = \log a$$

when a graph is plotted between $\log(a-x)$ vs time (t) then, a straight line with -ve slope is obtained which has y -intercept of $\log a$.

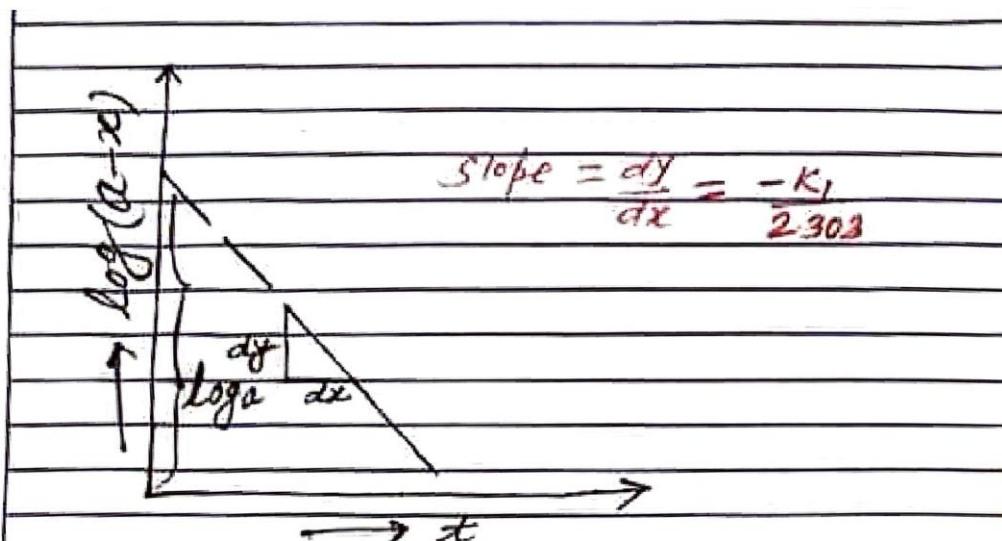


Fig: Plot of graph for $\log(a-x)$ vs t

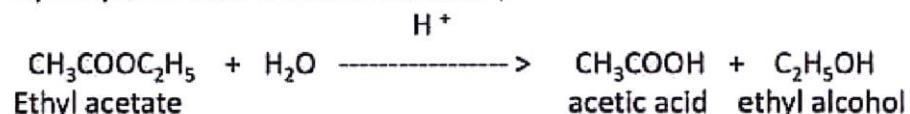
$$\text{Here, Slope} = -\frac{K_1}{2.303}$$

$$\therefore K_1 = -\text{Slope} \times 2.303$$

First order reaction : The reaction whose rate depends on 1st power of concentration of reactant is known as first order reaction .

For 1st order reaction , if the concentration of reactants is increased to double , rate of reaction also increases to double , and so on .

For e.g., 1) Hydrolysis of ester in acidic medium ,



2) Radioactive disintegration follows first order reaction.



3) Decomposition of N_2O_5



Characteristics of first order reaction :

- 1) It's rate depends on 1st power of concentration of reactants .
 - 2) Unit of rate constant is per time .
 - 3) Its half life period is independent of initial concentration of reactant.

Half life period ($t_{1/2}$) :

It is the time required to convert half of the reactant into product. It is denoted by ($t_{1/2}$) . During half period , $t = t_{1/2}$ and $x = a/2$

i) For zero order reaction : The integrated rate equation for zero order reaction is ,

$$x = K_0 * t \quad \text{Or, } K_0 = x/t \quad \rightarrow (i)$$

Using half life period, $t = t_{1/2}$, and $x = a/2$,

Putting the value of t and x in equation (i), we get,

$$K_0 = a/2 / t_{1/2}$$

$$\text{Or } t_{1/2} = a / 2 K_p$$

This is the expression of half life period for zero order reaction.

Now, $t_{1/2} \propto a$ [K_a is constant]

Thus, half life period of zero order reaction depends on initial concentration of reactant i.e. a .

Note: $K_s \propto a$ [$t_{1/2}$ is constant]

Thus, rate constant for zero order reaction is dependent to the initial concentration of reactant.

ii) For 1st order reaction : The integrated rate equation for 1st order reaction is ,

$$K_1 = 2.303 / t * \log (a / a - x) \longrightarrow (i)$$

During half life period , $t = 1/2$ and $x = a/2$

Putting the value of t and x in equation (i) , we get ,

$$K_1 = 2.303 / t_{1/2} * \log (a / a - a/2)$$

$$\text{Or, } K_1 = 2.303 / t_{1/2} * \log 2 a / a$$

$$\text{Or, } K_1 = 2.303 / t_{1/2} * \log 2$$

$$\text{Or, } K_1 = 0.693 / t_{1/2} \quad [2.303 * \log 2 = 0.693]$$

$$\text{Or, } t_{1/2} = 0.693 / K_1$$

This is the expression of half life period for 1st order reaction.

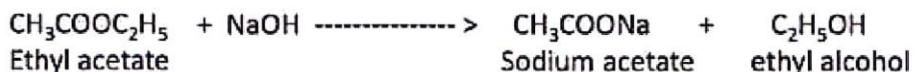
This equation indicates that half life period of 1st order reaction is independent of initial concentration of reactants.

Note : $K_1 \propto a^0$ [$t_{1/2}$ is constant]

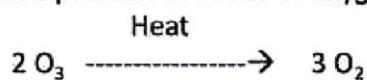
i.e. Rate constant for 1st order reaction is independent to the initial concentration of reactant.

iii) Second order reaction : The reaction in which reaction rate is dependent to the second power of concentration of reactant is called second order of reaction .

For e.g. , 1) Hydrolysis of ester in alkaline medium .



2) Thermal decomposition of ozone to oxygen .



The differential rate equation of 2nd order reaction is ,

$$\begin{aligned} \frac{dx}{dt} &= K_2 [A]^2 \\ \frac{dx}{dt} &= K_2 (a - x)^2 \quad \text{at time } t \end{aligned}$$

Integrated rate equation of 2nd order reaction is ,

$$K_2 = 1/t * x / a (a - x)$$

Half life period of 2nd order reaction is ,

$$t_{1/2} = 1 / K_2 a$$

iv) Third order reaction : The reaction in which reaction rate is dependent to the third power of concentration of reactant is called 3rd order of reaction .

For e.g. , 1) Combination of NO and Cl₂



2) Formation of NO₂ from NO and O₂



The differential rate equation of 3rd order reaction is ,

$$\frac{dx}{dt} = K_3 [A]^3$$

$$\frac{dx}{dt} = K_3 (a-x)^3 \text{ at time } t$$

Integrated rate equation of 3rd order reaction is ,

$$K_3 = 1/t * x(2a-x)/2a^2(a-x)^2$$

Half life period of 3rd order reaction is ,

$$t_{1/2} = 3/2K_3 a^2$$

Pseudo order of reaction : The chemical reaction which seems to be of higher order but actually follows the lower order kinetics is called pseudo order reaction. In pseudo order reaction , at least one of the reactants is present in large excess .

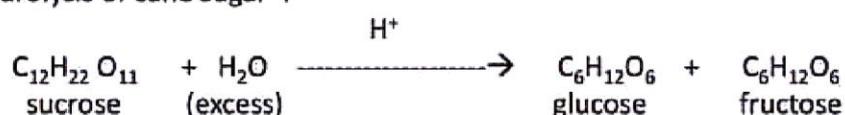
Note : If higher order reaction follows 1st order kinetics, then the reaction is known as pseudo 1st order reaction.

For e.g , 1) CH₃COOC₂H₅ + H₂O $\xrightarrow{\text{H}^+}$ CH₃COOH + C₂H₅OH
Ethyl acetate (excess) acetic acid ethyl alcohol

In this reaction , the concentration of water is very high as compared to the concentration of ester (ethyl acetate). So, the rate of reaction is independent of water used .

$$\text{Rate} = K [\text{CH}_3\text{COOC}_2\text{H}_5]^1 [\text{H}_2\text{O}]^0 = K [\text{CH}_3\text{COOC}_2\text{H}_5]$$

2) Hydrolysis of cane sugar :



Temperature Coefficient :

It is the ratio of two rate constant for a given reaction performed at two different temperatures usually separated by 10° C . It is denoted by μ . Its value ranges from 2 – 3 times and usually nearer to 2 .

Mathematically ,

$$\begin{aligned}\text{Temperature coefficient } (\mu) &= \frac{\text{Rate constant at higher temperature}}{\text{Rate constant at lower temperature}} \\ &= K(t+10^\circ\text{C}) / K(t^\circ\text{C}) \\ &= K_{35\text{ OC}} / K_{25\text{ OC}} \\ &= 2 - 3\end{aligned}$$

i.e. , Reaction rate increases or decreases by 2 to 3 times for every 10°C change in temperature .

Dependence of rate of reaction with temperature:

The dependence of rate of reaction with temperature can be studied under two headings :

1) Arrhenius Equation :

Rate constant is dependent upon temperature. This dependency of rate constant with temperature is shown by Arrhenius equation which is given below :

$$\text{i.e. } K = A * e^{-E_a/RT} \longrightarrow \text{(i)}$$

Where , K = Rate constant

E_a = Activation energy

R = Universal gas constant

T = Absolute temperature

A = Pre-exponential factor or frequency factor which depends upon the frequency and orientation of collision between reactant molecules.

$e^{-E_a/RT}$ = Arrhenius factor or Boltzmann's factors which shows the fraction of reactant molecules that can cross energy barrier.

In Arrhenius equation , rate constant is exponentially related with temperature. So, a small change in temperature causes large effect in the value of rate constant. It is estimated that the reaction rate increases by 2-3 times for every 10°C rise in temperature.

Taking In on both sides in eqⁿ. (i) , we get ,

$$\ln K = \ln (A * e^{-E_a/RT})$$

$$\text{Or } \ln K = \ln A + \ln e^{-E_a/RT}$$

$$\text{Or } \ln K = \ln A - E_a / RT \quad [\ln e = 1]$$

$$\text{Or } 2.303 \log K = 2.303 \log A - E_a / RT \quad [\ln = 2.303 \log]$$

Dividing this equation by 2.303 on both sides, we get,

$$\log K = \log A - E_a / 2.303 RT$$

$$\log K = -E_a / 2.303 R * 1/T + \log A \longrightarrow \text{(ii)}$$

Comparing eq^n. (ii) with $y = m x + c$, we get,

$$y = \log K, \quad m = -E_a / 2.303 R, \quad x = 1/T \quad \text{and} \quad c = \log A$$

When a graph is plotted between $\log K$ vs $1/T$, then, st. line with -ve slope is obtained which has y -intercept of $\log A$.

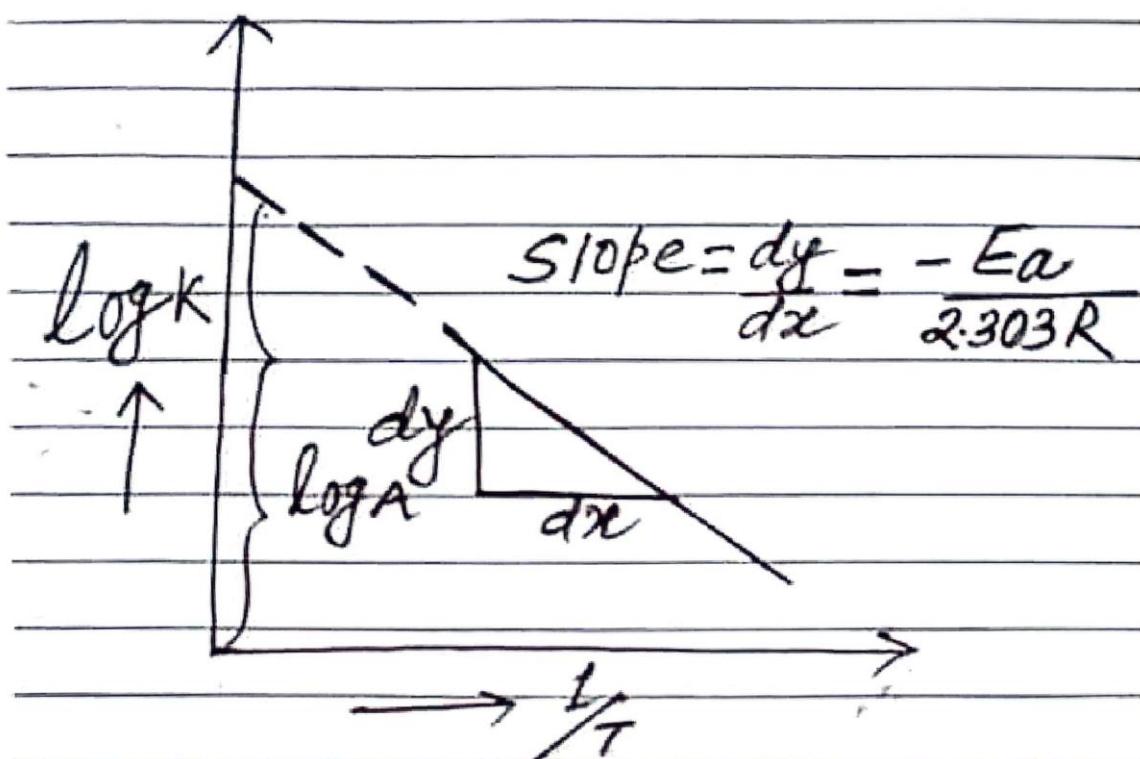


Fig: Plot of graph for
 $\log K$ vs $1/T$

Here, slope (m) = - $E_a / 2.303 R$

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

If the reaction is carried out at two different temperatures T_1 and T_2 then eqⁿ. (ii) can be written as ,

$$\log K_1 = - E_a / 2.303 R * 1/T_1 + \log A \longrightarrow \text{(iii)}$$

$$\log K_2 = - E_a / 2.303 R * 1/T_2 + \log A \longrightarrow \text{(iv)}$$

Subtracting (iii) from (iv) , we get,

$$\text{Or, } \log K_2 - \log K_1 = - E_a / 2.303 R * 1/T_2 + \log A - \{- E_a / 2.303 R * 1/T_1 + \log A\}$$

$$\text{Or, } \log K_2 / K_1 = - E_a / 2.303 R * 1/T_2 + \log A + E_a / 2.303 R * 1/T_1 - \log A$$

$$\text{Or } \log K_2 / K_1 = E_a / 2.303 R [1/T_1 - 1/T_2]$$

This is the final form of Arrhenius equation from which activation energy can be determined , if T_1 , T_2 , K_1 and k_2 are known.

V.Imp. 2) Collision Theory for Bimolecular Reaction :

This theory is applicable only for bimolecular reaction . According to this theory, the reaction is possible only when there is collision between reactant molecules. The no. of collisions occurring between the reactant molecules per time is called collision frequency (Z). This collision frequency is dependent upon temperature . At higher temperature , K.E. of reactant increases which causes the increase in collision frequency. The collision that can form the product is called effective collision (N) . For the collision to become effective , reactant molecules must collide with each other in proper orientation and with suitable energy. This factor is called Steric factor (P) .

$$\text{i.e. Rate of reaction } \propto N$$

$$\text{Also, rate of reaction } \propto P * Z$$

$$\text{So, } N \propto P * Z$$

To illustrate collision theory , let us consider a reaction between Hydrogen and Iodine molecules to from hydrogen iodide. It can be represented as ,



To form the product HI , there must be collision between H₂ and I₂ molecules. When H₂ and I₂ molecules collide with each other in proper orientation and with suitable energy , existing bond in reactant starts disappearing and new bond in product starts appearing . During this process , a highly unstable short lived molecular complex (H₂ I₂) is formed which is called activated complex or transition state complex . In activated complex , existing bonds in reactants are in the state of disappearance and new bonds in products are in the state of appearance. This activated complex finally decomposes to form the product. It can be represented diagrammatically as,

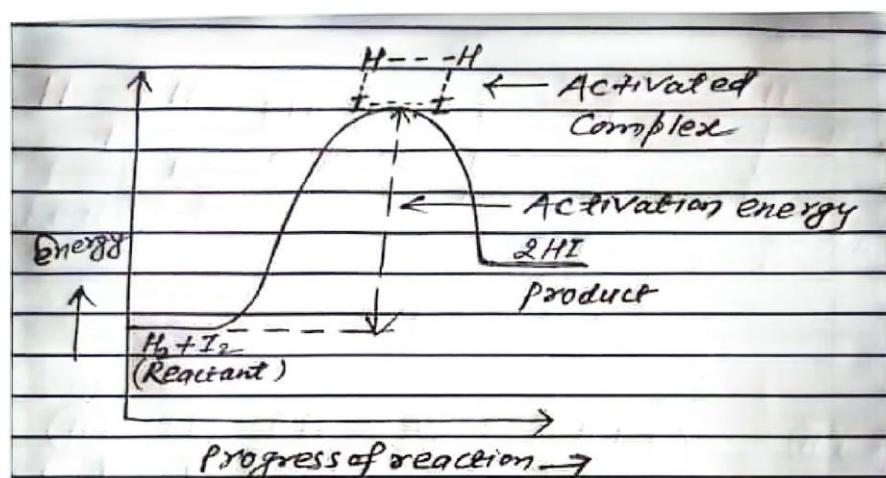
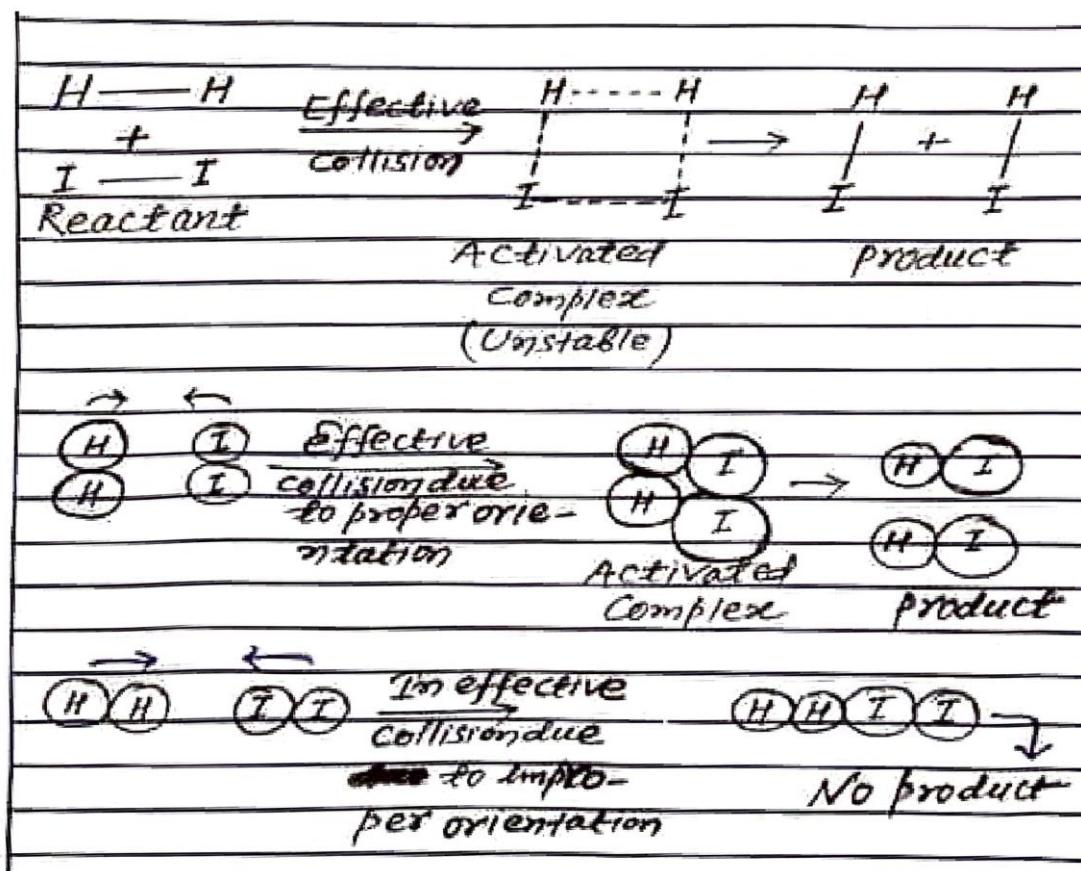


Fig: Energy profile diagram showing formation of H_2O

Catalysis and Types of Catalysis

Catalysis

*A **catalyst** is defined as a substance which alters the rate of a chemical reaction, itself remaining chemically unchanged at the end of the reaction.

*The action of catalyst to influence the rate of reaction is called **catalysis**.

*A catalyst may increase or decrease the rate of a reaction.

*A catalyst which enhances the rate of reaction is called a positive catalyst and the process is positive catalysis.

*A catalyst which retards the rate of a reaction is called a Negative catalyst and the process is Negative catalysis.

Characteristics of catalyst

1. Catalyst remain chemically unchanged at the end of the reaction but may change physically like change of colour, particle size etc.
2. Even a small quantity of catalyst is sufficient to cause unlimited reaction.
3. Efficiency of a catalyst is increased when used in finely divided form.
4. A catalyst is specific in its action. This means particular catalyst catalysis particular type of reaction.
5. A catalyst does not affect the final position of equilibrium, although it shortens the time required to establish the equilibrium.
6. Catalytic activity is found to be higher at particular point of catalyst called the active centre of catalyst. The active centre may be entire catalyst surface, corner of catalyst or cracks present on catalyst.
7. Greater the surface area of catalyst, greater will be the rate of reaction. For this reason, catalyst is used in finely divided form.

Types of Catalysis:

Catalysis is of two types;

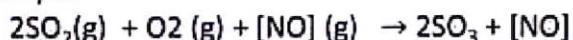
- a. Homogeneous catalysis
 - b. Heterogeneous catalysis

a. Homogeneous catalysis

In this type of catalysis, the reactants and the catalyst used are in same phase.

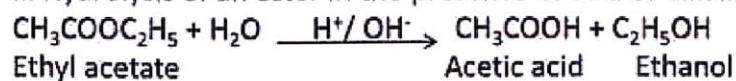
Examples:

- i. Oxidation of sulphur dioxide to sulphur trioxide with nitric oxide as catalyst.



This is homogeneous catalysis in gas phase

- #### ii. Hydrolysis of an ester in the presence of acid or alkali.



This is homogeneous catalysis in solution or liquid phase.

b . Heterogeneous catalysis

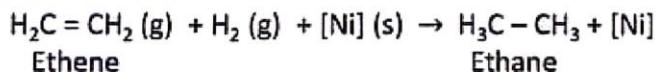
- ❖ The catalysis in which the reactants and the catalyst are in different phases is called Heterogeneous catalysis.
 - ❖ In most such cases, reactants are in gas phase while the catalyst in solid phase.
 - ❖ The process is also called contact or surface catalysis, since the reaction occurs by contact of reactants on the catalyst surface.
 - ❖ Usually the catalyst is a finely divided metal or a gauze.
 - ❖ This type of catalysis has great industrial importance.

Examples:

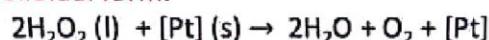
- i. Combination of sulphur dioxide and oxygen in the presence of finely divided Pt or V_2O_5



- ii. Hydrogenation reaction of unsaturated organic compounds catalyzed by finely divided nickel.



- ### iii. Decomposition of aqueous solution of H_2O_2 catalyzed by MnO_2 or Pt in colloidal form.

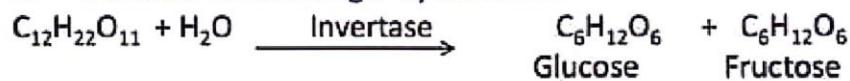


Enzyme Catalysis

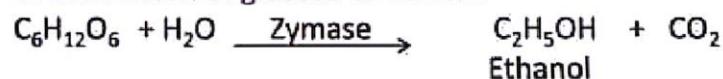
- ❖ The catalysis brought about by enzymes are called Enzyme catalysis.
 - ❖ Enzymes are complex protein molecules which act as catalyst to speed up organic reactions in living cells.
 - ❖ Numerous organic reactions are taking place in the body of plants and animals to maintain the life processes. These reactions are catalysed by enzymes.

Examples;

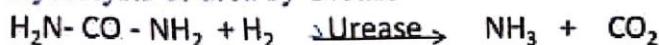
a. Inversion of cane sugar by Invertase



b. Conversion of glucose to ethanol



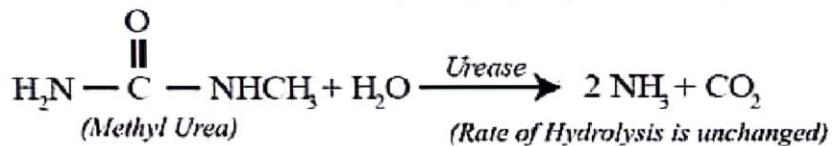
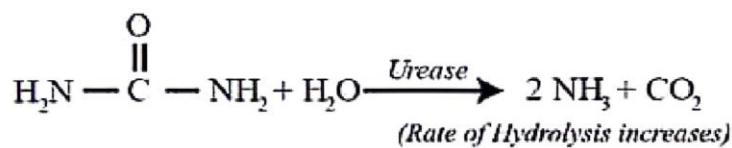
c. Hydrolysis of urea by *Urease*



General characteristics of enzyme catalysis:

- Enzyme catalysts are more specific in comparison to normal catalysts. This means, particular enzyme catalysis particular reaction only.

Example; Urease enzyme catalysis the hydrolysis of only urea but not substituted or methylated urea.



- The action (or efficiency) of an enzyme is found to be maximum at a particular temperature and pH called optimum temperature and optimum pH.
- ✓ The optimum temperature for enzymes in a human body is 37°C (98.6°F), which is the normal body temperature.
- ✓ Optimum pH for most enzymes in our body is 7.4, which is the pH of our blood and body fluids

