



Chapter # 07

Reaction Kinetics



Reaction Kinetics

Reaction kinetics is the study of the rates of chemical reactions.

It includes a variety of experimental methods for measuring reaction rates, orders and mechanisms of reactions.

Importance of Reaction Kinetics

- Reaction kinetics helps to determine the mechanism of reactions.
- Reaction control is important in industry. The factors affecting a reaction must be known in order to discover conditions for economic use of reactions.

Short Question

What is the importance of reaction kinetics?

Types of Reactions

The rates of different chemical reactions differ greatly.

Fast reaction

Some reactions take almost no time to complete. Such reactions are fast reactions.

e.g., The reaction of NaCl with AgNO_3

Slow reactions

Some reactions occur very slowly and may take even years to complete

e.g. Rusting of iron

Moderate reactions

Some reactions occur at moderate rates, neither too slow nor too fast.

e.g. The hydrolysis of ester.

Information

An explosion is a swift reaction that happens within a fraction of a second, the rusting of iron is a slow process that may take days or months. The rates of reactions occurring during the explosion are enormous.

COLLISION THEORY

Collision theory explains how reactions occur.

Postulates of Collision Theory

- For a chemical reaction to take place, the particles (atoms, ions or molecules) of reactants must form a homogeneous mixture and collide with one another.
- The collisions may be effective or ineffective depending upon the energy of the colliding particles.
- When the collisions are effective, they give the products. Otherwise the colliding particles just bounce back.
- For effective collisions, the colliding particles must possess a certain amount of energy. Also they approach each other with the proper orientation.

Short Question

What are the postulates of collision theory?

The minimum amount of energy, required for an effective collision between the reacting species, is called activation energy.

Most of the reactions are slow, showing that all the collisions are not equally effective.



Quick Check 7.1

a. The collision frequency and the orientation of molecules are necessary conditions a reaction to occur. Justify the statement.

The rate of a reaction is directly related to the collision frequency. Thus more collisions increases the chances of reaction. However, for reaction to occur, molecules must have activation energy and must collide with proper orientation. Bonds breaking and making only occurs at some proper orientation. If the molecules do not have proper orientation, then the collision may not be result in reaction.



b) What role does the activation energy play in chemical reactions?

The minimum amount of energy, required for an effective collision between the reacting species, is called activation energy. Only the colliding molecules with proper activation energy cross the energy barrier and give the products.

c) How does the activation energy affect the rate of reaction?

Activation energy is required to cross the energy barrier and give products. So, generally, higher the activation energy, lower will be the rate of reaction. It is because, number of high energy molecules decreases as the activation energy increases.

RATE OF REACTION

It is the change in concentration of reactants or products divided by the time taken for the change

$$\text{Rate} = \frac{\text{Change in conc. of the substance}}{\text{Time taken for change}} = \frac{\Delta x}{\Delta t}$$

Where Δx is a very small change in the concentration of a reactant or a product in a very small time interval Δt .

As the reaction proceeds, the concentration of reactant decreases with time and the concentration of product increases with time.

Short Question

Define rate of chemical reaction and give its units.

Units

- Usually, concentrations are expressed in mol dm^{-3} and time in sec. Therefore, units of rate are

$$\text{Rate} = \frac{\Delta x}{\Delta t} = \frac{\text{mol dm}^{-3}}{\text{sec}} = \text{mol dm}^{-3} \text{ sec}^{-1}$$

- However, for a slow reaction the units may be $\text{mol dm}^{-3} \text{ min}^{-1}$ or even $\text{mol dm}^{-3} \text{ h}^{-1}$.
- For a gas phase reaction, units of pressure are used in the place of molar concentrations.

Explanation

Consider a reactant A which is changing irreversibly to the product B.



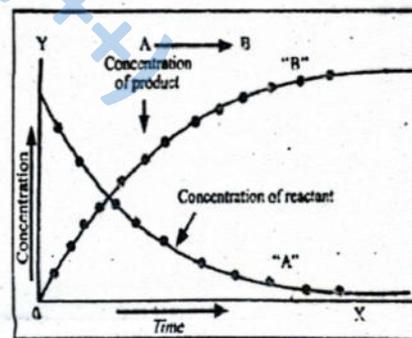
The rate of a general reaction, $A \rightarrow B$, can be expressed in terms of rate of disappearance of the reactant A or the rate of appearance of the product B.

$$\text{Mathematically, Rate of reaction} = -\frac{\Delta[A]}{\Delta t} = +\frac{\Delta[B]}{\Delta t}$$

where [A] and [B] are the concentrations of A and B, respectively.

In rate equation

- The negative sign indicates a decrease in the concentration of the reactant.
- The positive sign indicates an increase in the concentration of the product.



Average Rate

The rate of reaction between two specific time intervals or the rate over a time period.

Instantaneous Rate

The rate at any one instant during the interval is called the instantaneous rate.

- The average rate and instantaneous rate are equal for only one instant in any time interval.
- At the start of reaction, the instantaneous rate is higher than the average rate.
- At the end of the interval the instantaneous rate becomes lower than the average rate.

Differentiate between Average and Instantaneous rate of reaction

AVERAGE RATE OF REACTION		INSTANTANEOUS RATE OF REACTION	
1	The rate of a reaction between specific time intervals is called the average rate of reaction.	1	The rate at any one instant during the interval is called instantaneous rate.
2	It is given as $\text{Average rate} = \frac{c_2 - c_1}{t_2 - t_1} = \frac{\Delta c}{\Delta t}$	2	It is given as $\text{Instantaneous Rate} = \frac{dx}{dt}$
3	It remains constant during an interval	3	It may change during an interval.
4	It is a constant quantity for an interval	4	It is higher than average rate in the beginning of reaction and lower at the end of reaction.
4	It cannot give order of reaction.	5	It can be used to determine the order of reaction.

Sample Problem 7.1

The reaction for the formation of ammonia in Haber process is:



i. Calculate the instantaneous rate after 1.0 min

ii. What is the average rate of production of ammonia for the system, between 1.0 and 4.0 minutes?

Solution (i)

The instantaneous rate at 1.0 min can be calculated as

$$\text{Instantaneous Rate} = \frac{\Delta C}{\Delta t} = \frac{2.7 \text{ mol dm}^{-3}}{1 \text{ min}} = 2.7 \text{ mol dm}^{-3} \text{ min}^{-1}$$

If the concentration of ammonia is 3.5 mol dm^{-3} after 1.0 min and 6.2 mol dm^{-3} after 4.0 minutes?

Solution (ii)

$$\begin{aligned}\Delta C &= \Delta[\text{NH}_3] = (6.2 - 3.5) \text{ mol dm}^{-3} \\ &= 2.7 \text{ mol dm}^{-3}\end{aligned}$$

$$\begin{aligned}\Delta t &= (4.0 - 1.0) \\ &= 3.0 \text{ min}\end{aligned}$$

$$\text{Rate of formation of NH}_3 = \frac{\Delta[\text{NH}_3]}{\Delta t} = \frac{2.7 \text{ mol dm}^{-3}}{3 \text{ min}} = 0.90 \text{ mol dm}^{-3} \text{ min}^{-1}$$

The rate of production of NH_3 gas over the given time interval is $0.90 \text{ mol dm}^{-3} \text{ min}^{-1}$.

Quick Check 7.2

The reaction of hydrogen and iodine to make hydrogen iodide at a particular temperature, $\text{H}_{2(g)} + \text{I}_{2(g)} \rightarrow 2\text{HI}_{(g)}$ was studied at various times. At 100.0s after the start of the reaction, the iodine concentration had fallen from $0.010 \text{ mol dm}^{-3}$ to $0.0080 \text{ mol dm}^{-3}$. What is the average rate of reaction during this period?



Solution

$$\Delta C = \Delta[\text{I}_2] = (0.010 - 0.0080) \text{ mol dm}^{-3}$$

$$= 0.002 \text{ mol dm}^{-3}$$

$$\Delta t = 100 \text{ s}$$

$$\text{Rate of reaction} = \frac{\Delta[\text{I}_2]}{\Delta t} = \frac{0.002 \text{ mol dm}^{-3}}{100 \text{ s}} = 2.0 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$$

ADDITIONAL MCQs

(Answers on Page 229)

- The rate of reaction determined at any given time is called
(A) instantaneous rate (B) average rate (C) both (D) none
- The minimum amount of energy required for an effective collision is called
(A) Activation energy (B) Internal energy (C) Translational energy (D) None
- The energy of activated complex is:
(A) Greater than the reactants & products (B) Less than the reactants & products
(C) Equal to the products (D) Equal to the reactants

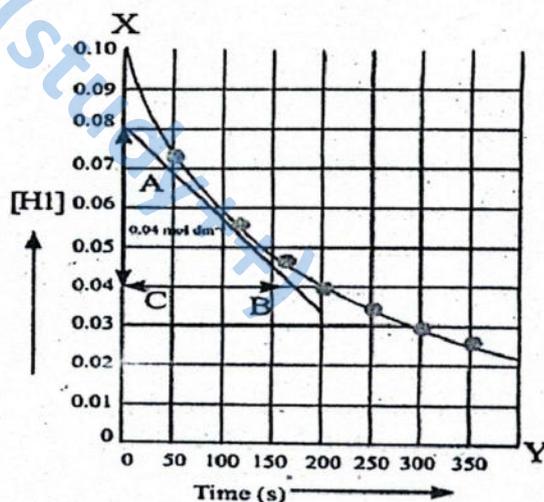
MEASURING THE RATE OF A CHEMICAL REACTION

- The measurement of rate of a chemical reaction involves the determination of the concentration of reactants or products at regular time intervals as the reaction progresses.
- To determine the rate of reaction for a given length of time, a graph is plotted between time on x-axis and concentration of a reactant on y-axis. So, a curve is obtained.

Example:

Consider the decomposition of HI to H_2 and I_2 at 500°C . The data is given in the table. A graph is plotted as shown between time on x-axis and concentration of HI in mol dm^{-3} on y-axis. Since HI is a reactant, so a falling curve is obtained.

Concentration of HI (mol dm^{-3})	Time (s)
0.100	0
0.0716	50
0.0558	100
0.0457	150
0.0387	200
0.0336	250
0.0296	300



- The steepness or slope of the concentration-time curve reflects the progress of reaction.
- Greater the slope of curve near the start of reaction, greater is the rate of reaction.
- To measure the rate of reaction, a tangent is drawn say, at 100 seconds.
- The slope of that tangent is measured by drawing a triangle.



- The slope of the tangent is the rate of reaction at that point.
- A right-angled triangle ABC is completed with a tangent as hypotenuse.
e.g. from graph, in 100 sec, the change in concentration is 0.04 mol dm^{-3} . The rate is then calculated by using the following expression.

$$\text{Rate of reaction} = \frac{\Delta C}{\Delta t}$$

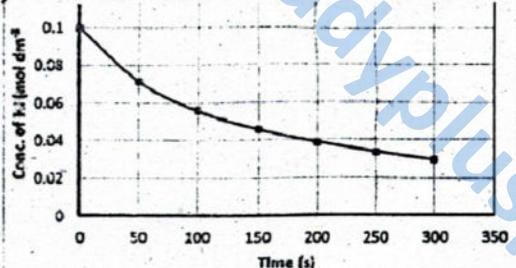
$$\text{Rate} = \frac{0.04 \text{ mol dm}^{-3}}{100 \text{ sec}}$$

$$= 4 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$$

- This shows that the concentration of HI is decreasing by $2.5 \text{ moles per dm}^{-3}$ per second during the given interval.
- If a graph is plotted between time on x-axis and concentration of any of the products i.e H_2 or I_2 , then a rising curve is obtained.
- The value of the tangent at 100 seconds will give the same value of rate of reaction as $4 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$.

Quick Check 7.3

a) Plot the data in Table 7.1 for HI.



b) Calculate the rate after 300 s (when the concentration is $0.0296 \text{ mol dm}^{-3}$) by drawing a tangent.

At 300 s, the rate is

$$\text{Rate of reaction} = \frac{\Delta C}{\Delta t} = \frac{0.007 \text{ mol dm}^{-3}}{140 \text{ sec}} = 5 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$$

c) Use the same method to calculate the rate of reaction at HI concentrations of 0.10 mol dm^{-3} , $0.050 \text{ mol dm}^{-3}$ and 0.02 mol dm^{-3} .

- At 0.1 mol dm^{-3} s, the rate is

$$\text{Rate of reaction} = \frac{\Delta C}{\Delta t} = \frac{0.03 \text{ mol dm}^{-3}}{42 \text{ sec}} = 7.14 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$$

- At 0.05 mol dm^{-3} s, the rate is

$$\text{Rate of reaction} = \frac{\Delta C}{\Delta t} = \frac{0.065 \text{ mol dm}^{-3}}{200 \text{ sec}} = 3.25 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$$

- The conc. 0.02 mol dm^{-3} is outside the given data range

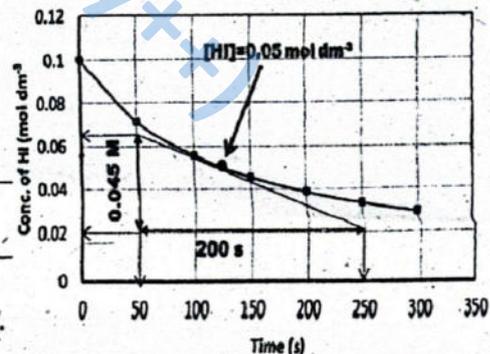
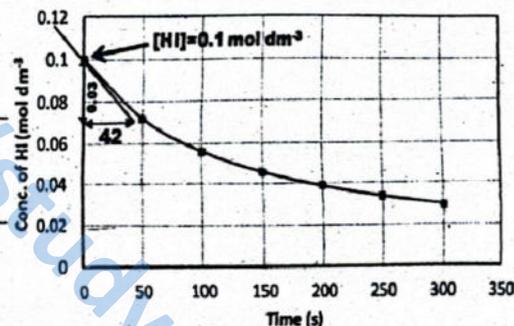
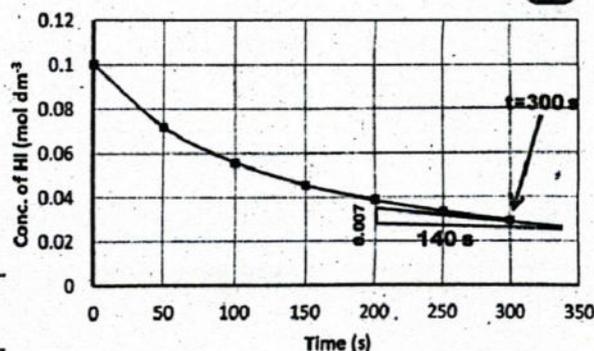
d) What do you deduce about the rate of the reaction with time from these calculations?

The rate of reaction decreases with time.

In the beginning, the rate is very fast. At the end the rate is very slow.

e) At which concentration, the rate is highest, and lowest?

The rate is highest in the beginning (at 0.1 mol dm^{-3}). It is because a large amount of reactant is present in the beginning. Hence, rate is very fast. At the end ($t=300 \text{ s}$), the rate is very slow, because amount of reactant is small.



MEASUREMENT OF CONCENTRATION

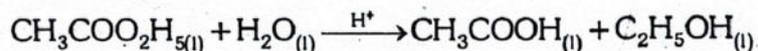
The change in concentrations of reactants or products can be determined by both physical and chemical methods. This depends upon the type of reactants or products involved.

a) Chemical Method

This is particularly suitable for reactions in solution. In this method, chemical analysis of a reactant or a product is done.

Example:

- Consider the acid hydrolysis of an ester (ethyl acetate) in the presence of a small amount of an acid.



- In hydrolysis of an ester, the solution of ester in water and the acid catalyst are allowed to react.
- After some time, a sample of reaction mixture is withdrawn by a pipette. It is poured into a flask along with about four times of ice-cold water. The dilution and chilling stop the reaction.
- The acid formed is titrated against a standard alkali, say NaOH, using phenolphthalein as an indicator.
- The analysis is repeated at various time intervals after the start of reaction.
- This gives data about the change in concentration of acetic acid formed during the reaction at different time.
- Finally, a graph is plotted between concentration of acetic acid and time to find rate of reaction.

b) Physical Methods

Some of the methods used for the measurement of concentration are as follows:

i) Spectrophotometry or colorimetry

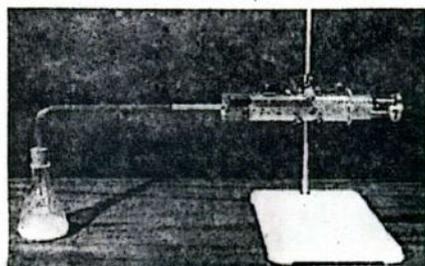
- This method is used if a reactant or a product absorbs ultraviolet, visible or infrared radiation.
- The rate of reaction can be measured by measuring the amount of radiations absorbed.
- e.g. The concentration can be measured using the colorimetry.

ii) Electrical conductivity method

- This method is used if ions are involved in a reaction.
- The conductivity of such a solution depends upon the rate of change of concentration of the reacting ions or the product ions.
- The conductivity will be proportional to the rate of change in the concentration of such ions.

iii) Volume change method

- This method is used when gases are involved and so changes in volumes occur.
- The volume change is directly proportional to the extent of reaction, and changes in concentration.



Interesting information

The rates of some very fast reactions can be monitored using stopped-flow spectrophotometry. In this technique, very small volumes of reactants are driven at high speed into a mixing chamber. From here they go to an observation cell, where the progress of the reaction is monitored (usually by measuring the transmission of ultraviolet radiation through the sample). A graph of rate of reaction against time can be generated automatically.

FACTORS AFFECTING RATE OF A CHEMICAL REACTION

- The rates at which reactants are consumed and products are formed during chemical reactions vary greatly.
- Even a chemical reaction involving the same reactants may have different rates under different conditions.
- The factors affecting the rates of reactions are
 1. Concentrations of the reactants
 2. Temperature of the system
 3. Surface area
 4. Catalyst

Did You Know?

In the case of reactions that involve gaseous reactants, an increase in pressure increases the concentration of the gases which leads to an increase in the rate of reaction. However, pressure change has no effect on the rate of reaction if the reactants are either solids or liquids.

1. Concentration

- According to the law of mass action, the greater the concentration of the reactants, the higher the rate of reaction.
- So, when the concentration of one or more reactants increases, rate of reaction increases.
- This is because increasing the concentration increases the collisions between the reacting particles. Thus, rate of reaction increases.

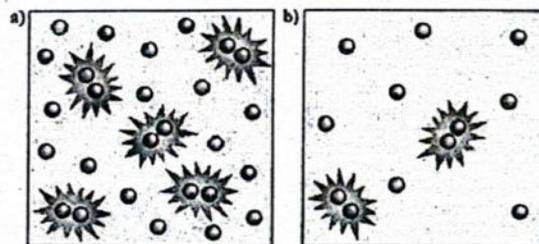


Fig. The reaction in box a) will occur faster than that in, b) due to the higher concentration.

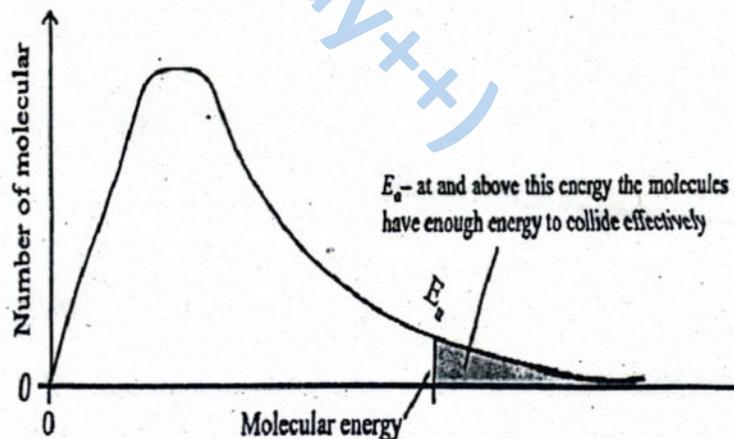
Exercise Q2 (e) Briefly summarize the effects of temperature and surface area on the rates of reactions

2. Temperature

- Increase in temperature increases the reaction rate.
- The rate either doubles or triples for every 10 °C rise in temperature.
- Molecules at higher temperatures have more thermal energy. So, they collide more frequently and with greater energy. So, the rate of reaction increases.

(Maxwell-Boltzmann distribution curve)

- The distribution of energies at a given temperature is shown in graph. This is called the Boltzmann distribution.
- This, graph shows that in a reaction,
 - ✓ a few particles have very low energy
 - ✓ a few particles have very high energy,
 - ✓ many particles have intermediate energy.
- The total area under the whole curve represents the total number of particles.
- The activation energy is the minimum energy required for colliding particles to convert into the product.



- The shaded area shows the number of particles with energy greater than the activation energy, E_a .
- When the temperature of a reaction is raised, the average kinetic energy of the particles increases.
- The reacting particles move more quickly at a higher temperature. So, more frequent collisions occur. Thus, the ratio of fruitful collisions also increases greatly.
- Hence, following changes occur in the Boltzmann distribution curve.
 - ✓ At the higher temperature, the curve becomes flatter and the peak shifts to the right.
 - ✓ For 10°C rise in temperature, the shaded area under the curve approximately doubles.
- Hence, increasing the temperature increases the rate of a reaction.

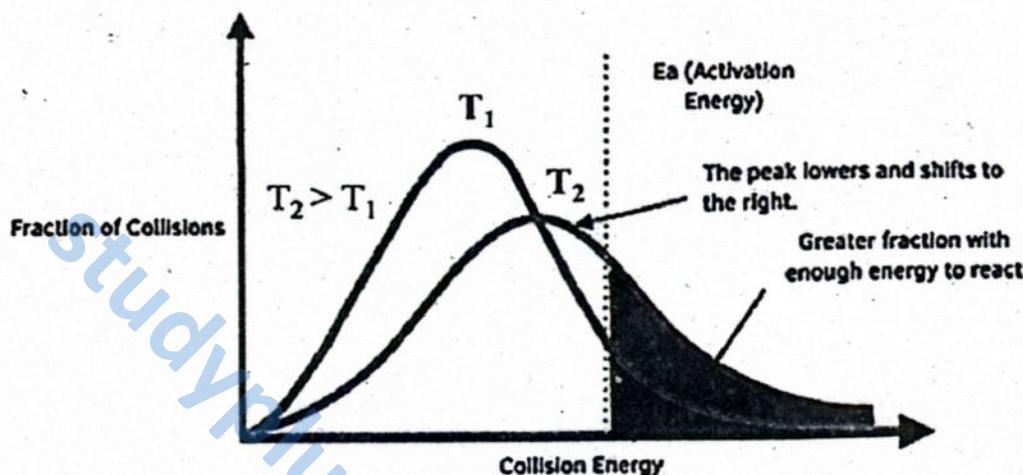


Fig. The Boltzmann distribution of molecular energies at temperatures T_1 and T_2

Quick Check 7.4

a) What is the Boltzmann distribution curve?

See page 214. Effect of temperature

b) Explain why a 10°C rise in temperature approximately doubles the rate of a reaction.

When the temperature of a reaction is raised, the average kinetic energy of the particles increases.

For 10°C rise in temperature, the number of molecules having activation energy becomes double. Thus, the rate of reaction is also becomes double.



3. Surface area

By increasing surface area, contact between reacting molecules increases. Hence rate of reaction increases.

Example

CaCO_3 in big pieces react slowly with H_2SO_4 , but it reacts rapidly in powdered form.

Short Question

How surface area affects the rate of a chemical reaction?

4. Catalyst

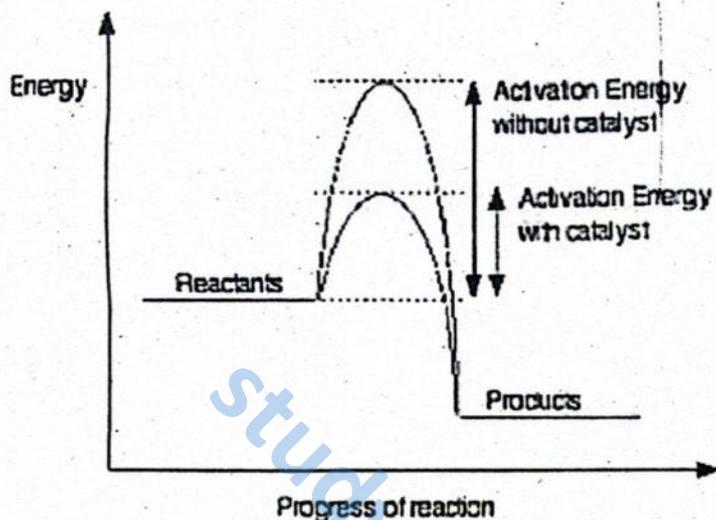
[What is a catalyst? How it increases the rate of reaction?]

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

A catalyst is often present in a very small amount.

Function of a Catalyst

- A catalyst provides a new reaction path with a low activation energy. Thus, a greater number of molecules are now able to get over the new energy barrier and reaction rate increases.



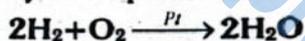
Exercise Q.5 : How does the activation energy profile of an unanalyzed reaction compare with that of the catalyzed reaction?

Exercise Q2(j) : A catalyst lowers the activation energy of a chemical reaction. Illustrate it.

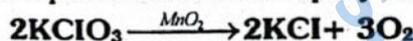
Figure: The energy path diagram for an uncatalyzed and a catalyzed reaction

Examples:

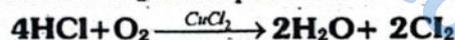
Example 1: the reaction between H_2 and O_2 to form water is very slow at ordinary temperature, but proceeds more rapidly in the presence of small amount of Pt which act as a catalyst.



Example 2: KClO_3 decomposes much more rapidly in the presence of a small amount of MnO_2 .



Example 3: HCl is oxidized to Cl_2 in the presence of CuCl_2 .



ADDITIONAL MCQs

(Answers on Page 229)

- With the 10°C rise in temperature, the rate of reaction
(A) becomes double (B) becomes triple (C) becomes half (D) remains unchanged
- The energy of activation of forward reaction is less than that of backward reaction in.
(A) Endothermic reactions (B) Exothermic reaction
(C) Isothermic reaction (D) None of the above
- The energy of activated complex is:
(A) Greater than the reactants & products (B) Less than the reactants & products
(C) Equal to the products (D) Equal to the reactants
- For reactions involving gaseous reactants, the concentration can be increased by
(A) increasing pressure (B) decreasing pressure
(C) increasing temperature (D) decreasing temperature
- The units of rate of reaction are
(A) mol dm^{-3} (B) $\text{mol dm}^{-3} \text{ s}^{-1}$ (C) $\text{mol dm}^3 \text{ s}$ (D) $\text{mol dm}^{-3} \text{ s}$
- For which type of reactions, the volume change method is suitable to follow the reaction.
(A) ionic reactions (B) solid state reactions (C) gaseous reactions (D) solution phase reactions

CATALYSIS

The process, which takes place in the presence of a catalyst, is called catalysis.

Types of Catalysis

i) Homogeneous Catalysis

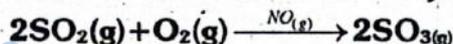
[What is homogeneous catalysis? Give an example]

In this process, the catalyst and the reactants are in the same phase and the reacting system is homogeneous throughout.

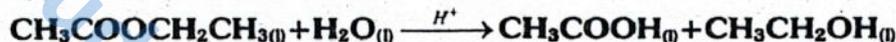
The catalyst is distributed uniformly throughout the system.

Examples

- i. In Lead Chamber process of manufacturing sulphuric acid, the formation of $\text{SO}_{3(g)}$ from $\text{SO}_{2(g)}$ and $\text{O}_{2(g)}$ needs $\text{NO}_{(g)}$ as a catalyst. Both the reactants and the catalyst are gases.



- ii. Esters are hydrolyzed in the presence of H_2SO_4 . Both the reactants and the catalyst are in the solution state.



Did you Know!

Biochemical catalysts, commonly known as enzymes (nature's catalyst) are essential molecules in living organisms' function by lowering the activation energy required for a chemical reaction to proceed, thereby increasing the reaction rate. Enzymes are typically proteins. Factors such as pH, temperature, and the concentration of substrate molecules can influence enzyme activity.

ii) Heterogeneous Catalysis

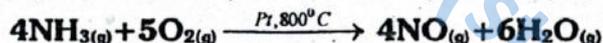
[What is heterogeneous catalysis? Give an example]

In such systems, the catalyst and the reactants are in different phases.

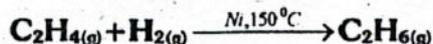
Mostly, the catalysts are in the solid phase, while the reactants are in the gaseous or liquid phase.

Examples:

1. Oxidation of ammonia to NO in the presence of platinum gauze during manufacturing of HNO_3 .



2. Hydrogenation of unsaturated organic compounds are catalysed by finely divided Ni, Pd or Pt.



Differentiate between homogeneous and heterogeneous catalysis.

HOMOGENOUS CATALYSIS		HETEROGENOUS CATALYSIS	
1	In this catalysis, reactants and catalyst are in the same phase.	1	In this catalysis, reactants and catalyst are in the different phase.
2	In this, system remains homogenous during the reaction.	2	In this, system remains heterogeneous during the reaction
3	Mostly liquids and gases are used as homogeneous catalysts.	3	Mostly solid are used as heterogeneous catalysts.
4	Example: $2\text{SO}_{2(g)} + \text{O}_{2(g)} \xrightleftharpoons{\text{NO}} 2\text{SO}_{3(g)}$ NO is homogeneous catalyst	4	Example: $2\text{SO}_{2(g)} + \text{O}_{2(g)} \xrightleftharpoons{\text{V}_2\text{O}_5} 2\text{SO}_{3(g)}$ V_2O_5 is heterogeneous catalyst

Interesting Information

Vitamins are organic compounds that act as catalysts in biochemical reactions, especially when they function as coenzymes. Coenzymes are organic molecules that help enzymes catalyze reactions more efficiently. For example, Vitamin K, is necessary for blood clotting. Low levels of vitamin K can cause bleeding diathesis. A lack of vitamins can disrupt metabolic balance in cells and organisms. Vitamin deficiency is an example of a cofactor deficiency.

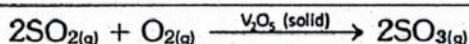
Quick Check 7.5

a) Can a catalyst be consumed in a chemical reaction? Why or why not?

A catalyst increases the rate of reaction. It provides a new reaction path with a low activation energy barrier. It is not changed in mass and composition at the end of reaction. So, it is not consumed in the reaction.



b) Explain whether the reaction below is an example of heterogeneous or homogeneous catalysis:



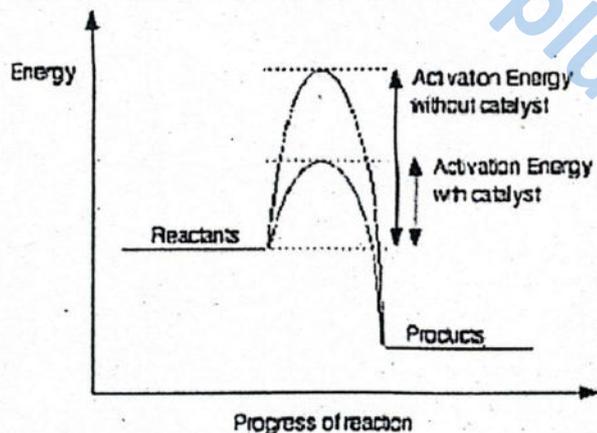
In this reaction both reactants are gases but the catalyst is solid. So, reactants and catalyst are in different phases. Hence, this is an example of heterogeneous catalysis.

c) Draw an energy profile diagram to show a typical uncatalysed reaction and an enzyme catalysed reaction.

On your diagram show:

i) the activation energy for the catalysed and

ii) uncatalysed reactions



ADDITIONAL MCQs

(Answers on Page 229)

- In homogeneous catalysis, which are in same phase?
(A) reactants and products (B) reactants and catalyst
(C) products and catalyst (D) none of these
- A catalyst increases the rate of reacting by
(A) decreasing activation energy (B) decreasing enthalpy
(A) decreasing reactants energy (B) decreasing products energy
- Hydrolysis of ester by sulphuric acid is an example of
(A) homogeneous catalysis (B) heterogeneous catalysis
(C) promotion catalysis (D) activation catalysis



RATE LAW

The representation of rate of a reaction in terms of concentration of the reactants is known as rate law.

- The rate of a chemical reaction at a given temperature may depend on the concentration of one or more reactants and products.
- A rate law is an equation that relates the rate of a reaction to the concentrations of reactants raised to various powers according to the experimental data.

Example and Explanation

- For a general reaction between A and B where 'a' moles of A and 'b' moles of B react to form 'c' moles of C and 'd' moles of D.



The rate equation as:

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

- Where x and y are the experimentally determined values. These may or may not be equal to the coefficient of reactants in the balanced chemical equation, as 'a' and 'b' in the above equation.
- This expression is called rate equation.
- The brackets [] represent the molar concentrations.
- The proportionality constant k is called rate constant for the reaction.

RATE CONSTANT

The specific rate constant of a chemical reaction is the rate of reaction when the concentrations of the reactants are unity

Explanation

- Consider general reaction



The rate equation as:

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

- The proportionality constant k is called rate constant for the reaction.

$$\text{If } [A] = 1 \text{ mol dm}^{-3} \text{ and } [B] = 1 \text{ mol dm}^{-3}$$

$$\text{Then Rate} = k \times 1^x \times 1^y = k$$

- The rate constant is defined as the rate of reaction when the concentrations of the reactants are unity.
- Under the given conditions, k remains constant, but it changes with temperature.

Short Question

What is specific rate constant?

OR

Explain what is meant by the specific rate (rate constant) of a reaction and how it is represented in rate equation [Quick Check 7.6(b)]

ADDITIONAL MCQs

(Answers on Page 229)

10. Specific rate constant is equal to rate of reaction, when concentration of reactants are:

- (A) Zero (B) Four (C) Three (D) Unity

11. The unit of rate constant depends on

- (A) order of reaction (B) molecularity (C) number of reactants (D) all of above

12. The unit of rate constant for a first order reaction is

- (A) mol dm^{-3} (B) $\text{mol dm}^{-3} \text{ s}^{-1}$ (C) $\text{mol dm}^{-3} \text{ s}^{-1}$ (D) s^{-1}



REACTION ORDER

Exercise Q.3 Relate the order of a reaction to the rate law for the reaction. How do you distinguish between zero order, first order and second order reaction?

The order of a reaction with respect to a specific reactant is the exponent applied to that reactant's concentration within the rate equation

Example and Explanation

- Consider a general reaction



The rate equation is:

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

In this equation

x is the order of reaction with respect to A

y is the order of reaction with respect to B

The sum is $x + y = n$. The sum ' n ' is called overall order of reaction

- x and y are the experimentally determined values.
- The order of a reaction defines how the reactant concentration influences its rate.
- For a single-reactant, the order is simply the concentration's power in the rate equation.
- The order of reaction provides valuable information about the mechanism of a reaction.

Example

Consider the reaction of nitrogen (II) oxide (NO) with H_2 and oxygen:



The experimental rate equation is

$$\text{rate} = k[H_2][NO]^2$$

This rate equation shows that the reaction is

- first-order with respect to H_2
- second-order with respect to NO
- third-order overall ($1 + 2 = 3$)

Keep in Mind

The order of reaction is given by the sum of all the exponents to which the concentrations in the rate equation are raised. It is important to note that the order of a reaction is an experimentally determined quantity and cannot be inferred simply by looking at the reaction equation. The sum of the exponents in the rate equation may or may not be the same as in a balanced chemical equation

Quick Check 7.6

a) How order of reaction is derived from the rate law?

Consider the reaction



The experimental rate law is

$$\text{rate} = k[H_2][NO]^2$$

This rate equation shows that the reaction is

- first-order with respect to H_2
 - second-order with respect to NO
- So, the overall order of reaction = $1 + 2 = 3$
Hence, the reaction is third order



b) Explain what is meant by the specific rate (rate constant) of a reaction and how it is represented in rate equation.

See 219

Types of Reaction Order

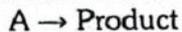
Zero Order Reaction

[What is a zero order reaction? Give an example]

The rate of a zero order reaction is independent of the concentration of the reactants.

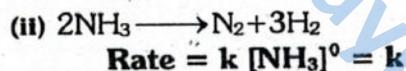
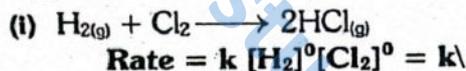
A change in the concentration has no effect on the speed of the reaction.

For the general reaction:



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

Examples:



(iii) Photochemical reactions are usually zero order

First Order Reaction

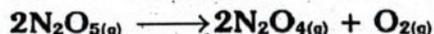
[What are first order reactions? Give an example]

A reaction for which sum of the exponents of the concentrations in the rate equation is 1.

In these reactions, there may be multiple reactants present, but concentration of only one reactant will change during the reaction.

Example:

Decomposition of nitrogen pentoxide involves the following equation.



The experimentally determined rate equation for this reaction is:

$$\text{Rate} = k[\text{N}_2\text{O}_5] \quad (n = 1)$$

This equation suggests that the reaction is first order with respect to the concentration of N_2O_5 .

Second Order Reaction

[What are second order reactions? Give an example]

A reaction for which sum of the exponents of the concentrations in the rate equation is 2.

A reaction can be second order in two ways.

- When rate depends either on the concentration of one reactant raised to the second power
 $\text{Rate} = k[A]^2 \quad (n = 2)$
- When rate depends upon the concentrations of two different reactants, each raised to the first power.
 $\text{Rate} = k [A]^1[B]^1 \quad (n = 1+1 = 2)$
- The simpler type of reaction involves one kind of reactant molecule.

Example

Oxidation of nitric oxide with ozone is first order with respect to NO and first order with respect to O_3 .



$$\text{Rate} = k[\text{NO}][\text{O}_3]$$

The sum of the individual orders is two ($1+1=2$). Hence it is a second order reaction.

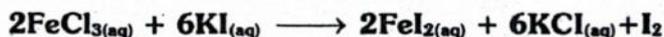
Third Order Reaction

[What are third order reactions? Give an example]

A reaction for which sum of the exponents of the concentrations in the rate equation is 3.

$$\begin{aligned}\text{Rate} &= k[A]^3 & (n = 3) \\ \text{or Rate} &= k[A]^2[B]^1 & (n = 2+1 = 3) \\ \text{or Rate} &= k[A]^1[B]^1[C]^1 & (n = 1+1+1 = 3)\end{aligned}$$

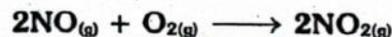
Example 1



This reaction involves eight reactant molecules. However, experimentally it is a third order reaction.

$$\text{Rate} = k[\text{FeCl}_3][\text{KI}]^2$$

Example 2



$$\text{Rate} = k[\text{NO}]^2[\text{O}_2]^1$$

Fractional Order Reaction

[Order of reaction can be in fraction. Justify?]

A reaction in which the sum of exponents of rate equation is in fraction.

Example: Consider the formation of carbon tetrachloride from chloroform.



$$\text{Rate} = k[\text{CHCl}_3][\text{Cl}_2]^{1/2} \quad n = 1 + 1/2 = 1.5$$

Reactions involving free radicals often show fractional orders.

UNITS OF RATE CONSTANT

- The rate constant is specific for a particular reaction at a certain temperature.
- The units of concentrations are mol dm^{-3} . The units of the rate of reaction are $\text{mol dm}^{-3} \text{ s}^{-1}$.
- The units for k depend on the order of the reaction and the units of time.

General Equation for n^{th} order reaction is:

$$\text{Rate} = k [\text{Reactants}]^n$$

where n = order of reaction

So,

$$k = \frac{\text{Rate}}{[\text{Reactants}]^n} = \frac{\text{mol dm}^{-3} \text{ s}^{-1}}{(\text{mol dm}^{-3})^n}$$

$$\text{or } k = (\text{mol dm}^{-3})^{1-n} \text{ s}^{-1}$$

$$\text{or } k = (\text{concentration})^{1-n} \text{ s}^{-1}$$

This equation can be used to determine units of any order of reaction.

- For a zero order reaction ($n = 0$),
 $k = (\text{mol} \cdot \text{dm}^{-3})^{1-0} \text{ s}^{-1}$
 $k = \text{mol} \cdot \text{dm}^{-3} \text{ s}^{-1}$

- For a first order reaction ($n = 1$), the rate is directly proportional to the concentration of one reactant.

$$\text{Thus, } k = (\text{concentration})^{1-n} \text{ s}^{-1}$$

$$k = (\text{mol} \cdot \text{dm}^{-3})^{1-1} \text{ s}^{-1}$$

$$k = (\text{mol} \cdot \text{dm}^{-3})^0 \text{ s}^{-1}$$

$$k = \text{s}^{-1}$$

Therefore, the units of k for a first order rate constant are s^{-1} .

- For a second order reaction ($n = 2$),

$$\text{Thus, } k = (\text{concentration})^{1-n} \text{ s}^{-1}$$

$$k = (\text{mol} \cdot \text{dm}^{-3})^{1-2} \text{ s}^{-1}$$

$$k = \text{mol}^{-1} \text{ dm}^3 \text{ s}^{-1}$$

$$k = \text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$$

The units of k for a second order rate constant are $\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$.

Short Question

Determine the units of rate constant for a second order reaction?

- For a third order reaction ($n = 3$),

$$\text{Thus, } k = (\text{concentration})^{1-n} \text{ s}^{-1}$$

$$k = (\text{mol} \cdot \text{dm}^{-3})^{1-3} \text{ s}^{-1}$$

$$k = \text{mol}^{-2} \text{ dm}^6 \text{ s}^{-1}$$

$$k = \text{dm}^6 \text{ mol}^{-2} \text{ s}^{-1}$$

Therefore, the units of k for a third order rate constant are $\text{dm}^6 \text{ mol}^{-2} \text{ s}^{-1}$.

Short Question

Determine the units of rate constant for a third order reaction?

Quick Check 7.7

a) Calculate the overall order of reactions which have the rate expressions:

i) $\text{rate} = k[\text{NO}]^2 [\text{NH}_3]^0$

This rate equation shows that the reaction is

- Second order with respect to NO
- zero order with respect to NH_3

So, the overall order of reaction = $2 + 0 = 2$

Hence, the reaction is second order

ii) $\text{rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]^2$

This rate equation shows that the reaction is

- first order with respect to BrO_3^-
- first order with respect to Br^-
- second order with respect to H^+

So, the overall order of reaction = $1 + 1 + 2 = 4$

Hence, theoretically it is 4th order



b) Why do you think chemists want to know the order of a reaction and the rate constant for a reaction?

The order of reaction help find out the mechanism of reaction

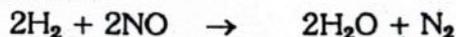
The rate constant describes the speed of reaction at a given temperature.

Thus, for any reaction a chemist always determine both these parameters

c) Why the sum of the coefficients of a balanced chemical equation is not necessarily important to give the order of a reaction?

The order of reaction is determined from the experimental rate law. Therefore, the co-efficients in balanced chemical equation has no relation with order of reaction.

e.g., consider the reaction



The experimental rate law of this reaction is

$$\text{Rate} = k[\text{H}_2][\text{NO}]^2$$

Hence, its order is $1+2 = 3$ (third order reaction)

However, the sum of co-efficients in balanced chemical equation = $2+2 = 4$.

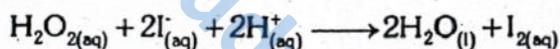
DETERMINATION OF RATE CONSTANT

Exercise Q.4 How do you find the numerical value of a rate constant by initial rate and half-life methods?

The rate constant (k) of a reaction can be calculated using the following two methods:

1. Initial Concentration Method

In the presence of hydrogen ions, hydrogen peroxide, H_2O_2 , reacts with iodide ions to form water and iodine:



- The rate equation for this reaction is:

$$\text{rate of reaction} = \frac{k[\text{H}_2\text{O}_2]}{[\text{I}^-]}$$

- The progress of the reaction can be followed by measuring the initial rate of formation of iodine.
- The data is given in the table below.
- The k is calculated using the data for experiment 1.

Step 1: Write out the rate equation.

$$\text{rate of reaction} = \frac{k[\text{H}_2\text{O}_2]}{[\text{I}^-]}$$

Step 2: Rearrange the equation in terms of k

$$k = \frac{\text{rate} \times [\text{I}^-]}{[\text{H}_2\text{O}_2]}$$

Step 3: Substitute the values

$$k = \frac{3.5 \times 10^{-6} \times (0.0100)}{(0.0200)} = 1.7 \times 10^{-2} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$$

Table 7.2: Effect of change in concentrations of reactants on the rate of reaction

Experiment	$[\text{H}_2\text{O}_2]/\text{mol dm}^{-3}$	$[\text{I}^-]/\text{mol dm}^{-3}$	$[\text{H}^+]/\text{mol dm}^{-3}$	Initial rate of reaction/ $\text{mol dm}^{-3} \text{ s}^{-1}$
1	0.0200	0.0100	0.0100	3.5×10^{-6}
2	0.0300	0.0100	0.0100	5.3×10^{-6}
3	0.0050	0.0200	0.0200	1.75×10^{-6}

The concentration of hydrogen ions is ignored because $[\text{H}^+]$ does not appear in the rate equation. The reaction is zero order with respect to $[\text{H}^+]$.

2. Half-Life Method

[Define half-life time of a reaction. How half-life time is related to the rate constant of a first order reaction?]

Half-life is the time taken for the concentration of a reactant to fall to half of its original value.

- It is denoted by $t_{1/2}$
- To determine rate constant (k) half-life time is measured for a reaction.
- For a first-order, then the rate constant and the half-life of the reaction are related as

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

Example

- A sample of hydrogen peroxide has a half-life of 2 hours.
- It decomposes in a first-order reaction.
- Half-life of the reaction = ($t_{1/2}$) = 2 hours = $2 \times 60 \times 60 = 7200$ s
- Hence rate constant is

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{7200 \text{ s}} = 9.6 \times 10^{-5} \text{ s}^{-1}$$

Short Question

How rate constant of a first order reaction can be determined by half-life method?

Sample Problem 7.2

The first-order reaction cyclopropane to propene, for which the half-life is 17.0 min, Calculate the rate constant of this reaction.

Solution:

Half-life of the reaction = $t_{1/2} = 17 \text{ min} = 17 \times 60 = 1020 \text{ s}$

Hence rate constant is

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

$$k = \frac{0.693}{1020 \text{ s}} = 679 \times 10^{-4} \text{ s}^{-1}$$

Quick Check 7.8

Consider a first-order reaction with a half-life of 15 minutes. If the initial concentration of the reactant is $0.100 \text{ mol dm}^{-3}$, calculate the rate constant (k) for the reaction.

Solution

Half-life of the reaction = $t_{1/2} = 15 \text{ min} = 15 \times 60 = 900 \text{ s}$

Hence rate constant is

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

$$k = \frac{0.693}{900 \text{ s}} = 7.7 \times 10^{-4} \text{ s}^{-1}$$



13. If 75% of any given amount of radioactive element disintegrates in 60 minutes, the half life of radioactive element is:
 (A) 20 minutes (B) 30 minutes (C) 45 minutes (D) 25 minutes
14. When a reaction proceeds in sequence of steps, the overall rate is determined by:
 (A) fastest step (B) slowest step (C) molecularity of all steps (D) order of different steps
15. Half-life period for ${}_{92}^{235}\text{U}$ is:
 (A) 710 million years (B) 720 million years (C) 810 million years (D) 820 million years
16. The half life period of ${}_{92}^{235}\text{U}$ is 710 million years. 100 kg of sample of ${}_{92}^{235}\text{U}$ will reduce to 50kg in
 (A) 710 million years (B) 1420 million years (C) 2130 million years (D) 2840 million years
17. The half life period of ${}_{6}^{14}\text{C}$ is 5760 years. 100 mg of sample of ${}_{6}^{14}\text{C}$ will reduce to 25 mg in
 (A) 11520 years (B) 2880 years (C) 576600 years (D) 5760 years
18. All radioactive disintegration nuclear reactions are of
 (A) first order (B) second order (C) Third order (D) zero order
19. The unit of rate constant for zero order reaction is:
 (A) $\text{dm}^3 \text{s}^{-1}$ (B) $\text{mole dm}^{-3} \text{s}^{-1}$ (C) $\text{dm}^3 \text{mole}^{-1} \text{s}^{-1}$ (D) mole s^{-1}
20. $2\text{A} + \text{B} \rightarrow \text{Product}$: If the reactant 'B' is in excess, the order of reaction with respect to 'A' in given rate law, $\text{Rate} = k[\text{A}]^2[\text{B}]$ is:
 (A) 2nd order reaction (B) 1st order reaction (C) Pseudo 1st order reaction (D) 3rd order reaction
21. Which one of the following correctly represents the units of the rate constant, k for a first order reaction?
 (A) s^{-1} (B) $\text{mol dm}^{-3} \text{s}$ (C) $\text{mol dm}^{-3} \text{s}^{-1}$ (D) $\text{mol}^{-1} \text{dm}^3 \text{s}$
22. The two steps involved in the gas phase reaction
 $\text{X} + 2\text{Y} \rightarrow \text{XY}_2$ are shown below
 i. $\text{X} + \text{Y} \xrightarrow{\text{slow}} \text{XY}$
 ii. $\text{XY} + \text{Y} \xrightarrow{\text{fast}} \text{XY}_2$ What is the rate equation for the overall reaction?
 (A) $\text{Rate} = K [\text{x}]^0 [\text{y}]^1$ (B) $\text{Rate} = K [\text{x}]^0 [\text{y}]^2$ (C) $\text{Rate} = K [\text{x}]^1 [\text{y}]^1$ (D) $\text{Rate} = K [\text{xy}]^1 [\text{y}]^1$
23. Decomposition of Ozone take place according to the following equation. $2\text{O}_3(\text{g}) \rightarrow 3\text{O}_2(\text{g})$,
 Rate equation for the reaction is $\text{Rate} = k[\text{O}_3]^2[\text{O}_2]^{-1}$, What is the order of the reaction?
 (A) 3 (B) Zero (C) 2 (D) 1
24. The unit of rate constant for zero order reaction is:
 (A) $\text{dm}^3 \text{s}^{-1}$ (B) $\text{mole dm}^{-3} \text{s}^{-1}$ (C) $\text{dm}^3 \text{mole}^{-1} \text{s}^{-1}$ (D) mole s^{-1}
25. The order of reaction for suppose a single step reaction $\text{NO} + \text{O}_3 \rightarrow \text{NO}_2 + \text{O}_2$ is
 (A) two (B) three (C) one (D) zero
26. A certain reaction is first order in H_2 and half order in Br_2 the rate law for the reaction is
 (A) $\text{Rate} = k [\text{H}_2] [\text{Br}_2]^{1/2}$ (B) $\text{Rate} = k [\text{H}_2] [\text{Br}_2]^1$
 (C) $\text{Rate} = k [\text{H}_2]^{1/2} [\text{Br}_2]^1$ (D) $\text{Rate} = k [\text{H}_2]^{1/2} [\text{Br}_2]^{1/2}$
27. The equation for the formation of CCl_4 is: $\text{CHCl}_3 + \text{Cl}_2 \rightarrow \text{CCl}_4 + \text{HCl}$
 The rate equation is $\text{Rate} = K[\text{CHCl}_3][\text{Cl}_2]^{1/2}$, The order reaction is
 (A) -1/2 (B) 3/2 (C) 2 (D) 3/2
28. When a reaction occurs in many steps than the slowest step is
 (A) Mechanism determining step (B) Enthalpy determining step
 (C) Rate determining step (D) None of the above.
29. Half-life period of a first order reaction is independent of
 (A) Initial concentration of the compound (B) Conditions of temperature
 (C) Presence of catalyst. (D) All the above

REACTION MECHANISM

A reaction mechanism is a detailed, step by step description of how a chemical reaction occurs at the molecular level to yield the product(s).

- Unlike the overall balanced chemical equation, which only shows the reactants and products, the reaction mechanism reveals the actually happening individual steps (called elementary steps) that lead to the formation of products.
- Each of these steps represents a single molecular event, such as the breaking or forming of bonds.
- Many reactions do not occur in a single step, but rather proceed through a series of steps.
- Each step is called an elementary reaction and is directly caused by the collision of atoms, ions or molecules.
- The number of reactant molecules involved in an elementary step is called its molecularity.
- A unimolecular elementary reaction involves only a single reactant molecule.

Example

The example of a unimolecular reactions is the decomposition of N_2O_5 .



An elementary reaction that involves two atoms, ions or molecules and is called bimolecular. For example;



A termolecular reaction step involves the simultaneous reaction of three molecules. Such reactions are rare. An example is the reaction between oxygen molecules and atomic oxygen to form ozone in the stratosphere or during smog formation.



Intermediates are short lived species (ions or free radicals) that are produced in one step of the mechanism and consumed in a subsequent step. They do not appear in the overall balanced equation because they are not stable products.

Rate Determining Step

[What is rate determining step of a reaction? Give an example]

In many reaction mechanisms, one step is significantly slower than all the others, this step is called the rate-determining step.

- This step controls the overall rate of the reaction because it limits the speed at which the reaction can proceed.
- The balanced equation for the overall reaction is equal to the net result of all the individual steps.
- In a chemical reaction, any step that occurs after the rate-determining step will not affect the rate, provided that it is compared with the rate-determining step.
- So the atoms, ions or molecules taking part in the mechanism after the rate-determining step do not appear in the rate expression.
- All reactants that appear in the rate-determining step will also appear in the rate equation. Because the rate-determining step limits the rate of the overall reaction.
- The order of a reaction can be obtained from the rate determining step.

Example 1

- Consider the following reaction:



- In first three experiments the concentration of H_2 is increased by keeping the concentration of NO constant.
- By doubling the concentration of H_2 , the rate is doubled. By tripling the concentration of H_2 , the rate is tripled.
- So, the rate of reaction is directly proportional to the first power of concentration of H_2 .



$$\text{Rate} \propto [\text{H}_2]$$

- In the next three experiments, the concentration of H_2 is kept constant.
- By doubling the concentration of NO , the rate increases four times. By tripling the concentration of NO the rate is increased nine times.
- So, the rate is proportional to the square of concentration of NO .

Table 7.3: Effect of change in concentrations of reactants on the rate of reaction

$[\text{NO}]$ in (mol dm^{-3})	$[\text{H}_2]$ in (mol dm^{-3})	Initial rate (atm min^{-1})
0.006^3	0.001^3	0.25^1
0.006	0.002	0.050
0.006	0.003	0.0075
0.001	0.009	0.0063
0.002	0.009	0.025
0.003	0.009	0.056

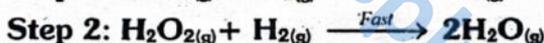
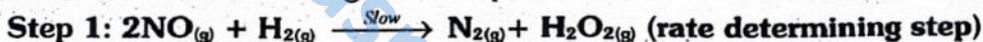
$$\text{Rate} \propto [\text{NO}]^2$$

- The overall rate equation of reaction is,

$$\text{Rate} \propto [\text{H}_2][\text{NO}]^2$$

$$\text{or Rate} = k[\text{H}_2]^1[\text{NO}]^2$$

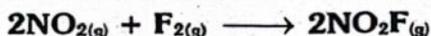
- Hence, the reaction is a third order one.
- The final equation is the rate law for this reaction.
- The possible mechanism consisting of two steps for the reaction is as follows.



- The step 1 is slow and rate determining.

Example 2

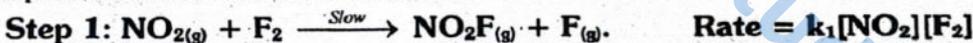
- Consider the reaction between nitrogen dioxide and fluorine gas:



- This reaction is first order in NO_2 , first order in F_2 and second order overall.
- The experimental rate law is first order in NO_2 and in F_2 :

$$\text{Rate} = k [\text{NO}_2][\text{F}_2] \quad \text{(Observed)}$$

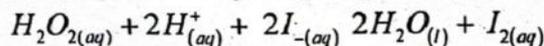
- The accepted mechanism for the reaction is:



- The first step is slow and determines the rate, in agreement with the observed rate expression.
- The second and fast step does not affect the reaction rate because fluorine atoms react with NO_2 as soon as they are produced.

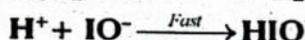
Quick Check 7.9

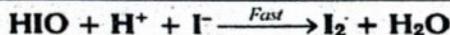
An acidified solution of hydrogen peroxide reacts with iodide ions.



The rate equation for this reaction is $\text{rate} = [\text{H}_2\text{O}_2][\text{I}^-]$

The mechanism below has been proposed for this reaction.





Explain why this mechanism is consistent with the rate equation

For any reaction, slow step is the rate determining step. Thus, the order of reaction depends upon this step.

In given slow step one molecule of H_2O_2 and one ion of I^- is taking part in the reaction.

Thus, the rate law from slow step is

$$\text{Rate} = k [\text{H}_2\text{O}_2] [\text{I}^-]$$

So, this equation is same as the experimental rate law.

Hence, the given mechanism is consistent with the rate equation.

ADDITIONAL MCQs

(Answers on Page 229)

30. The maximum point in a potential energy diagram represents
 (A) reactants (B) products (C) intermediate (D) activated complex
31. If a potential energy diagram has more than one maxima, then it shows that
 (A) the reaction is multi-step (B) intermediates are involved
 (C) the reaction has a slow step (D) all of the above
32. In a potential energy diagram having more than one maxima, the maxima with highest energy of activation is the
 (A) Mechanism determining step (B) Enthalpy determining step
 (C) Rate determining step (D) None of the above.

ANSWERS TO ADDITIONAL MCQs

Q#	Ans														
1	A	2	A	3	A	4	A	5	B	6	C	7	A	8	A
9	A	10	D	11	A	12	D	13	B	14	B	15	A	16	A
17	A	18	A	19	B	20	A	21	A	22	C	23	D	24	B
25	A	26	A	27	B	28	C	29	A	30	D	31	A	32	C



Q1 MULTIPLE CHOICE QUESTIONS

- I. The rate of reaction:**
- Increases as the reaction proceeds
 - Decreases as the reaction proceeds
 - Remains the same as the reaction proceeds
 - May decrease or increase as the reaction proceeds
- II. Increasing the temperature of a chemical reaction increases the rate of reaction because:**
- Both the collision frequency and collision energies of reactant molecules increase
 - Collision frequency of reactant molecules increases
 - Activation energy increase
 - Activation energy decrease
- III. Consider two reactions with different activation energies at the same temperature. The reaction with the lower activation energy will have:**
- A smaller rate constant
 - A larger rate constant
 - The same rate constant
 - A rate constant that depends on the enthalpy change
- IV. The order of a chemical reaction, that is independent of concentration is:**
- Second order reaction
 - First order reaction
 - Zero order reaction
 - Pseudo first order reaction
- V. On a Boltzmann distribution curve, the area under the curve represents:**
- Activation energy of the reaction.
 - Total number of molecules in the sample.
 - Average kinetic energy of the molecules.
 - Rate constant of the reaction.
- VI. On a Boltzmann distribution curve, the activation energy (E_a) is represented by:**
- The height of the peak
 - The area under the entire curve
 - A vertical line drawn at a specific kinetic energy value
 - The difference between the peak and the X-axis
- VII. If we double the concentration of a reactant, the rate increases by four times, the reaction is:**
- Second order
 - First order
 - Third order
 - Zero order
- VIII. The rate determining step in a multi-step reaction is:**
- Always the first step
 - Always the last step
 - The slowest step
 - The fastest step
- IX. The reaction $\text{NO}_2 + \text{CO} \rightarrow \text{NO} + \text{CO}_2$ occurs in two steps. What is the rate law equation for this reaction?**
- $$2\text{NO}_2 \rightarrow \text{NO} + \text{NO}_3 \quad (k_1) \quad \text{slow}$$
- $$\text{NO}_2 + \text{CO} \rightarrow \text{CO}_2 + \text{NO}_2 \quad (k_2) \quad \text{fast}$$
- $R = k_1 [\text{NO}_2]^3$
 - $R = k_2 [\text{NO}_3][\text{CO}]$
 - $R = k_1 [\text{NO}_2]$
 - $R = k_1 [\text{NO}_2]^2$
- X. How does the presence of a catalyst affect the rate of a chemical reaction?**
- It always decreases the rate of the reaction.
 - It always increases the rate of the reaction.
 - It increases the rate of the forward and decreases the rate of the reverse reaction.
 - It increases the rate of both the forward and reverse reactions.
- XI. On an energy profile diagram, the presence of a catalyst is represented by:**
- A higher peak representing the activation energy.
 - A lower peak representing the activation energy.
 - A change in the energy level of the reactants or products.
 - A shift in the equilibrium position.

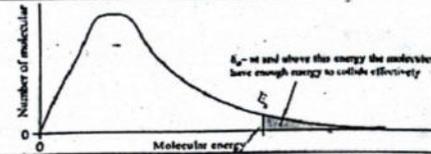
XII. The units of the rate constant (k) for a reaction depend on the:

- Activation energy of the reaction
- Temperature of the reaction
- Overall order of the reaction
- Stoichiometry of the balanced chemical equation

XIII. A first-order reaction has a half-life ($t_{1/2}$) of 20 minutes. What is the value of its rate constant (k)?

- 0.05 min^{-1}
- 0.693 min^{-1}
- 0.0347 min^{-1}
- 13.86 min^{-1}

ANSWERS TO MULTIPLE CHOICE QUESTIONS

No.	Ans	EXPLANATION
I.	b	According to the law of mass action, rate of a reaction is directly proportional to the product of concentrations of reactants. As the reaction proceeds, the concentration of reactants decreases. Hence, rate of reaction also decreases.
II.	a	Increase in temperature increases the collision frequency and collision energies of reactant molecules. Thus, rate of reaction increases.
III.	b	Increase in activation energy decreases the rate of reaction. Hence, reaction with lower activation energy will have a larger rate constant.
IV.	c	Consider a reaction $A \rightarrow \text{Products}$ If this reaction is a zero order reaction then its rate equation is given as $\text{Rate} = k[A]^0 = k$ It means that the rate of reaction is independent of the concentration of reactant.
V.	b	The area under the curve represent the total number of molecules in the sample.
VI.	c	The activation energy is shown by a vertical line drawn at a specific kinetic energy value. 
VII.	a	If conc. of reactant is doubled then rate of reaction becomes four times. Thus, rate depends upon 2^{nd} power of concentration of reactant. So, $\text{Rate} \propto [\text{reactant}]^2$, if conc. is doubled then $\text{Rate} \propto [2]^2 = 4$ times Hence, the reaction is second order.
VIII.	a	The slowest step is the rate determining step.
IX.	a	Since, the slowest step is the rate determining step. Hence, according to slow step the rate law for the reaction is $R = k_1 [\text{NO}_2]^2$
X.	b	By definition, a catalyst increases the rate of reaction without being consumed in the reaction.
XI.	b	A catalyst provides an alternate path with low activation energy. Hence, in an energy profile diagram it is represented by a relatively lower peak representing the activation energy.
XII.	c	The units of rate constant depends upon overall order of reaction. e.g. for zero order reaction $\text{Rate} = k[A]^0 = k$. Hence, units of 'k' is $\text{mol dm}^{-3} \text{s}^{-1}$
XIII.	c	For a first order reaction, $t_{1/2} = 0.693/k$ or $k = 0.693/t_{1/2} = 0.693/20 = 0.0347 \text{ min}^{-1}$

Q2 SHORT QUESTIONS

(a) What do you understand by the rate of a reaction?

It is the change in concentration of reactants or products divided by the time taken for the change

$$\text{Rate} = \frac{\text{Change in conc. of the substance}}{\text{Time taken for change}} = \frac{\Delta x}{\Delta t}$$

where Δx is a very small change in concentration of a reactant or a product in a very small time interval Δt .

(b) Give the difference between enthalpy change of reaction and energy of activation of reaction

Enthalpy Change	Activation Energy
It is the heat absorbed or released during a reaction at constant pressure.	It is the minimum energy required to form an activated complex and thus to start a chemical reaction.
It is denoted by ΔH	It is denoted by E_a
It is the energy difference between reactants and products.	It is the energy difference between reactants and the transition state (activated complex).
It gives the net heat transfer during a reaction. It shows the endothermic or exothermic nature of reaction.	It is related to the rate of reaction. Higher the E_a , slower the rate of reaction.
It is a thermodynamic property.	It is a kinetic property.
Catalysts do not affect ΔH .	Catalysts lower E_a . Thus, rate of reaction increases.

(c) Differentiate clearly between order and molecularity of a reaction.

ORDER OF REACTION	MOLECULARITY
1 It is the number of atoms, ions or molecules whose concentration changes during reaction.	1 It is the number of atoms, ions or molecules used to form activated complex.
2 It is assigned to the reaction as a whole.	2 For multistep reactions, it is assigned to each step separately
3 It can be zero	3 It cannot be zero
4 It can be in fraction	4 It cannot be in fraction
5 It is determined experimentally	5 It is assigned theoretically
6 The reactions are classified as zero order, first order etc.	6 The reactions are classified as unimolecular, bimolecular etc.

(d) Why the instantaneous rate changes during a reaction?

According to the law of mass action, the rate of a chemical reaction is directly proportional to the concentration of reactants.

When the reaction starts, the concentration of reactants is high, therefore, rate of reaction is fast. As the concentration of reactants is decreased, the rate of a reaction is also decreased. At the end of reaction, the reaction becomes very slow. Hence, instantaneous rate of reaction is changes during a reaction.

(e) Briefly summarize the effects of temperature and surface area on the rates of reactions

(f) Justify that the radioactive decay is always a first order reaction

The rate of radioactive decay depends on the amount of radioactive substance. Since only one substance (reactant) is involved in this process, therefore, it is always a first order reaction. Moreover, the half-life time for the radioactive decay of a particular substance is also constant. Hence, it is a first order process.

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

For a general second order reaction, the rate equation is

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

If conc. is doubled then $\text{Rate} \propto [2]^2 = 4$ times.

Hence, the rate of reaction will increase four times of the original rate.

If conc. is reduced to one half then $\text{Rate} \propto [1/2]^2 = 1/4$ times.

Hence, the rate of reaction will decrease to $1/4$ th of its original rate.

(h) What is meant by half-life and what is it used for?

Half-life is the time taken for the concentration of a reactant to fall to half of its original value.

It is denoted by $t_{1/2}$

It can be used to determine the rate constant (k) for a reaction.

Example:

If the reaction is first-order, then the rate constant and the half-life of the reaction are related in the following way:

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

Thus, knowing $t_{1/2}$, the rate constant (k) for a first order reaction can be determined.

(i) Why does wood burn more rapidly in pure oxygen than in air?

According to the law of mass action, the rate of a chemical reaction is directly proportional to the concentration of reactants.

In air, oxygen is about 21% by volume. In pure oxygen, the concentration of oxygen is increased. Hence, according to Law of mass action, the wood burn more rapidly in pure oxygen than in air.

(j) A catalyst lowers the activation energy of a chemical reaction. Illustrate it.

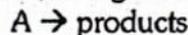
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(k) The rate constant for a certain reaction is $3.5 \times 10^{-4} \text{ s}^{-1}$ at 25°C . What is the order of the reaction? Explain based on the units of the rate constant.

This is a first order reaction because for a first order reaction the unit of rate constant is s^{-1}

Proof

Consider a general first order reaction



Rate equation is

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

where 'k' is the rate constant.

$$\text{So } k = \frac{\text{Rate}}{[A]} = \frac{\text{mol dm}^{-3}\text{s}^{-1}}{\text{mol dm}^{-3}} = \text{s}^{-1}$$

- (l) **If the initial concentration of the reactant is 0.50 mol dm^{-3} , calculate the initial rate of the reaction.**

From above data

$$\text{Rate constant} = k = 3.5 \times 10^{-4} \text{ s}^{-1}$$

$$\text{Initial conc. of reactant} = [A] = 0.50 \text{ mol dm}^{-3}$$

The initial rate of reaction is given by

$$\text{Rate} = k[A] = 3.5 \times 10^{-4} \text{ s}^{-1} \times 0.50 \text{ mol dm}^{-3} = 1.75 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$$

- (m) **How would the rate of this reaction change if the concentration of the reactant were doubled?**

Since it is a first order reaction. Thus, if the concentration of the reactant is doubled then the rate of reaction is also doubled.

$$\text{i.e., Rate} = k[A] \text{ or Rate} \propto [2]^1 = 2 \text{ times}$$

- (n) **A certain first-order reaction has a rate constant of $2.5 \times 10^{-3} \text{ s}^{-1}$. Calculate the half-life of the reaction in minutes.**

$$k = 2.5 \times 10^{-3} \text{ s}^{-1}$$

So, for a first order reaction

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{2.5 \times 10^{-3} \text{ s}^{-1}} = 277.2 \text{ s}$$

- (o) **A radioactive isotope decays by a first-order process with a half-life of 12 hours. Calculate the rate constant for the decay in s^{-1} .**

$$t_{1/2} = 12 \text{ h} = 12 \times 60 \times 60 = 43200 \text{ s}$$

The radioactive decay is a first order reaction. So, for a first order reaction

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

$$\text{or } k = \frac{0.693}{t_{1/2}} = \frac{0.693}{43200} = 1.604 \times 10^{-5} \text{ s}^{-1}$$

DESCRIPTIVE QUESTIONS

- Q.3 Relate the order of a reaction to the rate law for the reaction. How do you distinguish between zero order, first order and second order reaction?**

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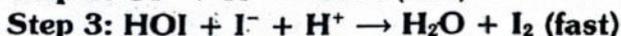
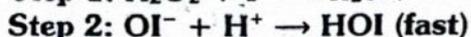
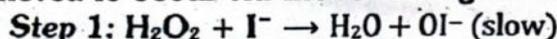
- Q.4 How do you find the numerical value of a rate constant by initial rate and half-life methods?**

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- Q.5 How does the activation energy profile of an unanalyzed reaction compare with that of the catalyzed reaction?**

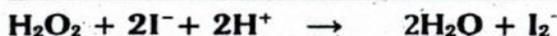
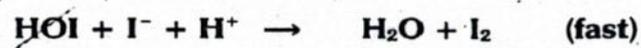
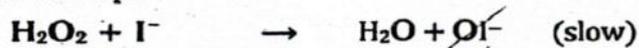
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- Q.6 The reaction between hydrogen peroxide (H_2O_2) and iodide ions (I^-) in acidic solution is believed to occur via the following mechanism:**



- i) Write the overall balanced equation for the reaction.

Add all steps



- ii) Identify any intermediates and catalysts in this mechanism.

Catalyst: There is no substance which is used as a reactant in the first step and regenerates later. Thus, there is no catalyst involved in the reaction.

Reaction intermediate: The OI^- ion is produced in the first step. However, it consumes in the second step. The HOI is produced in the second step. However, it consumes in the third step. Thus, both OI^- ions HOI are not present in the overall reactants and products. So, these are reaction intermediate

- iii) What is the rate-determining step?

Slow step is always the rate determining step. Hence step 1 is the rate determining step.

- iv) Write the rate equation for the reaction, expressing it in terms of the reactants in the overall reaction.

Since, step 1 is the rate determining step. So, the rate equation can be derived from this step

Hence, $\text{Rate} = k [\text{H}_2\text{O}_2][\text{I}^-]$.

The reaction is first order in H_2O_2 and first order in I^- . Hence, it is 2nd order overall.

NUMERICAL PROBLEMS

- Q.6 (a) Calculate the reaction rate if the concentration of A is 0.5 M, the concentration of B is 0.2 and the rate constant k is $4.0 \text{ M}^{-2} \text{ s}^{-1}$. Given the rate law for a reaction: $\text{Rate} = k[\text{A}][\text{B}]^2$.

Solution

$$[\text{A}] = 0.5 \text{ M}$$

$$[\text{B}] = 0.2 \text{ M}$$

$$k = 4.0 \text{ M}^{-2} \text{ s}^{-1}$$

So, rate is given as

$$\begin{aligned} \text{Rate} &= k[\text{A}][\text{B}]^2 \\ &= 4.0 \text{ M}^{-2} \text{ s}^{-1} \times 0.5 \text{ M} \times [0.2 \text{ M}]^2 \\ &= 0.08 \text{ M s}^{-1} \end{aligned}$$

- Q.7 A first order reaction is found to have a rate constant, $k = 5.5 \times 10^{-14} \text{ s}^{-1}$. Find the half-life of the reaction.

Temp. (K)	Rate constant ($\text{cm}^3 \text{ mol}^{-1} \text{ s}^{-1}$)(K)
500	6.814×10^4
550	2.64×10^{-2}
600	0.56×10^0
650	7.31×10^0
700	66.67×10^0

Solution

$$k = 5.5 \times 10^{-14} \text{ s}^{-1}$$

So, half life time for a first order reaction is

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

$$t_{1/2} = \frac{0.693}{5.5 \times 10^{-14}} = 1.26 \times 10^{13} \text{ seconds}$$

or $t_{1/2} = \frac{1.26 \times 10^{13}}{60 \times 60 \times 24 \times 365} = 3.995 \times 10^5 \text{ years}$

Q.8 Three experiments that have identical conditions were performed to measure the initial rate of the reaction. $2\text{HI}_{(g)} \rightarrow \text{H}_{2(g)} + \text{I}_{2(g)}$

Experiment	[HI] (M)	Rate (M/s)
1	0.015	1.1×10^{-3}
2	0.030	4.4×10^{-3}
3	0.045	9.9×10^{-3}

Write the rate law for the reaction. Find the value and units of the specific rate constant, k.

Solution

- Comparing experiment 1 and 2, the ratio between concentration and the ratio between the rates is

$$\frac{(\text{HI})_2}{(\text{HI})_1} = \frac{0.030}{0.015} = 2 \quad \text{and} \quad \frac{(\text{Rate})_2}{(\text{Rate})_1} = \frac{4.4 \times 10^{-3}}{1.1 \times 10^{-3}} = 4$$

So, when the concentration of HI doubles, the rate of reaction increases four times. i.e. $2^2 = 4$

- Comparing experiment 1 and 3, the ratio between concentration and the ratio between the rates is

$$\frac{(\text{HI})_3}{(\text{HI})_1} = \frac{0.045}{0.015} = 3 \quad \text{and} \quad \frac{(\text{Rate})_3}{(\text{Rate})_1} = \frac{9.9 \times 10^{-3}}{1.1 \times 10^{-3}} = 9$$

So, when the concentration of HI increase three times, the rate of reaction increases nine times. i.e. $3^2 = 9$

Hence, rate of reaction is directly proportional to the second power of concentration of HI.

i.e., $\text{Rate} = k [\text{HI}]^2$ It is a second order reaction

Determination of rate constant

The rate equation is

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

or $k = \frac{\text{Rate}}{[\text{HI}]^2}$

From experiment 1 we have

$$k = \frac{1.1 \times 10^{-3} \text{ Ms}^{-1}}{[0.015 \text{ M}]^2}$$

$$k = 4.89 \text{ M}^{-1} \text{ s}^{-1}$$

Tests MCQs

- The rate equation determined experimentally for this reaction: $(\text{CH}_3)_3\text{CBr} + \text{H}_2\text{O} \rightarrow (\text{CH}_3)_3\text{COH} + \text{HBr}$ is $\text{Rate} = k[(\text{CH}_3)_3\text{CBr}]$. Hence it is which of the following? **MCAT 2009****
(A) Fractional Order (B) Pseudo First Order (C) First Order (D) Second Order
- The reaction rate in forward direction decreases with the passage of time because **MCAT 2012****
(A) Concentration of reactants decrease (C) The order of reaction changes
(B) Concentration of product decreases (D) Temperature of the system changes
- When the change in concentration is $6 \times 10^{-4} \text{ mol dm}^{-3}$ and time for that change is 10 seconds, the rate of reaction will be **MCAT 2015****
(A) $6 \times 10^{-3} \text{ mol dm}^{-3} \text{ sec}^{-1}$ (C) $6 \times 10^{-2} \text{ mol dm}^{-3} \text{ sec}^{-1}$ (B) $6 \times 10^{-4} \text{ mol dm}^{-3} \text{ sec}^{-1}$ (D) $6 \times 10^{-5} \text{ mol dm}^{-3} \text{ sec}^{-1}$
- The unit of the rate constant is the same as that of the rate of reaction in **MDCAT 2020****
(A) First order reaction (B) Second order reaction (C) Zero order reaction (D) Third order reaction.
- The rate of reaction involving ions can be studied by _____ method **MCAT 2010****
(A) Dilatometric (B) Refractometric (C) Optical rotation (D) Electrical conductivity
- In zero order reactions, the rate is independent of: **MCAT 2013****
(A) Concentration of the product (C) Temperature of the reaction
(B) Concentration of the reactant (D) Surface area of the product
- For the reaction $2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2$, the rate equation for the forward reaction is **MCAT 2014****
(A) $\text{Rate} = k[\text{NO}][\text{O}_2]$ (B) $\text{Rate} = k[\text{NO}]^2[\text{O}_2]$ (C) $\text{Rate} = k[\text{NO}_2]^2$ (D) $\text{Rate} = k[\text{NO}_2]$
- $2\text{A} + \text{B} \rightarrow \text{Product}$. If the reactant 'B' is in excess, the order of reaction with respect to 'A' in given rate law, $\text{Rate} = k[\text{A}]^2[\text{B}]$ is: **MCAT 2016****
(A) 2nd order reaction (B) 1st order reaction (C) Pseudo 1st order reaction (D) 3rd order reaction
- Unit of k in first order reaction is **MDCAT 2017****
(A) s^{-1} (B) moles dm^{-3} (C) $\text{moles dm}^{-3} \text{ s}^{-1}$ (D) $\text{mol}^{-2} \text{ dm}^{-3}$
- Rate of first order reaction depends on _____ **MDCAT 2017****
(A) concentration of one reactant (B) concentrations of two reactants
(C) concentrations of three reactants (D) independent of the initial concentration
- For which of the following order of reaction, rate of reaction is inversely proportional to the conc. **MDCAT 2020****
(A) 1st order reaction (B) 2nd order reaction (C) negative order reaction (D) zero order reaction
- For the reaction, $\text{A}_{(g)} \rightarrow \text{products}$, when the conc. of A doubles, the rate of reaction increases four folds, which means it is **MDCAT 2020****
(A) negative order reaction (B) zero order reaction (C) first order reaction (D) second order reaction
- The study of rate of chemical reactions and the factors that affect the rates of reactions is called **MDCAT 2020****
(A) Thermodynamics (B) Stoichiometry (C) Electrochemistry (D) Chemical kinetics
- When concentration-time graph of a reactant indicates a constant half-life, then the order with respect to that reactant is **MDCAT 2018****
(A) Zero order (B) Half order (C) Second order (D) First order
- Glucose is converted into ethanol by the enzyme _____ present in yeast: **MCAT 2010****
(A) Urease (B) Invertase (C) Sucrase (D) Zymase
- It is experimentally found that a catalyst is used to: **MCAT 2011****
(A) Lower the activation energy (C) Lower the pH
(B) Increase the activation energy (D) Decrease the temp of the reaction
- Catalyst helps in a reaction by **ECAT 2017****

- (A) Increasing the rate of reaction (B) Lowering the activation energy barrier
 (C) Increasing the activation energy barrier (D) Both options A and B are correct
- 18. Role of a catalyst in a chemical reaction is** **MDCAT 2018**
 (A) decrease yield of reaction (B) decrease rate of reaction (C) increase yield of product (D) increase rate of a process
- 19. If the energy of activation of a chemical reaction is very low, the rate of that reaction is high because** **MDCAT 2019**
 (A) concentration of the reactants becomes irrelevant (B) reaction proceeds without any transition state
 (C) number of effective collisions increases (D) molecules of the reactants move slowly
- 20. The rate of reaction between two specific time intervals is called?** **MDCAT 2023**
 (A) Instantaneous rate of reaction (B) Rate of reaction (C) Average rate of reaction (D) Initial rate
- 21. How will be the rate of reaction, if the slope of the curve is greater near the start of the reaction?** **MDCAT 2023**
 (A) Constant (B) Equilibrium (C) Greater (D) Lesser
- 22. If the rate does not change with concentration then it is?** **MDCAT 2024**
 (A) 3rd order (B) 2nd order (C) 1st order (D) zero order

Answers

Q#	Ans														
1	B	2	A	3	D	4	C	5	D	6	B	7	B	8	A
9	A	10	A	11	C	12	D	13	D	14	D	15	D	16	A
17	D	18	D	19	C	20	C	21	C	22	D				

Test Your Skills

OBJECTIVE: Time: 10 Minutes: Marks: 08

Q1. Choose the correct answer and encircle it.

- The unit of rate constant 'k' for a first order reaction are?
(A) s^{-1} (B) $\text{mol dm}^{-1}\text{s}$ (C) $\text{mol dm}^{-3}\text{s}^{-1}$ (D) $\text{mol}^{-1}\text{dm}^3\text{s}$
- The rate of reaction between two specific time intervals is called.
(A) Rate of a reaction (B) Average rate of a reaction.
(C) Instantaneous rate of reaction. (D) None
- The Unit of rate of reaction is
(A) mol dm^{-3} (B) mol kg^{-1} (C) grams dm^{-3} (D) $\text{mol dm}^{-3} \text{sec}^{-1}$
- If the rate equation of a reaction $2A + B \rightarrow \text{Products}$ is. $\text{Rate} = k[A]^2[B]$, and A is present in large excess, then order of reaction is
(A) 1 (B) 2 (C) 3 (D) None
- If the rate of reaction becomes four time when the conc. of reactant A is doubled, keeping concentrations of other reactants constant. Then the order of reaction w.r.t. A will be
(A) 4 (B) 2 (C) 3 (D) 1
- In zero order reaction, the rate is independent of
(A) Temperature of reaction (B) Concentration of reactants.
(C) Concentration of products (D) None of these.
- The half life period of is ${}^{238}_{92}\text{U}$ 710 million years. 100 kg of sample of ${}^{238}_{92}\text{U}$ will reduce to 25 kg in
(A) 710 million years (B) 1420 million years (C) 2130 million years (D) 2840 million years
- The rate of reaction for a multistep reaction is controlled by
(A) fast step (B) slow step (C) first step (D) last step

Fill in the correct option					Write Correct option here
1.	(A)	(B)	(C)	(D)	
2.	(A)	(B)	(C)	(D)	
3.	(A)	(B)	(C)	(D)	
4.	(A)	(B)	(C)	(D)	
5.	(A)	(B)	(C)	(D)	
6.	(A)	(B)	(C)	(D)	
7.	(A)	(B)	(C)	(D)	
8.	(A)	(B)	(C)	(D)	

SUBJECTIVE: Time: 60 minutes Marks: 32

Section - I

Q2. Answer the following short questions.

(2 × 12 = 24)

- Differentiate between average and instantaneous rate of reaction.
- What is a second order reaction?
- What is reaction mechanism?
- A catalyst is specific in its action. Explain
- What is half-life period? Give one example.
- How surface area affects the rate of a chemical reaction?
- Define (a) activation energy (b) specific rate constant
- The rate of a chemical reaction is an ever changing parameter under the given conditions. Justify
- The sum of the co-efficients of a balanced chemical equation is not necessarily important to give the order of reaction.
- What is rate determining step? Give an example.
- How does a catalyst work?
- Name the factors which affect the rate of reaction.

Section - II

(8 × 1 = 08)

- What is catalysis? Write down its types and explain any one of them. (04)
- A first order reaction is found to have a rate constant, $k = 5.5 \times 10^{-14} \text{ s}^{-1}$. Find the half-life of the reaction (04)

ANSWERS TO MCQs: TEST YOUR SKILLS-II

Q#	Ans														
1	A	2	B	3	D	4	A	5	B	6	B	7	B	8	B

