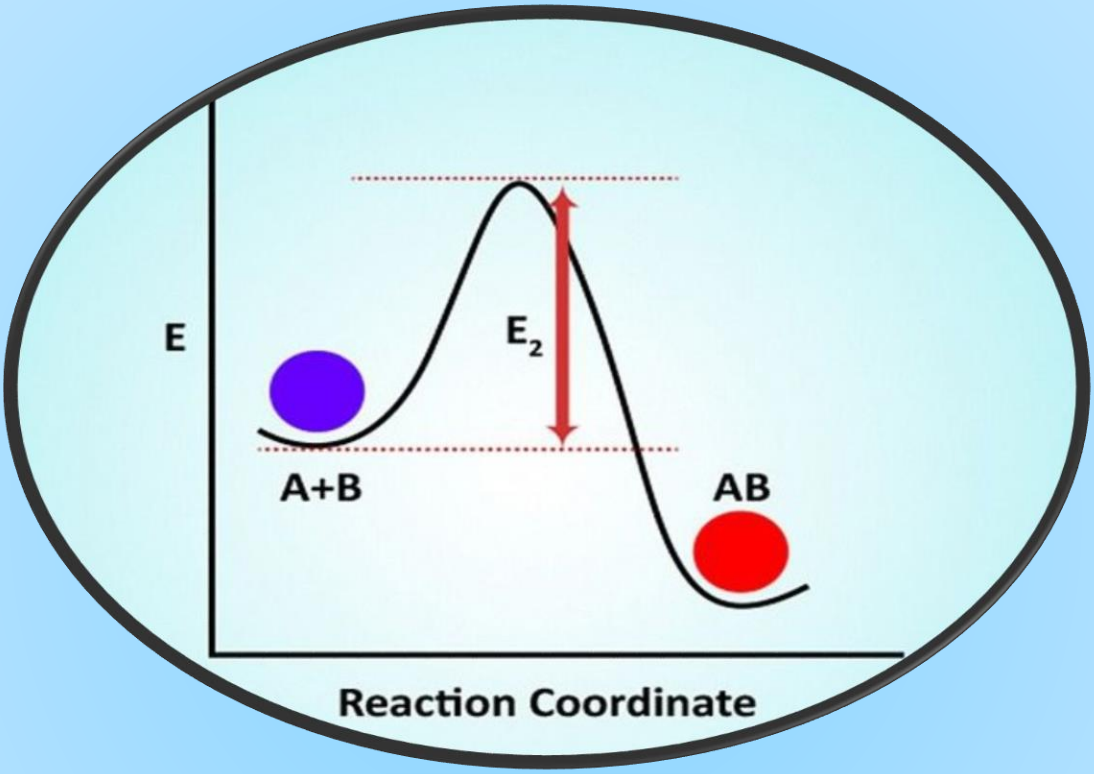
4. CHEMICAL KINETICS



Chemistry Smart Booklet

Theory + NCERT MCQs + Topic Wise Practice MCQs + NEET PYQs

CHEMICAL KINETICS

RATE OF RXN

for : $aA + bB \longrightarrow xX + yY$

$$rate = \frac{-1}{a}\frac{d[A]}{dt} = \frac{-1}{b}\frac{d[B]}{dt} = \frac{1}{x}\frac{d[X]}{dt} = \frac{1}{y}\frac{d[Y]}{dt}$$

INSTANTANEOUS RATE

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

AVERAGE RATE

$$rate = \frac{-\Delta R}{\Delta t} = \frac{-([R_2]-[R_1])}{t_2 - t_1}$$

Reaction	Unit of Rate Constant (K)
Zero Order	mol $L^{-1}S^{-1}$
First Order	S^{-1}
Second Order	$mol^{-1} LS^{-1}$

ORDER OF REACTION

Sum of powers of concentration of reactants in the rate law.

Order of a reaction may be whole number or a fraction.

$aA + bB \longrightarrow$ Product

rate = $k[A]^x [B]^y$

order = x + y

$$aA + bB \rightleftharpoons cC + dD$$

MOLECULARITY

- Number of reacting species taking party in an elementary reaction
- It cannot be zero or a fraction.

Molecularity = a + b

RATE CONSTANT

rate = $K[A]^x[B]^y$

RATE EXPRESSION

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

RATE CONSTANT UNIT FOR NTH ORDER REACTION

$K = mol^{1-n}\ Litre^{n-1}\ sec^{-1}$

COLLISION THEORY

A chemical reaction occurs when molecules collide with sufficient Energy.

EFFECTIVE COLLISION

Molecular collide with sufficient kinetic energy and proper orientation

COLLISION FREQUENCY

Number of collisions per unit volume of reaction mixture.

FOR BIMOLECULAR REACTION

rate = $PZ_{AB}e^{-Ea/RT}$

P → steric factor

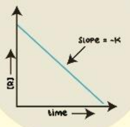
Z_{AB} → collisio frequency

INTEGRATED RATE EQUATION

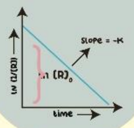
Reaction	Differential rate law	Integrated rate law	Half life
Zero Order	$\frac{d[R]}{dt} = -K$	$Kt = [R]_0 - [R]$	$t_{1/2} = \frac{[R]_0}{2K}$
First Order	$\frac{d[R]}{dt} = -K[R]$	$Kt = \ln\left(\frac{[R]_0}{[R]}\right)$	$t_{1/2} = \frac{0.693}{K}$

Radioactive decay are first order reaction.

ZERO ORDER



FIRST ORDER



Half life of nth order reaction

$t_{1/2} \propto [R_0]^{1-N}$

PSEUDO FIRST ORDER REACTION

- are not truely first order reaction but in certain conditions behaves like those.
- hydrolysis of ethyl acetate in acidic medium

FACTORS INFLUENCING RATE

CONCENTRATION

Higher the concentration of reactants. faster is the rate reaction.

TEMPERATURE

Rate of reaction increases with increase in temperature.

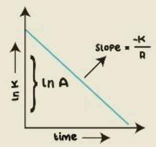
SURFACE AREA

Greater is the surface area. faster is the reaction rate.

PRESENCE OF CATALYST

Rate of reaction increases in presence of a catalyst.

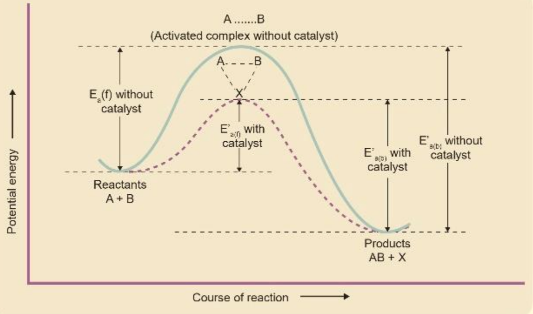
ARRHENIUS EQUATION



$$\log\left(\frac{K_1}{K_2}\right) = \frac{E_a}{2.303}\left[\frac{T_1 - T_2}{T_1T_2}\right]$$

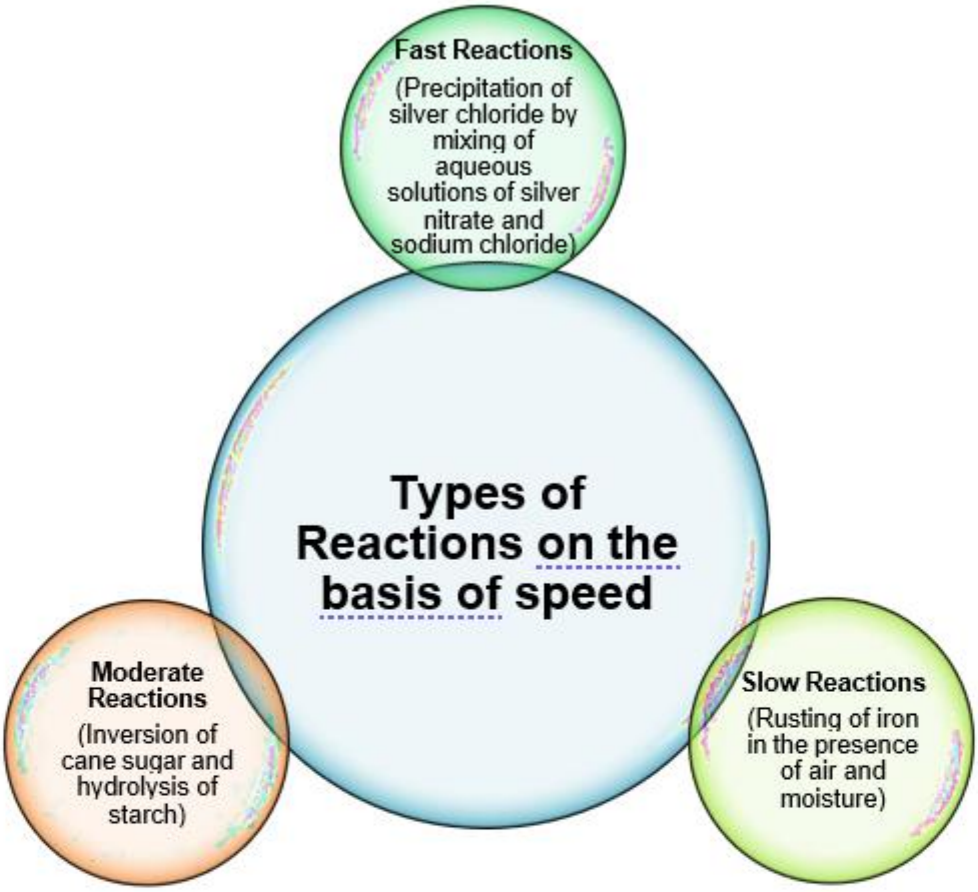
$$K = Ae^{-Ea/RT}$$

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



CHEMICAL KINETICS

Rate of Chemical Reaction



Rate of Chemical Reaction

The **rate of reaction** is the change in concentration of a reactant or productin unit time.

- Alternatively, the rate of reaction can also be expressed as

The rate of decrease in concentration of any one of the reactants.

The rate of increase in concentration of any one of the products.

- Consider a hypothetical reaction, assuming that the volume of the system remains constant.

$$R \rightarrow P$$

One mole of the reactant R produces one mole of the product P.

- If $[R]_1$ and $[P]_1$ are the concentrations of R and P at time t_1 and $[R]_2$ and $[P]_2$ are theirconcentrations at time t_2, then

$$\Delta t = t_2 - t_1$$
$$\Delta[R] = [R]_2 - [R]_1$$
$$\Delta[P] = [P]_2 - [P]_1$$

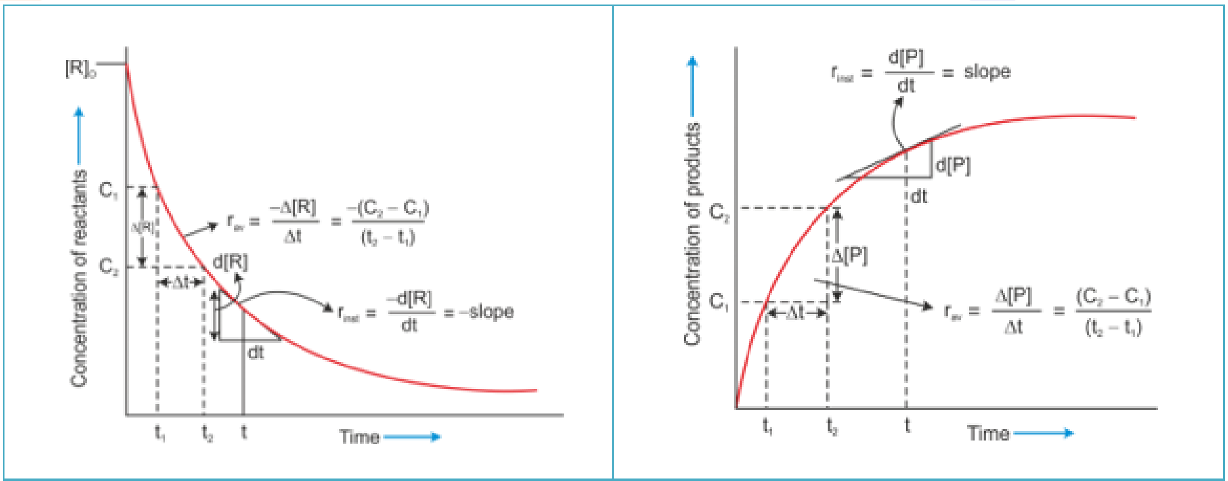
The square brackets in the above expressions are used to express molar concentration.

$$\text{Rate of disappearance of R} = \frac{\text{Decrease in concentration of R}}{\text{Time taken}} = -\frac{\Delta[R]}{\Delta t} \quad \text{\_\_\_\_\_(1)}$$

- $\Delta[R]$ is a negative quantity because the concentration of reactants is decreasing.

$$\text{Rate of appearance of P} = \frac{\text{Increase in concentration of P}}{\text{Time taken}} = +\frac{\Delta[P]}{\Delta t} \quad \text{\_\_\_\_\_(2)}$$

- Equations 1 and 2 represent the average rate of a reaction, r_{av}.
  This average rate depends on the change in concentration of reactants or products and the timetaken for that change to occur.



Instantaneous and average rate of a reaction

Units of Rate of a Reaction

- From Equations 1 and 2, it is clear that the units of rate are concentration time^{-1}.

- For example, if concentration is in mol L^{-1} and time is in seconds, then the units are mol $L^{-1}s^{-1}$.
- In gaseous reactions, the concentration of gases is expressed in terms of their partial pressures;hence, the units of the rate equation will be atm s^{-1}.

Instantaneous Rate of Reaction

- Consider the hydrolysis of butyl chloride (C_4H_9Cl).

$$C_4H_9Cl + H_2O \rightarrow C_4H_9OH + HCl$$

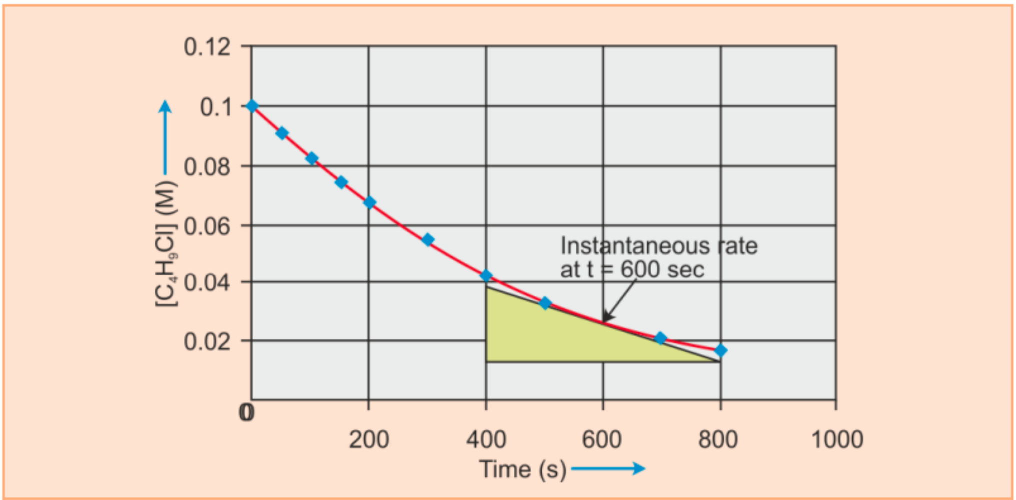
- We have provided the concentrations over different intervals of time below.

Time (s^{-1})	0	50	100	150	200	300	400	700	800
Concentration (mol L^{-1})	0.100	0.0905	0.0820	0.0741	0.0671	0.0549	0.0439	0.0210	0.017

- We can determine the difference in concentration over different intervals of time, and thus, wedetermine the average rate by dividing Δ[R] by Δt.
- It can be seen from experimental data that the average rate falls from 1.90×10^{-4} mol $L^{-1}s^{-1}$ to 0.4×10^{-4} mol $L^{-1}s^{-1}$.
- However, the average rate cannot be used to predict the rate of reaction at a particular instant as itwould be constant for the time interval for which it is calculated.
- Hence, to express the rate at a particular moment of time, we determine the instantaneous rate.
- It is obtained when we consider the average rate at the smallest time interval, say dt, when Δt approaches zero.

  Therefore, for an infinitesimally small dt, the instantaneous rate is given by

$$r_{inst} = -\frac{d[R]}{dt} = \frac{d[P]}{dt}$$



Instantaneous rate of hydrolysis of butyl chloride(C_4H_9Cl)

- By drawing the tangent at time t on either of the curves for the concentration of R versus time t orconcentration of P versus time t and calculating the slope of the curve, we can determine the instantaneous rate of reaction.
- Hence, in this example, r_{inst} at 600 s is calculated by plotting the graph of the concentration of butylchloride as against time t.
- A tangent is drawn on the curve at a point t = 600 s.

$$\therefore r_{inst} \text{ at } 600s = \left[\frac{0.0165-0.037}{(800-400)}\right] molL^{-1} = 5.12\times10^{-5} molL^{-1}s^{-1}$$

$$\begin{aligned} \text{At } t&=250s & r_{inst} &= 1.22\times10^{-4} molL^{-1}s^{-1} \\ t&=350s & r_{inst} &= 1.0\times10^{-4} molL^{-1}s^{-1} \\ t&=450s & r_{inst} &= 6.4\times10^{-5} molL^{-1}s^{-1} \end{aligned}$$

- Now consider a reaction,

$$Hg(l) + Cl_2(g) \rightarrow HgCl_2(s)$$

Here, the stoichiometric coefficients of the reactants and products are the same; hence, the rate of reaction is given as

$$\text{Rate of reaction} = -\frac{\Delta[Hg]}{\Delta t} = -\frac{\Delta[Cl_2]}{\Delta t} = \frac{\Delta[HgCl_2]}{\Delta t}$$

Therefore, we can say that from the above equation that the rate of disappearance of any of the reactants is the same as the rate of appearance of the products.

- Consider another reaction,

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

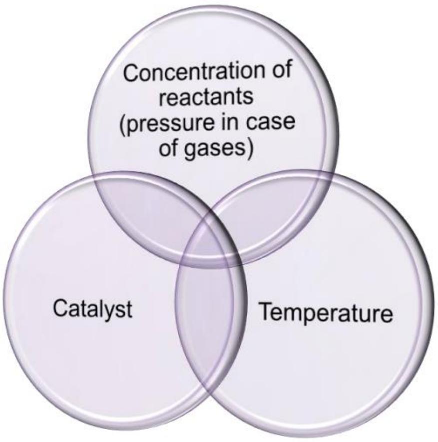
In this reaction, two moles of HI decompose to produce one mole each of H2 and I2, i.e. the stoichiometric coefficients of reactants or products are not equal to one; hence, we need to divide the rate of disappearance of any of the reactants or the rate of appearance of products by their respective stoichiometric coefficients.
Because the rate of consumption of HI is twice the rate of formation of H2 or I2, to make them equal, the term Δ[HI] is divided by 2.

The rate of this reaction is given by

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

- For a gaseous reaction at constant temperature, concentration is directly proportional to the partial pressure of a species, and hence, the rate can be expressed as the rate of change in partial pressure of the reactant or the product.

Instantaneous Rate of Reaction



NCERT LINE BY LINE QUESTIONS

(1.) **Assertion:** Catalyst changes Gibbs free energy of system.
Reason: Catalyst changes pre-exponential factor of a chemical reaction. **[Page: 115]**

(a.) Both A and R are true and R is correct explanation of A.

(b.) Both A and R are true and R is not correct explanation of A.

(c.) A is true but R is false.

(d.) Both A and R are false.

(2.) The reaction, $X \rightarrow$ Products follows first order kinetics. In 40 minutes, the concentration of X changes from 0.1 M to 0.025 M then the rate of reaction when concentration of X is 0.01 M is:

(a.) $1.73 \times 10^{-4} \mathrm{M}/\ \mathrm{min}$

(b.) $3.47 \times 10^{-5} \mathrm{M}/\ \mathrm{min}$

(c.) $3.47 \times 10^{-4} \mathrm{M}/\ \mathrm{min}$

(d.) $1.73 \times 10^{-5} \mathrm{M}/\ \mathrm{min}$

(3.) The time for half-life period of a certain reaction $A \rightarrow$ Products is 1 hour. When the initial concentration of the reactant 'A' is $2.0\ \mathrm{mol\ L^{-1}}$, how much time does it take for its concentration to come from 0.50 to$0.25\ \mathrm{mol\ L^{-1}}$ if it is a zero order reaction)

(a.) 1 h

(b.) 4 h

(c.) 0.5 h

(d.) 0.25 h

(4.) For the reaction $A + 2B \rightarrow C$ the reaction rate is doubled if the concentration of A is doubled. The rate is increased by four times when concentrations of both A and B are increased by four times. The order of the reaction is: **[Page: 103]**

(a.) 1

(b.) 2

(c.) 3

(d.) 0

(5.) When initial concentration of a reactant is doubled, the order of the reaction is:
[NEET-2015, Page: 107, 1O8]

(a.) First

(b.) Second

(c.) More than zero but less than first

(d.) Zero

(6.) In a first order reaction 80% of the reactant at an instant was reduced to 8% in 4606 seconds. The rate constant of the reaction is **[Page: 108]**

(a.) $2.303 \times 10^{-4} s^{-1}$

(b.) $4.606 \times 10^{-3} s^{-1}$

(c.) $5.000 \times 10^{-4} s^{-1}$

(d.) $5.000 \times 10^{-3} s^{-1}$

(7.) On the reaction $3A \rightarrow 2B$, rate of reaction $+\frac{d[B]}{\Delta t}$ is equal to: **[Page: 98]**

(a.) $+2\frac{d[A]}{dt}$

(b.) $-\frac{3}{2}\frac{d[A]}{dt}$

(c.) $-\frac{1}{3}\frac{d[A]}{dt}$

(d.) $-\frac{2}{3}\frac{d[A]}{dt}$

(8.) A chemical reaction proceeds following the formula $K = Pze^{-Ea/RT}$. Which of the following processes will increase the rate of reaction . **[Page: 116]**

(a.) Increasing of Ea

(b.) Lowering of P

(c.) Lowering of Z (d.) Independent of all the above factors.

(9.) A first order reaction has a specific reaction rate of $10^{-2}s^{-1}$. How much time will it take for $20g$ of the reactant to reduce to $5g$? **[NEET-2017, Page: 106]**

(a.) 693.0 s (b.) 238.6 s

(c.) 138.6 s (d.) 346.5 s

(10.) Consider the reaction $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$ **[Page: 10l, 103]**

The rate law for this is Rate $= K[N_2O_5]$.

Which of the following statements is true regarding the above reaction?

(a.) Its order is 1 and molecularity 1 (b.) Its order is 1 and molecularity 2

(c.) Its order is 2 and molecularity 2 (d.) Its order is 2 and molecularity 1

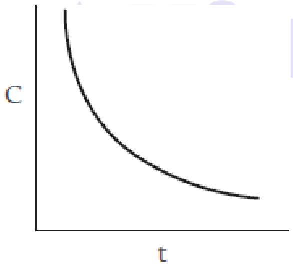
(11.) For a given reaction of first order, it takes 15 minute for the concentration to drop from $0.8\ \text{M litre}^{-1}$ to $0.4\ \text{M litre}^{-1}$. The time required for the concentration to drop from $0.1\ \text{M litre}^{-1}$ to $0.025\ \text{M litre}^{-1}$ will be:

(a.) 60 minute (b.) 15 minute

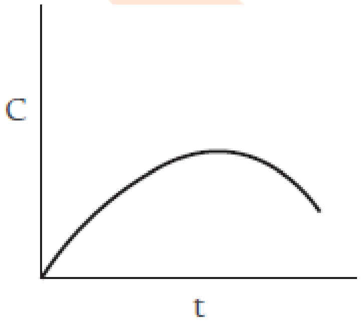
(c.) 7.5 minute (d.) 30 minute

(12.) The plot of concentration of the reactant versus time for a zero-order reaction is given by **[Page: 107]**

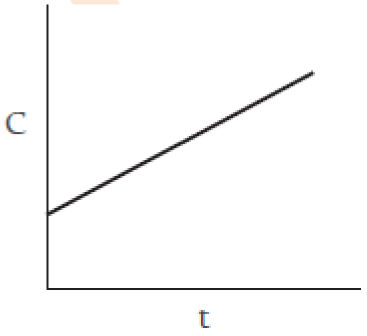
(a.)



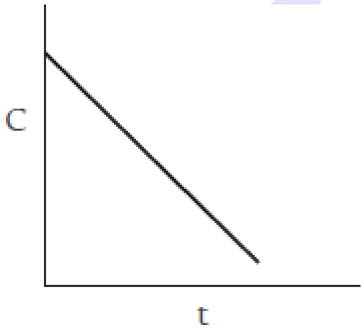
(b.)



(c.)



(d.)



(13.) The activation energy of a reaction at a given temperature is found to be 2.303 RT J mol^{-1}.The ratio of rate constant to the Arrhenius factor is: **[Page: 113, 114]**

(a.) 0.01 (b.) 0.02

(c.) 0.001 (d.) 0.1

(14.) If the rate constant for a first order reaction is K, the time 't' required for the completion of 99% of the reaction is given by **[NEET-2O19, Page 107, l08]**

(a.) $t = 0.693/K$ (b.) $t = 6.909/K$

(c.) $t = 4.606/K$ (d.) $t = 2.303/K$

(15.) **Assertion:** The kinetics of the reaction $m\text{A}+n\text{B}+p\text{C}\rightarrow m'\text{X}+n'\text{Y}+p'\text{Z}$ obeys the rate expression as $\frac{\text{d}X}{\text{dt}}=K[\text{A}]^m[\text{B}]^n$.

Reason: The rate of reaction does not depend upon the concentration of C. **[Page: 106]**

(a.) Both A and R are true and R is correct explanation of A.

(b.) Both A and R are true and R is not correct explanation of A.

(c.) A is true but R is false.

(d.) Both A and R are false.

(16.) For the non-stoichiometric reaction $2\text{A}+\text{B}\rightarrow\text{C}+\text{D}$, the following kinetic data were obtained in three separate experiments, al at 298 K.

Initial concentration (A)	**Initial concentration (B)**	**Initial rate of formation C $\left(\textbf{mol L}-1\textbf{s}-1\right)$**
$0.1M$	$0.1M$	1.2×10^3
$0.1M$	$0.2M$	1.2×10^3
$0.2M$	$0.1M$	2.4×10^3

The rate law for the formation of C, is:

(a.) $\frac{\text{d}[\text{C}]}{\text{dt}}=K[\text{A}][\text{B}]$

(b.) $\frac{\text{d}[\text{C}]}{\text{dt}}=K[\text{A}]^2[\text{B}]$

(c.) $\frac{\text{d}[\text{C}]}{\text{dt}}=K[\text{A}][\text{B}]^2$

(d.) $\frac{\text{d}[\text{C}]}{\text{dt}}=K[\text{A}]$

(17.) The energy of activation for forward and backward change for an endothermic reaction $\text{X}\rightarrow\text{Y}$ are E_f and E_b respectively. Which is correct

(a.) $E_b<E_f$

(b.) $E_b>E_f$

(c.) $E_b=E_f$

(d.) No reaction between them

(18.) According to collision theory of reaction rates, increase in the temperature of a reaction will increase the rate of the reaction because of. **[Page: 115]**

(a.) Increase in the velocity of the reacting molecules

(b.) Increase in the number of collisions

(c.) Increase in the number of molecules having the activation energy

(d.) None of these

(19.) The decomposition of phosphine $\left(\text{PH}_3\right)$ on tungsten at low pressure is first order reaction. It is because the **[NEET-Phase-II-(2016)]**

(a.) Rate is proportional to surface coverage

(b.) Rate is inversely proportional to surface coverage

(c.) Rate is independent of the surface coverage

(d.) Rate of decomposition is very slow

(20.) The value of rate constant of a pseudo-first order reaction **[QR Code, NCERT Exemplar, Page: 10l, 103]**

(a.) Depends on the concentration of reactants present in small amount

(b.) Depends on the concentration of reactants present in excess (c.) is independent of the concentration of reactants

(d.) Depends only on temperature

(21.) The rate law for a reaction between the substances A and B is given by rate $= K[\mathrm{A}]^n[\mathrm{B}]^m$. On doubling the concentration of A and halving the concentration of B, the ratio of the new rate to the earlier rate of the reaction will be as:

(a.) $1/2^{m+n}$

(b.) $(m+n)$

(c.) $(n-m)$

(d.) $2^{(n-m)}$

(22.) For a first order reaction $\mathrm{A} \to \mathrm{P}$, the temperature (*T*) dependent rate constant (*k*) was found to follow the equation $\log \log k = -(2000)\frac{1}{T} + 6.0$ The pre-exponential factor *A* and the activation energy E_a, respectively, are:

(a.) $1.0\times10^6\,\mathrm{s}^{-1}and9.2\ \mathrm{kJ\ mol}^{-1}$

(b.) $6.0\mathrm{s}^{-1}$and$16.6\mathrm{kJ\ mol}^{-1}$

(c.) $1.0\times10^6\,\mathrm{s}^{-1}and16.6\ \mathrm{kJ\ mol}^{-1}$

(d.) $1.0\times10^6\,\mathrm{s}^{-1}and38.3\ \mathrm{kJ\ mol}^{-1}$

(23.) Energy of activation of forward reaction for an endothermic process is $90\,KJ$. If enthalpy change for the reaction is $50\,KJ$, then activation energy for backward reaction will be: **[Page: 110]**

(a.) $40\,KJ$

(b.) $50\,KJ$

(c.) $90\,KJ$

(d.) $140\,KJ$

(24.) $\mathrm{t}_{1/2}$ for a first order reaction is 14.26 min. Calculate the time when 5% of the reaction is left. **[Page: 109]**

(a.) 26 min

(b.) 42 min

(c.) 53 min

(d.) 62 min

(25.) For the reaction $\mathrm{A+B \to C+D}$, doubling the concentration of both the reactants increases the reaction rate by 8 times and doubling the concentration of only 'B' simply doubles the reaction rate. The rate law is given as:

(a.) $r = K[\mathrm{A}]^{1/2}[\mathrm{B}]^{1/2}$

(b.) $r = K[\mathrm{A}][\mathrm{B}]^2$

(c.) $\mathrm{r} = K[\mathrm{A}]^2[\mathrm{B}]$

(d.) $r = K[\mathrm{A}][\mathrm{B}]$

(26.) Activation energy ($\mathrm{E_a}$) and rate constants ($\mathrm{K_1}$ and $\mathrm{K_2}$) of chemical reactions at two different temperatures ($\mathrm{T_1}$ and $\mathrm{T_2}$) are related by: **[Page: 113]**

(a.) $\ln\frac{\mathrm{K_2}}{\mathrm{K_1}} = -\frac{\mathrm{E_a}}{\mathrm{T}}\left(\frac{1}{\mathrm{T_1}} - \frac{1}{\mathrm{T_2}}\right)$

(b.) $\ln\frac{\mathrm{K_2}}{\mathrm{K_1}} = -\frac{\mathrm{E_a}}{\mathrm{T}}\left(\frac{1}{\mathrm{T_2}} - \frac{1}{\mathrm{T_1}}\right)$

(c.) $\ln\frac{\mathrm{K_2}}{\mathrm{K_1}} = -\frac{\mathrm{E_a}}{\mathrm{T}}\left(\frac{1}{\mathrm{T_2}} + \frac{1}{\mathrm{T_1}}\right)$

(d.) $\ln\frac{\mathrm{K_2}}{\mathrm{K_1}} = \frac{\mathrm{E_a}}{\mathrm{T}}\left(\frac{1}{\mathrm{T_1}} - \frac{1}{\mathrm{T_2}}\right)$

(27.) Decomposition of $\mathrm{H_2O_2}$ follows a first order reaction fifty minutes the concentration of $\mathrm{H_2O_2}$ decreases from 0.5 to 0.125 M in one such decomposition. When concentration of $\mathrm{H_2O_2}$ reaches 0.05 M, the rate formation of $\mathrm{O_2}$ will be:

(a.) $6.93\times10^4\,\mathrm{mol\ min}^{-1}$

(b.) $2.66\ \mathrm{L\ min}^{-1}$at STP

(c.) $1.34\times10^2\,\mathrm{mol\ min}^{-1}$

(d.) $6.93\times10^2\,\mathrm{mol\ min}^{-1}$

(28.) Assertion: Rate of reaction doubles when concentration of reactant is doubled if it is a first order reaction.
Reason: Rate constant a1so doubles. **[Page: 10l, 103]**

(a.) Both A and R are true and R is correct explanation of A.
(b.) Both A and R are true and R is not correct explanation of A.
(c.) A is true but R is false.
(d.) Both A and R are false.

(29.) The order of reaction is:

(a.) 3
(b.) 6
(c.) miss text
(d.) miss text

(30.) For a first order reaction, (A) → products, the concentration of A changes from $0.1M$ to $0.025M$ in 40 minutes. The rate of reaction when the concentration of A is 0.01M, is:

(a.) $3.47\times10^{-4}\text{M}/\text{ min}$
(b.) $3.47\times10^{-5}\text{M}/\text{ min}$
(c.) $1.73\times10^{-4}\text{M}/\text{ min}$
(d.) $1.73\times10^{-5}\text{M}/\text{ min}$

(31.) Kinetics of the reaction $\text{A}(\text{g})\rightarrow 2\text{B}(\text{g})+\text{c}(\text{g})$ is followed by measuring the total pressure at different times. It is given that.
Initial pressure of $\text{A}=0.5\text{ atm}$
Total pressure of A after $2\text{h}=0.7\text{ atm}$
Rate constant of the reaction $=1\times10^{-3}\text{s}^{-1}$
What is the rate of reaction $\dfrac{-\text{d}[\text{A}]}{\text{dt}}$
when the total pressure is $0.7\,atm$?

(a.) $2.0\times10^{-4}\text{M s}^{-1}$
(b.) $4.0\times10^{-4}\text{M s}^{-1}$
(c.) $5.0\times10^{-4}\text{M s}^{-1}$
(d.) $7.0\times10^{-4}\text{M s}^{-1}$

(32.) $t_{1/4}$ can be taken as the time taken for concentration of reactant to drop to 3/4 of its initial value. If the rate constant for a first order reaction is K, then $t_{1/4}$ can be written as:

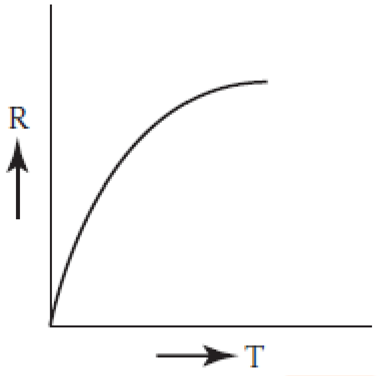
(a.) $0.10/K$
(b.) $0.29/K$
(c.) $0.69/K$
(d.) $0.75/K$

(33.) In a first order-reaction $\text{A}\rightarrow\text{B}$, if K is rate constant and initial concentration of the reactant 'A' is $0.5M$, then the half life is **[Page: 10.8]**

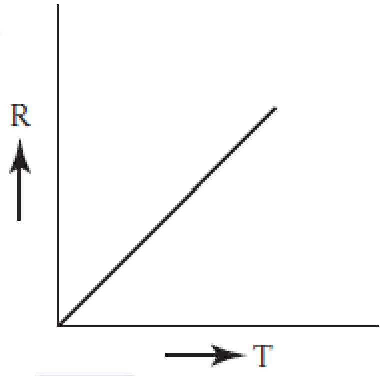
(a.) $\dfrac{\ln 2}{K}$
(b.) $\dfrac{\log 2}{K}$
(c.) $\dfrac{\log 2}{K\sqrt{0.5}}$
(d.) $\dfrac{0.693}{0.5K}$

(34.) Which curve corresponds to the temperature dependence of the rate R of a simple one step Reaction? **[Page: 113, 115]**

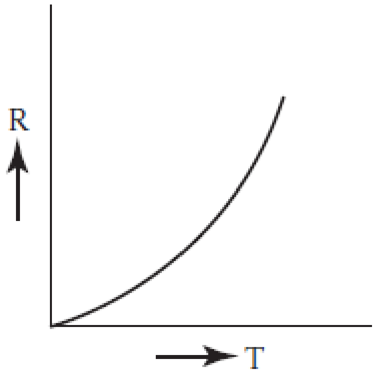
(a.)



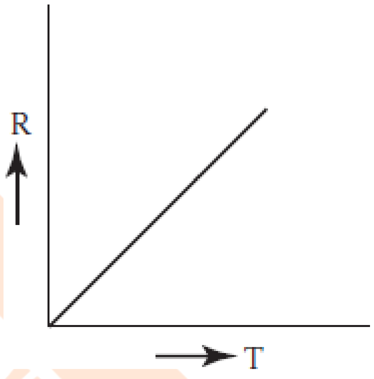
(b.)



(c.)



(d.)



(35.) The following mechanism has been proposed for the reaction of NO with $\mathrm{Br_2}$ to form $NOBr$:

$\mathrm{NO(g)+Br_2(g)} \rightleftharpoons \mathrm{NOBr_2(g)}$

$\mathrm{NOBr_2(g)+NO(g)\rightarrow 2NOBr}(g)$ (g)

If the second step is the rate determining step, the order of the reaction with respect to $NO(g)$ is:

(a.) 1
(b.) 0
(c.) 3
(d.) 2

(36.) Consider the chemical reaction, $\mathrm{N_2(g)+3H_2}(g)\rightarrow \mathrm{2NH_3(g)}$

The rate of this reaction can be expressed in terms of time derivative of concentration of $\mathrm{N_2(g), H_2(g)}$ and $\mathrm{NH_3(g)}$. Identify the correct relationship amongst the rate expression:

(a.) $Rate = -\mathrm{d[N_2]}/dt = -\frac{1}{3}\mathrm{d[H_2]/dt} = \frac{1}{2}\mathrm{d[NH_3]}/dt$

(b.) $Rate = -\mathrm{d[N_2]/dt} = -3\mathrm{d[H_2]}/dt = 2\mathrm{d[NH_3]/dt}$

(c.) $Rate = \mathrm{d[N_2]/dt} = \frac{1}{3}\mathrm{d[H_2]/dt} = \frac{1}{2}\mathrm{d[NH_3]/dt}$

(d.) $Rate = -\mathrm{d[N_3]/dt} = -\mathrm{d[H_2]/dt} = \mathrm{d[NH_3]/dt}$

(37.) Which of the following statement is not correct for the catalyst? **[QR Code, NCERT Exemplar, Page: 116]**

(a.) It catalyses the forward and backward reaction to the same extent.
(b.) It alters ΔG of the reaction.
(c.) It is a substance that does not change the equilibrium constant of a reaction.
(d.) It provides an alternate mechanism by reducing activation energy between reactants and products.

(38.) In an attempt to compare the half-lives of two radioactive elements A and B, a scientist set aside $400g$ of each. After 3 months, the scientist found $25g$ of ‘A’ and $200g$ of B. Which one of the following statements is true? **[HOTS]**

(a.) Half life of B is twice that of A
(b.) Half life of B is four times that of A
(c.) Half life of A is twice that of B
(d.) Half life of A is four times that of B

(39.) For a reaction $A+B \rightarrow C+2D$. Experimental results were collected for 3 trials and the data obtained are given below: **[Page: 106]**

Trial	[A], M	[B], M	Initial rate, $M\,s^{-1}$
1	0.40	0.20	5.5×10^4
2	0.80	0.20	5.5×10^4
3	0.40	0.40	2.2×10^{-3}

The correct rate law of the reaction is:

(a.) Rare $=K[A]^0[B]^2$
(b.) Rate $=K[A][B]^2$
(c.) Rate $=R[A][B]$
(d.) Rare $=K[A][B]^0$

(40.) In a reversible reaction, the enthalpy change and the activation energy is the forward direction are respectively $-x$ kJ mol^{-1}, and y kJ mol^{-1}. **[Page: 113, 115]**

Therefore, the energy of activation and in the backward direction is kJ mol^{-1} is:

(a.) $y-x$
(b.) $(x+y)$
(c.) $x-y$
(d.) $-x-y$

(41.) Which of the following expressions is correct for the rate of reaction given below? **[NCERT Exemplar, Page: 103]**

$$5Br^-_{(aq)} + BrO^-_{(3aq)} + 6H^+_{(aq)} \rightarrow 3Br_{(2aq)} + 3H_2O_{(I)}$$

(a.) $\frac{\Delta(Br^-)}{\Delta t} = 5\frac{\Delta(H^+)}{\Delta t}$

(b.) $\frac{\Delta(Br^-)}{\Delta t} = \frac{6}{5}\frac{\Delta(H^+)}{\Delta t}$

(c.) $\frac{\Delta(Br^-)}{\Delta t} = \frac{5}{6}\frac{\Delta(H^+)}{\Delta t}$

(d.) $\frac{\Delta(Br^-)}{\Delta t} = 6\frac{\Delta(H^+)}{\Delta t}$

(42.) The correct difference between first order reaction and second-reaction is that: **[NEET-2018, Page: 107, 108]**

(a.) A first-order reaction can be catalysed, a second order reaction cannot be catalysed
(b.) The rate of a first order reaction does not depend on reactant concentrations, the rate of a second order reaction does depend on reactant concentrations

(c.) The rare of a first order reaction does depend on reactant concentrations; the rate of a second order reaction does not depend on reactant-concentrations.

(d.) The half life of a first-order reaction does not depend on $[A]_0$,· the half life of a second order reaction does depend on $[A]_0$

(43.) For the reaction

$$Ag^+ + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]^+$$

the net rate of reaction is given by

$$\frac{dx}{dt} = 2\times10^7 [Ag^+][NH_3]^2 - 1\times10^{-2}[Ag(NH_3)_2]^+$$

Then which of the following statements is/are correct?

(a.) Rate constant for forward reaction $= 2\times10^7$

(b.) Rate constant for backward reaction $= 1\times10^{-2}$

(c.) Equilibrium constant of the reaction $= 2\times10^9$

(d.) All of the above

(44.) A first order reaction, which is 30% complete in 30 minutes has a half life period of. **[Page: 108, 109]**

(a.) 120.2 min

(b.) 102.2 min

(c.) 24.2 min

(d.) 58.2 min

(45.) For a first order reaction $A \rightarrow$ Product, the initial concentration of A is 0.1 M and after 40 minute it becomes 0.025 M. Calculate the rate of reaction at reactant concentration of 0.01 M:

(a.) $3.47\times10^4 \text{M min}^{-1}$

(b.) $3.47\times10^5 \text{M min}^{-1}$

(c.) $1.735\times10^6 \text{M min}^{-1}$

(d.) $1.735\times10^4 \text{M min}^{-1}$

(46.) Consider the following statement

(A) increase in concentration of reactant increases the rate of zero order reaction.

(B) rate constant *K* is equal to collision frequency A if $E_a = 0$.

(C) rate constant *K* is equal to collision frequency A if $E_a = \propto$.

(D) In *K* vs *T* is a straight line

(E) In *K* vs $1/T$ is a straight line

Correct statement are

(a.) **(a)** and **(d)**

(b.) **(b)** and (e)

(c.) **(c)** and **(d)**

(d.) **(b)** and **(c)**

(47.) In a first order reaction the concentration of reactant decreases from 800 mol/dm^3 to $50 mol/\text{dm}^3$ in 2×10^4sec. The rate constant of reaction in sec^{-1} is:

(a.) 2×10^4

(b.) 3.45×10^{-5}

(c.) 1.386×10^{-4}

(d.) 2×10^{-4}

(48.) For a chemical reaction, $mA \rightarrow xB$, the rate law is $r = K[A]^2$. If the concentration of A is doubled, the reaction rare will be: **[Page: 99]**

(a.) Increased by 8 times

(b.) Quadrupled

(c.) Doubled

(d.) Unchanged

(49.) For a reaction $X \to Y$, the graph of the product concentration X versus time (t) came out to be a straight line passing through the origin. Hence the graph of $-\frac{d[X]}{dt}$ and time would be: **[Page: 107]**

(a.) A hyperbola.

(b.) A straight line parallel to x-axis

(c.) Straight line with a negative slope and an intercept on y-axis

(d.) Straight line with a positive slope and an intercept on y-axis

(50.) **Assertion:** The overall order of the reaction is the sum of the exponents of all the reactants in the rate expression.

Reason: There are many higher order reactions. **[Page: 10l, 103]**

(a.) Both A and R are true and R is correct explanation of A.

(b.) Both A and R are true and R is not correct explanation of A.

(c.) A is true bur R is false.

(d.) Both A and R are false.

TOPIC WISE PRACTICE QUESTIONS

TOPIC 1: Rate of Reaction, Rate Laws and Rate Constant

1. Which of the following will react at the highest rate?
   1) 1 mole of A and 1 mole of B in a 1-L vessel 2) 2 mole of A and 2 mole of B in a 2-L vessel
   3) 3 mole of A and 3 mole of B in a 3-L vessel 4) All would react at the same rate

2. Burning of coal is represented as $C(s)+O_2(g)\to CO_2(g)$
   The rate of this reaction is increased by
   1) decrease in the concentration of oxygen. 2) powdering the lumps of coal.
   3) decreasing the temperature of coal. 4) providing inert atmosphere.

3. If n_A and n_B are the number of moles at any instant in the reaction : $2A(g)\to 3B(g)$ carried out in a vessel of VL, the rate of the reaction at that instant is given by
   1) $-\frac{1}{2}\frac{dn_A}{dt}=\frac{1}{3}\frac{dn_B}{dt}$ 2) $-\frac{1}{V}\frac{dn_A}{dt}=\frac{1}{V}\frac{dn_B}{dt}$ 3) $-\frac{1}{2V}\frac{dn_A}{dt}=\frac{1}{3V}\frac{dn_B}{dt}$ 4) $-\frac{1}{V}\frac{n_A}{t}=\frac{1}{V}\frac{n_B}{t}$

4. Rate of a reaction
   1) decreases with increase in temperature. 2) increases with increase in temperature.
   3) may increase or decrease with increase in temperature.
   4) does not depend on temperature.

5. $3A\to 2B$, rate of reaction $\frac{d[B]}{dt}$] is equal to
   1) $-\frac{3}{2}\frac{d[A]}{dt}$ 2) $-\frac{2}{3}\frac{d[A]}{dt}$ 3) $-\frac{1}{3}\frac{d[A]}{dt}$ 4) $+2\frac{d[A]}{dt}$

6. For the reaction $2A+B\to 3C+D$
   Which of the following does not express the reaction rate ?
   1) $-\frac{d[B]}{dt}$ 2) $\frac{d[D]}{dt}$ 3) $-\frac{1}{2}\frac{d[A]}{dt}$ 4) $-\frac{1}{3}\frac{d[C]}{dt}$

7. In the rate equation, when the conc. of reactants is unity then rate is equal to
   1) specific rate constant 2) average rate constant
   3) instantaneous rate constant 4) None of above

8. The rate of reaction between two specific time intervals is called
   1) instantaneous rate 2) average rate 3) specific rate 4) ordinary rate

9. $k_{34^0}/k_{35^0}<1$ then
   1) rate increases with the rise in temperature. 2) rate decreases with rise in temperature.
   3) rate does not change with rise in temperature. 4) none of these.

10. The rate of the reaction, $2NO + Cl_2 \rightarrow 2NOCl$ is given by the rate equation, rate = $k[NO]^2[Cl_2]$

The value of the rate constant can be increased by

1) increasing the concentration of NO. 2) increasing the temperature.
3) increasing the concentration of the Cl_2. 4) doing all of these.

TOPIC 2: Order of Reaction and Half Life Period

11. For the following homogeneous reaction, $A + B \xrightarrow{k} C$ the unit of rate constant is the unit of rate constant is

1) sec^{-1} 2) sec^{-1} mol L^{-1} 3) sec^{-1} mol^{-1} L 4) sec^{-1} mol^{-2} L^2

12. Which one of the following statements for the order of a reaction is incorrect ?

1) Order can be determined only experimentally.
2) Order is not influenced by stoichiometric coefficient of the reactants.
3) Order of reaction is sum of power to the concentration terms of reactants to express the rate of reaction.
4) Order of reaction is always whole number.

13. Units of rate constant of first and zero order reactions in terms of molarity M unit are respectively

1) sec^{-1}, $Msec^{-1}$ 2) sec^{-1}, M 3) $Msec^{-1}$, sec^{-1} 4) M, sec^{-1}.

14. Which one of the following reactions is a true first order reaction?

1) Alkaline hydrolysis of ethyl acetate
2) Acid catalyst hydrolysis of ethyl acetate
3) Decomposition of N_2O
4) Decomposition of gaseous ammonia on a hot platinum surface

15. The rate constant of a reaction is 0.0693 min^{-1}. Starting with 10 mol, the rate of the reaction after 10 min is

1) 0.0693 mol min^{-1} 2) 0.0693 × 2 mol min^{-1} 3) 0.0693 × 5 mol min^{-1} 4) 0.0693 × $(5)^2$ mol min^{-1}

16. The hydrolysis of ester in alkaline medium is a

1) first order reaction with molecularity 1 2) second order reaction with molecularity > 2
3) first order reaction with molecularity 2 4) second order reaction with molecularity 1

17. For the zero order reaction $A \rightarrow B + C$; initial concentration of A is 0.1 M. If [A] = 0.08 M after 10 minutes, then its half-life and completion time are respectively :

1) 10 min; 20 min 2) 2×10^{-3} min; 4×10^{-3} min
3) 25 min, 50 min 4) 250 min, 500 min

18. The half-life of a reaction is inversely proportional to the square of the initial concentration of the reactant. Then the order of the reaction is

1) 0 2) 1 3) 2 4) 3

19. Diazonium salt decomposes as

$C_6H_5N_2^+Cl^- \rightarrow C_6H_5Cl + N_2$ At 0°C, the evolution of N_2 becomes two times faster when the initial concentration of the salt is doubled. Therefore, it is

1) a first order reaction 2) a second order reaction
3) independent of the initial concentration of the salt 4) a zero order reaction

20. In the presence of an acid, the initial concentration of cane sugar was reduced from 0.20 to 0.10 M in 5 hours and from 0.2 to 0.05 M in 10 hours. The reaction is of

1) zero order 2) first order 3) second order 4) third order

21. For the reaction $3A(g) \xrightarrow{k} B(g) + C(g)$, k is 10^{-4} L/mol min.

If [A] = 0.5 M then the value of $-\dfrac{d[A]}{dt}\left(\text{in } Ms^{-1}\right)$ is

1) 7.5×10^{-5} 2) 3×10^{-4} 3) 2.5×10^{-5} 4) 1.25×10^{-6}

22. 99% of a first order reaction was completed in 32 minutes when 99.9% of the reaction will complete :

1) 50 min 2) 46 min 3) 48 min 4) 49 min

23. Which of the following represents the expression for 3/4th life of first order reaction

1) $\dfrac{k}{2.303}\log 4/3$ 2) $\dfrac{2.303}{k}\log 4/3$ 3) $\dfrac{2.303}{k}\log 4$ 4) $\dfrac{2.303}{k}\log 3$

24. In a 1st order reaction, reactant concentration C varies with time t as :
1) $1/C$ increases linearly with t 2) $\log C$ decreases linearly with t
3) C decreases with $1/t$ 4) $\log C$ decreases with $1/t$

25. A reaction proceeds by first order, 75% of this reaction was completed in 32 min. The time required for 50% completion is
1) 8 min 2) 16 min 3) 20 min 4) 24 min

26. A substance 'A' decomposes by a first order reaction starting initially with [A] = 2.00 M and after 200 min, [A] becomes 0.15 M. For this reaction $t1/2$ is
1) 53.72 min 2) 50.49 min 3) 48.45 min 4) 46.45 min

27. Half life of a first order reaction is 4s and the initial concentration of the reactant is 0.12 M. The concentration of the reactant left after 16 s is
1) 0.0075 M 2) 0.06 M 3) 0.03 M 4) 0.015 M

28. $3A \rightarrow B + C$, it would be a zero order reaction when
1) the rate of reaction is proportional to square of concentration of A.
2) the rate of reaction remains same at any concentration of A.
3) the rate remains unchanged at any concentration of B and C.
4) the rate of reaction doubles if concentration of B is increased to double.

29. The reaction $A \rightarrow B$ follows first order kinetics. The time taken for 0.8 mole of A to produce 0.6 mole of B is 1 hour. What is the time taken for conversion of 0.9 mole of A to produce 0.675 mole of B?
1) 2 hours 2) 1 hour 3) 0.5 hour 4) 0.25 hour

30. For a first order reaction $A \rightarrow B$ the reaction rate at reactant concentration of 0.01 M is found to be $2.0 \times 10^{-5} \text{molL}^{-1}\text{s}^{-1}$. The half life period of the reaction is
1) 30 s 2) 220 s 3) 300 s 4) 347 s

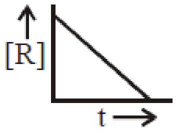
31. For the reaction $A + 2B \rightarrow C$, rate is given by $r = [A]\,[B]^2$ then the order of the reaction is
1) 3 2) 6 3) 5 4) 7

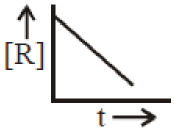
32. For a first order reaction, a plot of log (a – x) against time is a straight line with a negative slope equal to
1) $\frac{-k}{2.303}$ 2) $-2.303k$ 3) $\frac{2.303}{k}$ 4) $-\frac{E_a}{2.303R}$

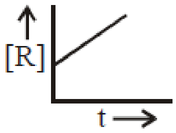
33. In a reaction, when the concentration of reactant is increased two times, the increase in rate of reaction was four times. Order of reaction is :
1) zero 2) 1 3) 2 4) 3

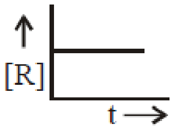
34. The rate constant of a first order reaction is $6.9 \times 10^{-3} \text{s}^{-1}$. How much time will it take to reduce the initial concentration to its 1/8th value?
1) 100 s 2) 200 s 3) 300 s 4) 400 s

35. The plot that represents the zero order reaction is :

(a)


(b)


(c)


(d)


36. Which of the following is correct for a first order reaction?
1) $t_{1/2} \propto \alpha$ 2) $t_{1/2} \propto \frac{1}{\alpha}$ 3) $t_{1/2} \propto \alpha^0$ 4) $t_{1/2} \propto \alpha^2$

37. Point out the wrong statement:
For a first order reaction
1) time for half-change ($t_{1/2}$) is independent of initial concentration
2) change in the concentration unit does not change the rate constant (k)
3) time for half-change × rate constant = 0.693
4) the unit of k is $\text{mole}^{-1}\ \text{min}^{-1}$

38. For a certain reaction, rate = $k \times [H^+]^n$. If pH of reaction changes from two to one, the rate becomes 100 times of its value at pH = 2, the order of reaction is –

1) 1 2) 2 3) 0 4) 3

39. Ethylene is produced by $C_4H_8 \xrightarrow{\Delta} 2C_2H_4$

The rate constant is 2.3×10^{-4} sec^{-1}. Approximately in what time will the molar ratio of ethylene to cyclobutane in mixture attain the value equal to one ?

(log 2 = 0.3, log 3 = 0.47)

1) 1700 sec 2) 1600 sec 3) 2000 sec 4) 1200 sec

40. Which one of the following statements for the order of a reaction is incorrect ?

1) Order can be determined only experimentally.
2) Order is not influenced by stoichiometric coefficient of the reactants.
3) Order of reaction is sum of power to the concentration terms of reactants to express the rate of reaction.
4) Order of reaction is always whole number.

41. For a given reaction $t_{1/2} = 1/ka$. The order of this reaction is

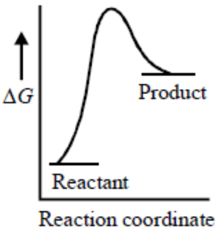
1) 0 2) 1 3) 2 4) 3

42. Half life of a reaction is found to be inversely proportional to the cube of its initial concentration. The order of reaction is

1) 2 2) 5 3) 3 4) 4

43. Which one of the following reaction energy diagrams best represents a reaction in the reverse direction, that it is the most endothermic?

(a)



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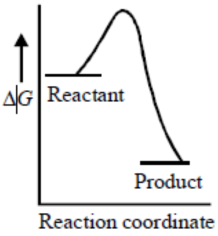
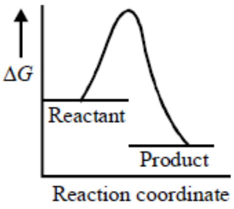
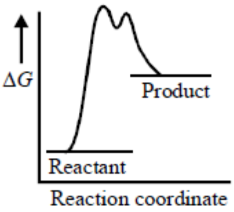


(c)



(d)



TOPIC 3 : Theories of Rate of Reaction

44. In the Haber process for the manufacture of ammonia the following catalyst is used

1) Platinized asbestos 2) Iron with molybdenum as promoter
3) Copper oxide 4) Alumina

45. A reaction having equal energies of activation for forward and reverse reaction has :

1) $\Delta G = 0$ 2) $\Delta H = 0$ 3) $\Delta H = \Delta G = \Delta S = 0$ 4) $\Delta S = 0$

46. A catalyst only

1) decreases activation energy 2) increases activation energy
3) both (1) and (2) 4) comes to equilibrium

47. An example of autocatalytic reaction is

1) decomposition of nitroglycerine
2) thermal decomposition of $KClO_3$ and MnO_2 mixture
3) breakdown of $^{14}C_6$ 4) hydrogenation of vegetable oil using nickel catalysts

48. For an endothermic reaction where ΔH represents the enthalpy of reaction in kJ/mol, the minimum value for the energy of activation will be

1) less than ΔH 2) more than ΔH 3) equal to ΔH 4) zero

49. The activation energies of the forward and backward reactions in the case of a chemical reaction are 30.5 and 45.4 kJ/mol respectively. The reaction is :

1) exothermic 2) endothermic
3) neither exothermic nor endothermic 4) independent of temperature

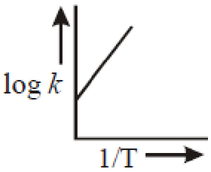
50. According to which theory activation energy and proper orientation of the molecules together determine the criteria for an effective collision?
1) Arrhenius theory 2) Activated complex theory 3) Collision theory 4) Both 1) and 3)

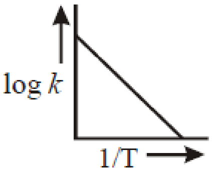
51. For an exothermic reaction, the energy of activation of the reactants is
1) equal to the energy of activation of products
2) less than the energy of activation of products
3) greater than the energy of activation of products
4) sometimes greater and sometimes less than that of the products

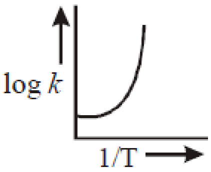
52. The activation energy for a simple chemical reaction $A \rightarrow B$ is E_a in forward direction. The activation energy for reverse reaction
1) is always double of E_a. 2) is negative of E_a.
3) is always less than E_a. 4) can be less than or more than E_a.

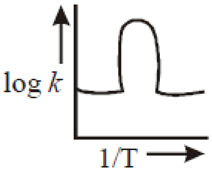
53. For a reaction, activation energy (E_a) = 0 and rate constant (k) = $(k) = 3.2 \times 10^{6} s^{-1}$ at 300 K. What is the value of the rate constant at 310 K?
1) $3.2 \times 10^{-12} s^{-1}$ 2) $3.2 \times 10^{6} s^{-1}$ 3) $6.4 \times 10^{12} s^{-1}$ 4) $6.4 \times 10^{6} s^{-1}$

54. A graph plotted between log k vs 1/T for calculating activation energy is shown by

(a)


(b)


(c)


(d)


55. When a biochemical reaction is carried out in laboratory in the absence of enzyme then rate of reaction obtained is 10^{-6} times, then activation energy of reaction in the presence of enzyme is
1)6/RT 2) different from E_a obtained in laboratory 3) P is required 4) can't say anything

56. For an endothermic reaction, energy of activation is E_a andenthalpy of reaction is ΔH (both of these in kJ/mol). Minimum value of E_a will be
1) less than ΔH 2) equal to ΔH 3) more than ΔH 4) equal to zero

57. In an exothermic reaction if ΔH is the enthalpy then activation energy is
1) more than ΔH 2) less than ΔH 3) equal to ΔH 4) none of the above

58. In respect of the equation $k = Ae^{-Ea/RT}$ in chemical kinetics, which one of the following statements is incorrect?
1) A is adsorption factor 2) E_a is energy of activation 3) R is Rydberg's constant 4) Both 1) and 3)

59. Match the columns.

Column-I	**Column-II**
1) Catalyst alters the rate of reaction	(p) cannot be fraction or zero
2) Molecularity	(q) proper orientation is not there always.
3) Second half life of first	(r) by lowering the activation order reaction energy
4) Energetically favourable	(s) is same as the first reactions are sometimes slow

1) A – (q), B – (r), C – (s), D – (p) 2) A – (r), B – (s), C – (p), D – (q)
3) A – (r), B – (p), C – (s), D – (q) 4) A – (p), B – (r), C – (s), D – (q)

60. In terms of the 'Collision Theory of Chemical Kinetics', the rate of a chemical reaction is proportional to
1) the change in free energy per second. 2) the change in temperature per second.
3) the number of collisions per second. 4) the number of products molecules.

NEET PREVIOUS YEARS QUESTIONS

1. When initial concentration of the reactant is doubled, the half-life period of a zero order reaction **[2018]**
1) is halved 2) is doubled 3) remains unchanged 4) is tripled

2. The correct difference between first and second order reactions is: **[2018]**
1) The rate of a first-order reaction does not depend on reactant concentrations, the rate of a second-order reaction does depend on reactant concentrations
2) The half-life of a first-order reaction does not depend on $[A]_0$, the half-life of a second-order reaction depends on $[A]_0$
3) The rate of a first-order reaction does depend on reactant concentrations, the rate of a second-order reaction does not depend on reactant concentrations
4) A first-order reaction can be catalyzed, a second-order reaction cannot be catalysed

3. Mechanism of a hypothetical reaction **[2017]**
$X_2 + Y_2 \rightarrow 2XY$ is given below :
i) $X_2 \rightarrow X + X(\text{fast})$ ii) $X + Y_2$ □ $XY + Y(\text{slow})$ iii) $X + Y \rightarrow XY(\text{fast})$
The overall order of the reaction will be :
1) 2 2) 0 3) 1.5 4) 1

4. A first order reaction has a specific reaction rate of 10^{-2} sec^{-1}. How much time will it take for 20g of the reactant to reduce to 5 g? **[2017]**
1) 138.6 sec 2) 346.5 sec 3) 693.0 sec 4) 238.6 sec

5. The rate of a first -order reaction is 0.04 mol $l^{-1}s^{-1}$ at 10 seconds and 0.03 mol $l^{-1}s^{-1}$ at 20 seconds after initiation of the reaction. The half-life period of the reaction is **[2016]**
1) 24.1 s 2) 34.1 s 3) 44.1 s 4) 54.1 s

6. The addition of a catalyst during a chemical reaction alters which of the following quantities? **[2016]**
1) Entropy 2) Internal energy 3) Enthalpy 4) Activation energy

7. The rate constant of the reaction $A \rightarrow B$ is 0.6×10^{-3} mole per second. If the concentration of A is 5 M then concentration of B after 20 minutes is : **[2015]**
1) 1.08 M 2) 3.60 M 3) 0.36 M 4) 0.72 M

8. The activation energy of a reaction can be determined from the slope of which of the following graphs? **[2015]**
1) $\frac{\ln K}{T}$ vs.T 2) $\ln K vs. \frac{1}{T}$ 3) $\frac{T}{\ln K} vs \frac{1}{T}$ 4) $\ln K$ vs .T

9. When initial concentration of a reactant is doubled in a reaction, its half-life period is not affected. The order of the reaction is : **[2015]**
1) First2) Second 3) More than zero but less than first 4) Zero

10. If the rate constant for a first order reaction is k, the time (t) required for the completion of 99% of the reaction is given by :- **[2019]**
(1) t = 0.693/k (2) t = 6.909/k (3) t = 4.606/k (4) t = 2.303/k

11. For the chemical reaction $N_2(g) + 3H_2(g)$□ □□ $2NH_3(g)$ the correct option is : **[2019]**
(1) $-\frac{1}{3}\frac{d[H_2]}{dt} = -\frac{1}{2}\frac{d[NH_3]}{dt}$ (2) $-\frac{d[N_2]}{dt} = 2\frac{d[NH_3]}{dt}$
(3) $-\frac{d[N_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$ (4) $3\frac{d[H_2]}{dt} = 2\frac{d[NH_3]}{dt}$

12. A first order reaction has a rate constant of 2.303×10^{-3} s^{-1}. The time required for 40g of this reactant to reduce to 10 g will be- [Given that $\log_{10}2=0.3010$] **[2019 ODISSA]**
(1) 230.3 s (2) 301 s (3) 2000 s (4) 602 s

13. For a reaction, activation energy $E_a=0$ and the rate constant at 200K is $1.6 \times 10^6 s^{-1}$. The rate constant at 400K will be- [Given that gas constant] R= 8.314 J K^{-1} mol^{-1} **[2019 ODISSA]**

(1) 3.2×10^4 s^{-1} (2) 1.6×10^6 s^{-1} (3) 1.6×10^3 s^{-1} (4) 3.2×10^6 s^{-1}

14. The half-life for a zero order reaction having 0.02 M initial concentration of reactant is 100 s.

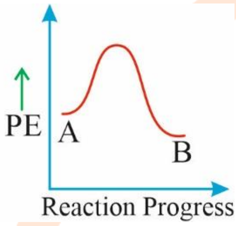
The rate constant (in mol $L^{-1}s^{-1}$) for the reaction is **[2020 – COVID-19]**

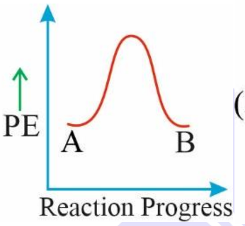
(1) 1.0×10^{-4} (2) 2.0×10^{-4} (3) 2.0×10^{-3} (4) 1.0×10^{-2}

15. In collision theory of chemical reaction, Z_{AB} represents **[2020 – COVID-19]**
(1) the fraction of molecules with energies greater than E_a
(2) the collision frequency of reactants, A and B
(3) steric factor
(4) the fraction of molecules with energies equal to E_a

16. An increase in the concentration of the reactants of a reaction leads to change in : **[2020]**
1. Collision frequency 2. Activation energy 3. heat of reaction 4. Threshold energy

17. For a reaction $A \rightarrow B$, enthalpy of reaction is -4.2 kJ mol^{-1} and enthalpy of activation is 9.6 kJ mol^{-1}. The correct potential energy profile for the reaction is shown in option **[NEET-2021]**

(1)


(2)


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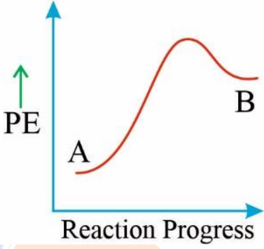
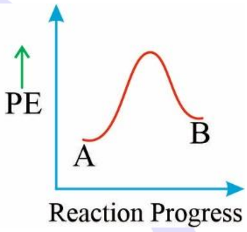


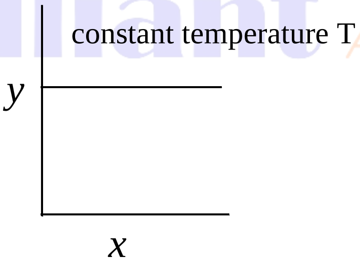
(4)


18. The slope of Arrhenius Plot $\left(\ln k\, v/s\, \frac{1}{T}\right)$ of first order reaction is. Choose the correct option for your answer. **[NEET-2021]**

$\left[Given\ R = 8.314 L\ JK^{-1} mol^{-1}\right]$

1) $83.0\, kJ\ mol^{-1}$ 2) $166\, kJ\ mol^{-1}$ 3) $-83\, kJ\ mol^{-1}$ 4) $41.5\, kJ\ mol^{-1}$

19. The given graph is a representation of kinetics of a reaction **[NEET-2022]**



The y and x axes for zero and first order reactions, respectively are
1)zero order (y=concentration and x= time), first order(y = $t_{1/2}$ and x = concentration)
2)zero order (y= concentration and x= time), first order (y= rate constant and x = concentration)
3)zero order (y= rate and x = concentration), first order (y= $t_{1/2}$ and x = concentration)
4)zero order (y= rate and x = concentration), first order (y= rate and x =$t_{1/2}$)

20. A 10.0 L flask contains 64g of oxygen at 27^{0}C. (Assume O_2 gas is behaving ideally).The pressure inside the flask in bar is (Given R = 0.0831 L bar K^{-1} mol^{-1}) **[NEET-2022]**
1)2.5 2)498.6 3)49.8 4)4.9

21. For a first order reaction A $\longrightarrow$ products, initial concentration of A is 0.1M which becomes 0.001M after 5 minutes. Rate constant for the reaction in $\min^{-1}$ is **[NEET-2022]**
1)1.3818 2)0.9212 3) 0.4606 4)0.2303

NCERT LINE BY LINE QUESTIONS – ANSWERS

(1.)	d	**(2.)**	c	**(3.)**	a	**(4.)**	a	**(5.)**	a
(6.)	c	**(7.)**	d	**(8.)**	a	**(9.)**	c	**(10.)**	b
(11.)	d	**(12.)**	d	**(13.)**	d	**(14.)**	c	**(15.)**	c
(16.)	d	**(17.)**	a	**(18.)**	c	**(19.)**	a	**(20.)**	b
(21.)	d	**(22.)**	a	**(23.)**	a	**(24.)**	d	**(25.)**	c
(26.)	b	**(27.)**	a	**(28.)**	b	**(29.)**	a	**(30.)**	a
(31.)	b	**(32.)**	b	**(33.)**	a	**(34.)**	b	**(35.)**	d
(36.)	a	**(37.)**	b	**(38.)**	b	**(39.)**	a	**(40.)**	b
(41.)	c	**(42.)**	d	**(43.)**	d	**(44.)**	d	**(45.)**	a
(46.)	b	**(47.)**	c	**(48.)**	b	**(49.)**	d	**(50.)**	a

TOPIC WISE PRACTICE QUESTIONS – ANSWERS

1) 4	**2)** 2	**3)** 3	**4)** 2	**5)** 2	**6)** 4	**7)** 1	**8)** 2	**9)** 1	**10)** 2
11) 3	**12)** 4	**13)** 1	**14)** 3	**15)** 3	**16)** 2	**17)** 3	**18)** 4	**19)** 1	**20)** 2
21) 4	**22)** 3	**23)** 3	**24)** 2	**25)** 2	**26)** 1	**27)** 1	**28)** 2	**29)** 2	**30)** 4
31) 1	**32)** 1	**33)** 3	**34)** 3	**35)** 4	**36)** 3	**37)** 4	**38)** 2	**39)** 1	**40)** 4
41) 3	**42)** 4	**43)** 2	**44)** 2	**45)** 2	**46)** 3	**47)** 2	**48)** 2	**49)** 1	**50)** 3
51) 2	**52)** 4	**53)** 2	**54)** 2	**55)** 2	**56)** 2	**57)** 4	**58)** 4	**59)** 3	**60)** 3

NEET PREVIOUS YEARS QUESTIONS-ANSWERS

1) 2	**2)** 2	**3)** 3	**4)** 1	**5)** 1	**6)** 4	**7)** 4	**8)** 2	**9)** 1
10) 3	**11)** 3	**12)** 4	**13)** 2	**14)** 1	**15)** 1	**16)** 1	**17)** 1	**18)** 4
19) 2	**20)** 4	**21)** 2						

NCERT LINE BY LINE QUESTIONS – SOLUTIONS

(1.) (d) Catalyst has no effect on Gibb's free energy of system and pre-exponential factor of a chemical reaction.

(4.) (a) $A + 2B \rightarrow C$

Reaction rate is doubled when the concentration of A is doubled again when both A and B are made four times, reaction rate also become four times.

(5.) (a) $t_{t/2} = \frac{0.693}{K}$ For first order $t_{1/2}$ is independent of concentration.

(6.) (c)

$$K = \frac{2.303}{t} \log \frac{[A]_0}{[A]_t} = \frac{2.303}{4606} \log \frac{80}{8} = 5 \times 10^{-4} s^{-1}$$

(7.) (d) In a reaction $3A \rightarrow 2B$

$$\text{Rate of reaction } = \frac{1}{3}\frac{d[A]}{dt} = +\frac{1}{2}\frac{d[B]}{dt}$$

$$\text{Therefore, } \frac{d[B]}{dt} = -\frac{2}{3}\frac{d[A]}{dt}$$

(8.) (a) Lowering of Ea, raises the value of K.

(9.) (c) For first order reaction

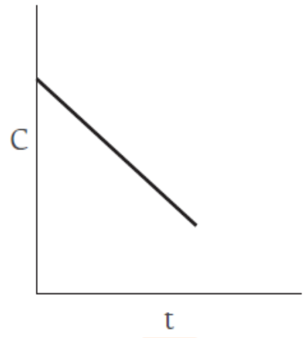
$$k = \frac{2.303}{t} \log \frac{[A]_0}{[A]_t}$$

$$10^{-2} = \frac{2.303}{t} \log \frac{20}{5}$$

$$r = \frac{2.303}{10^{-2}} \log 4 = \frac{2.303 \times 0.6020}{10^{-2}} = 138.64 \text{ s}$$

(10.) (b)

(12.) (d) For zero-order reaction the plot of rate vs concentration is given by the graph (iv)



(13.) (d) We known that,

$$\log \frac{K}{A} = -\frac{Ea}{2.303RT} = -\frac{2.303RT}{2.303RT} = -1$$

$$\log \frac{K}{A} = -1$$

$$\frac{K}{A} = \text{antilog}(-1) = 0.1$$

(14.) (c) For First order reaction,

$$K = \frac{1}{t} \text{In}\left(\frac{A_0}{A_r}\right)$$

$$K = \frac{1}{t} \ln\left[\frac{100}{1}\right]$$

$$K = \frac{2.303 \log_{10} 100}{t}$$

$$K = \frac{2.303 \times 2}{t}$$

$K = \frac{4.606}{t}$ For 99% completion

$[A]_0 = 100$

$t = \frac{4.606}{K}$ $[A]_r = 1$

(15.) (c) For first order reaction $\text{Rate}_1 = K[A_1]$

According *to* question,

$[A_2] = [2A_1]$

$\therefore \text{Rate}_2 = K[2A_1] = \text{Rate}_2 = 2\text{Rate}_1$

For a given reaction, rate constant is constant and independent of the concentration of reactant.

(18.) (c) When temperature is increased, the number of active molecules (i.e. the molecules with activation energy) increases. Therefore, the number of effective collisions will increase and the rate of reaction will also increase.

(19.) (a) At low pressure, rate is proportional to the surface-coverage and is of first order while at high pressure, it follow zero order kinetics due to complete coverage of surface area.

(20.) (b) The value of rate constant of a pseudo first order reaction depends not only on temperature but also on concentration of reactant present in excess. Hence, (b) is correct.

It means reaction rate is independent of concentration of B.

With respect to 'A' the order is definitely one.

$$\text{w.r.t.'A' order } = 1$$
$$\text{w.r.r.'B' order } = 0$$
$$\text{Overall order } = 1$$

(23.) (a) $(Ea) = 90 - 50 = 40\text{kJ}$

(24.) (d) $K = \dfrac{0.693}{t_{1/2}} = \dfrac{0.693}{14.26} = 0.0486$

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

$$t = \frac{2.303}{0.0486}\log\frac{100}{100-95}$$

$$t = 61.65\text{ min} = 62\text{ min}$$

(26.) (b) $K_1 = Ae^{-Ea/RT_1}$

$$K_2 = Ae^{-Ea/RT_2}$$

$$\ln K_1 = \ln A - \frac{Ea}{RT_1}$$

$$\text{In } K_2 = \text{In A} - \frac{Ea}{RT_2}$$

From equation (i) and (ii), we have

$$\text{In } K_2 - \text{In } K_1 = \text{In } A - \frac{Ea}{RT_2} - \text{In } A + \frac{Ea}{RT_1} = \text{In}\frac{K_2}{K_1} = \frac{Ea}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$= \text{In}\frac{K_2}{K_1} = \frac{-Ea}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

(28.) (b) $(r_{1/2})_{\text{zero}} = -\dfrac{[A]_0}{2k}$

$$\therefore \text{If } [A]_0 = \text{Doubled}$$

$$t_{1/2} = \text{Doubled}$$

(33.) (a) For a 1st order kinetics,

$$K = \frac{2.303}{r}\log 10\frac{a}{a-x}$$

$$\text{At } t_{1/2}, K = \frac{2.303}{t_{1/2}}\log 10\frac{a}{a-\frac{a}{2}}$$

$$0 - t_{1/2} = \frac{2.303}{K}\log 10^2 = \frac{\ln 2}{K}$$

(34.) (b)

(37.) (b) Presence of catalyst does not change ΔG or ΔH value. Hence,(b) is not correct.

(38.) (b) After 3 months,

$t_{1/2} =$

$$\begin{array}{ccccccccc} & & 1 & & 2 & & 3 & & 4 \\ 400g & \rightarrow & 200g & \rightarrow & 100g & \rightarrow & 50g & \rightarrow & 25g \\ \text{Sample} & & B & & & & & & A \end{array}$$

'A' or 'B'

For element A, 4 half lives $= 3$ months (faster)

For element B, 1 half life $= 3$ months (slower)

Hence, half life of 'B' is 4 times that of < A'.

(39.) (a) Let the rate according to the rate law $=K[A]^X[B]^Y$

Then from experiments

$$5.5\times10^4 = K[0.4]^x(0.2)^Y \text{…..(i)}$$

$$5.5\times10^4 = K[0.8]^X(0.2)^Y \text{…..(ii)}$$

$$2.2\times10^3 = K[0.4]^X(0.4)^Y \text{…..(iii)}$$

From equations (i) and (ii)

$$\frac{5.5\times10^{-4}}{5.5\times10^{-4}} = \frac{K[0.4]^X[0.2]^Y}{K[0.8]^x[0.2]^Y}$$

$$1=\left(\frac{0.4}{0.8}\right)^X = \left(\frac{1}{2}\right)^X = X = 0$$

From equations (i) and (iii), we have

$$\frac{5.5\times10^{-4}}{2.2\times10^{-3}} = \frac{K[0.4]^X[0.2]^Y}{K[0.4]^X[0.4]^Y}$$

$$\frac{1}{4} = \left(\frac{0.2}{0.4}\right)^Y = \left(\frac{1}{2}\right)^Y$$

$y = 2$

Thus, rate law

Rate $=K[A]^0[B]^2$

(40.) (b)

(41.) (c) Rate $=\frac{1}{5}\frac{\Delta[Br^-]}{\Delta t} = -\frac{\Delta[BrO_3^-]}{\Delta t} = -\frac{1}{6}\frac{\Delta[H^+]}{\Delta t}$

$$\frac{\Delta[Br^-]}{\Delta t} = 5\frac{\Delta[BrO_3^-]}{\Delta t} = \frac{5}{6}\frac{\Delta[H^+]}{\Delta t}$$

(42.) (d) For first-order reaction $t_{1/2} = \frac{0.693}{K}$ which is independent of initial concentration of reactant.

*For second order reaction $r_{1/2} = \frac{1}{K}[A_0]$ which depends on initial concentration of reactant.

(44.) (d) For first order reaction,

$$c = \frac{2.303}{K}\log\frac{a}{a-x} = \frac{t30\%}{t50\%} = \frac{\frac{2.303}{K}\log\frac{100}{100-30}}{\frac{2.303}{K}\log\frac{100}{100-50}}$$

$$= \frac{30}{t_{1/2}} = \frac{\log\frac{10}{7}}{\log 2} = \frac{30}{t_{1/2}} = \frac{0.1549}{0.3010}$$

$$t_{1/2} = \frac{0.3010\times30}{0.1549} = 58.29 \text{ min}$$

(48.) (b) $r = K[A]^2$

When concentration of ‘A’ is doubled,

$r = K[2A]^2 = 4K[A]^2$

(49.) (d) If product concentration is X.

For a zero order reaction $\frac{X}{t} = k$

Thus graph would be a straight line passing through origin.
The given information is for zero-order reaction.
For a zero-order reaction, rate of the reaction is constant.

Plot of rate vs time i.e., $-\frac{d[X]}{dt}$ vs time will be a straight line parallel to X-axis.

(50.) (a) Rate expression $\frac{dX}{dt} = K[A]^m[B]^n$ shows that the overall order of reaction is $m+n+0=m+n$ as the rate of reaction is independent of concentration of *C*, i.e., the order with respect to *C* is zero.

TOPIC WISE PRACTICE QUESTIONS - SOLUTIONS

1. (4) Since all have same concentration of reactants, all would react at same time.
2. (2) Surface area increases.
3. (3)
4. (2)
5. 2) $3A \rightarrow 2B$ Rate of appearance of B is equal to rate of disappearance of A.

$\frac{1}{2}\frac{d[B]}{dt} = -\frac{1}{3}\frac{d[A]}{dt} \Rightarrow \frac{d[B]}{dt} = -\frac{2}{3}\frac{d[A]}{dt}$

6. 4)
7. 1)
8. 2)
9. 1) $\frac{k_{34^0}}{k_{35^0}} < 1$ also, $k \propto T$ Hence, rate increase with rise of temperature.
10. 2) $2NO(g)+Cl_2(g)$ □ $2NOCl(g)$

Rate $= k[NO]^2[Cl]$

Rate constant is independent of the initial concentration of the reactants. It has a constant value at fixed temperature. Hence the value of rate constant can be increased by increasing the temperature.

11. 3) $k = \left(\text{mol lit}^{-1}\right)^{1-n}$ time^{-1} for given reaction $n = 2$

$\therefore k = \text{mol}^{-1}\text{lit sec}^{-1}$

12. 4) order of reaction may be zero, whole number or fractional.
13. 1) For a zero order reaction

Rate $= k[A]^0$ i.e. rate $= k$

Hence unit of k = M. sec^{-1}
For a first order reaction
Rate $= k[A]$

$k = M.\text{sec}^{-1}/M = \text{sec}^{-1}$

14. 3) $N_2O \rightarrow N_2 + \frac{1}{2}O_2$

$\frac{dx}{dt} \propto [N_2O]^1$

i.e. order of reaction $= 1$

15. 3) $t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0693} = 10\,\text{min}$

Reactant after 10 min $= 5$ mol

Rate $\left(\frac{dx}{dt}\right) = k[A] = 0.0693 \times 5\text{mol min}^{-1}$

16. 2)
17. 3) $x = 0.1 - 0.08 = 0.02M$;

$$k=\frac{x}{t}=\frac{0.02}{10}=2\times10^{-3}\text{M min}^{-1}$$

$$\therefore t_{1/2}=\frac{[A]_0}{2k}=\frac{0.1}{2\times2\times10^{-3}}=25\,\text{min}$$

$$t_{\text{completion}}=2\times t_{1/2}=50\,\text{min}$$

18. 4) $t_{1/2}\propto\frac{1}{a^2}$ we know that $t_{1/2}\propto\frac{1}{a^{n-1}}$ i.e. n = 3 Thus reaction is of 3rd order

19. 1) as doubling the initial conc. Doubles the rate of reaction, order $=1$

20. 2) $k_1=\frac{1}{5}\text{In}\frac{0.2}{0.1};k_2=\frac{1}{10}\text{In}\frac{0.2}{0.05};k_1=k_2$

21. 4) Reaction is second order; $k=10^{-4}\,\text{L/mol.min}$

$$\therefore -\frac{1}{3}\frac{d[A]}{dt}=k[A]^2$$

$$-\frac{d[A]}{dt}=10^{-4}\times3\times(0.5)^2$$

$$=7.5\times10^{-5}\,\text{M}\,\text{min}^{-1}=\frac{7.5\times10^{-5}}{60}=1.25\times10^{-6}\,\text{Ms}^{-1}$$

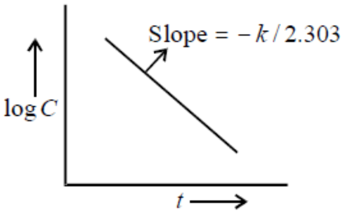
22. 3) $k=\frac{2.303}{t}\log\left(\frac{a}{a-x}\right)$

$$k=\frac{2.303}{t}\log\left(\frac{100}{1}\right)\Rightarrow\frac{2.303}{16}$$

$$t=\frac{2.303\times16}{2.303}\log\left(\frac{100}{0.1}\right)=48\,\text{min}$$

23. 3) $t_{3/4}=\frac{1}{k}\ln\frac{100}{\frac{1}{4}\times100}=t_{3/4}=\frac{2.303}{k}\log\frac{1}{\frac{1}{4}\times1}\Rightarrow\frac{2.303}{k}\log4$

24. 2)



25. 2) Given : 75% reaction gets completed in 32 min thus,

$$k=\frac{2.303}{t}\log\frac{a}{(a-x)}=\frac{2.303}{32}\log\frac{100}{(100-75)}=\frac{2.303}{32}\log4=0.0433\,\text{min}^{-1}$$

Now we can use this value of *k* to get the value of time required for 50% completion of reaction

$$t=\frac{2.303}{k}\log\frac{a}{(a-x)}=\frac{2.303}{0.0433}\log\frac{100}{50}=\frac{2.303}{0.0433}\log2=16\,\text{min}$$

26. 1) Given initial concentration 1) = 2.00 M; Time taken (*t*) = 200 min and final concentration (*a* – *x*) = 0.15 M. For a first order reaction rate constant,

$$k=\frac{2.303}{200}\log\frac{a}{a-x}=\frac{2.303}{200}\log\frac{2.00}{0.15}=\frac{2.303}{200}\times(0.301+0.824)=1.29\times10^{-2}\,\text{min}^{-1}$$

Further

$$(t_{1/2})=\frac{0.693}{k}=\frac{0.693}{1.29\times10^{-2}}=53.72\,\text{min}$$

27. 1) $t_{1/2}=4\text{s}\quad T=16\text{s}$

$$n=\frac{T}{t_{1/2}}=\frac{16}{4}=4\left(\therefore T=n\times t_{1/2}\right)$$

$$[A]=[A]_0\left(\frac{1}{2}\right)^n=0.12\times\left(\frac{1}{2}\right)^4=\frac{0.12}{16}=0.0075M$$

28. 2) For reaction $3A\rightarrow B+C$

If it is zero order reaction $r=k[A]^0$,i.e., the rate remains same at any concentration of 'A'. i.e., independent of concentration of A.

29. 2) $A\rightarrow B$ For a first order reaction

Given a = 0.8 mol, (a-x) = 0.8-0.6 = 0.2

$$k=\frac{2.303}{1}\log\frac{0.8}{0.2}\text{ or }k=2.303\log 4$$

Again $a=0.9, a-x=0.9-0.675=0.225$

$$k=\frac{2.303}{t}\log\frac{0.9}{0.225}$$

$$2.303\log 4=\frac{2.303}{t}\log 4\text{ Hence }t=1\text{ hour}$$

30. 4) Given $[A]=0.01M$

Rate = $2.0\times10^{-5}\text{mol L}^{-1}\text{s}^{-1}$

For a first order reaction

Rate = $k[A]$

$$k=\frac{2.0\times10^{-5}}{[0.01]}=2\times10^{-3}$$

$$t_{1/2}=\frac{0.693}{2\times10^{-3}}=347\,\text{sec}$$

31. 1) Order is the sum of the power of the concentration terms in rate law expression.

32. 1) $t=\frac{2.303}{k}\log\frac{a}{a-x}$ or $t=\frac{2.303}{k}\log\ a-\frac{2.303}{k}\log(a-x)$

33. (3) Since rate of reaction becomes four times on doubling concentration of reactant, it is second order reaction.

34. (3)

35. (4) For a zero order, rate of reaction does not change with time.

36. 3) $t_{1/2}\propto\frac{1}{a^n-1}$ for first order reaction, n = 1

$\therefore t_{1/2}\propto a^0$

37. 4) Unit of k for Ist order reaction is (time)$^{-1}$

38. 2) $pH=2;r_1=k\times\left(10^{-2}\right)^n\{\because\left[H^+\right]=10^{-pH}\}$

$$pH=1;r_2=k\times\left(10^{-1}\right)^n$$

$$\text{Given }r_2=100r_1\Rightarrow\left(\frac{10^{-1}}{10^{-2}}\right)^n=100$$

$$10^n=100$$

$$\therefore n=2$$

39. 1) $C_4H_8\rightarrow 2C_2H_4$

Moles at $t=0$ $\quad a \quad\quad 0$

Moles at $t=t$ $\quad (a-x) \quad\quad 2x$

$$\text{when } \frac{2x}{a-x}=1 \text{ then } x=\frac{a}{3}$$

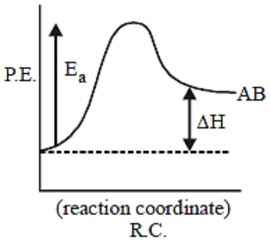
$$\therefore\ t=\frac{2.303}{k}\log\frac{a}{(a-x)}$$

$$=\frac{2.303}{2.3\times10^{-4}}\log\frac{a}{a-\frac{a}{3}}=1700\text{ sec}$$

40. (4) Order of reaction may be zero, whole number or fractional.
41. (3)
42. (4) Half life of a reaction is found to be inversely proportional to the cube of initial concentration. The order of reaction is 4.
43. (2) In reverse direction, product becomes reactant and vice versa.
44. (2) In Haber's process, ammonia is manufactured from N_2 and H_2 using iron as catalyst with molybdenum as promoter at high temperature and pressure

$$N_2+3H_2\xrightarrow[\text{Mo(catalytic promoter)}]{Fe_2O_3\text{(catalyst)}}2NH_3$$

45. 2) $\Delta H=E_{a_f}-E_{a_b}=0$
46. 3)
47. 2)
48. 2)



E_a is always greater than ΔH

49. (1) Exothermic because of activation energy $E_b > E_f$
50. (3)
51. (2) Activation energy of reactant is less than the energy of activation of products.
52. (4) The activation energy of reverse reaction will depend upon whether the forward reaction is exothermic or endothermic.
53. (2) When $E_a = 0$, rate constant is independent of temperature.
54. 2) $\log k=\log A-\dfrac{E_a}{2.303RT}$
55. 2)
56. 2)
57. 4) The activation energy depends upon the nature of chemical bonds undergoing rupture during chemical reaction and is independent of enthalpies of reactants and products
58. 4) In equation $k=Ae^{-E_a/RT}$; A = Frequency factor k = velocity constant, R = gas constant and E_a = energy of activation
59. (3)
60. (3)

1. 2) $\left(t_{1/2}\right)_{zero} = \frac{[A]_0}{2k}$; $\therefore$ If $[A]_0$ = doubled, $t_{1/2}$ = doubled

2. 2) $\left(t_{1/2}\right)$ 1^{st} order = Independent of concentration ; $\left(t_{1/2}\right)$ 2^{nd} order $\propto \frac{1}{[A]_0}$

3. 3) The overall reaction rate depends on the rate of the slowest step.
i.e., Overall rate = Rate of slowest step (ii)
$= k[X][Y_2]$..............(1)
K = rate constant
Assuming step (i) to be reversible, its equilibrium constant,
$$k_{eq} = \frac{[X]^2}{[X_2]} \Rightarrow [X]^2 = k_{eq}[X_2]; [X] = k_{eq}^{\frac{1}{2}}[X_2]^{1/2}$$
From eq(1) and (2)
Rate $= kk_{eq}^{\frac{1}{2}}[X_2]^{\frac{1}{2}}[Y_2]$; Overall order = $\frac{1}{2}+1=\frac{3}{2}=1.5$

4. 1) Half life for a first order reaction,
$t_{1/2} = \frac{0.693}{K}$ So, $t_{1/2} = \frac{0.693}{10^{-2}}$ sec Also, for the reduction of 20 g of reactant to 5 g, two half lives will be required.
$\therefore$ For 20 g of the reactant to reduce to 5g, time taken, $t = 2\times\frac{0.693}{10^{-2}}\text{sec} = 138.6\,\text{sec}$

5. 1) For a first order reaction
$$k = \frac{2.303}{(t_2-t_1)}\log\frac{(a-x_1)}{(a-x_2)}\ ; \quad k = \frac{2.303}{(20-10)}\log\frac{(0.04)}{(0.03)}$$
$$k = \frac{2.303\times0.1249}{10}$$
$$\frac{0.6932}{t_{1/2}} = \frac{2.303\times0.1249}{10}\ ;\ t_{1/2} = \frac{0.6932\times10}{2.303\times0.1249} = 24.1\,\text{sec}$$

6. (4) A catalyst provides an alternative route for the reaction with a lower activation energy.

7. (4) Rate constant k = 0.6 × 10^{-3} mole per second. (unit mole per second shows zero order reaction)
For a zero order reaction
$[A] = [A]_0 - kt$ and $[A]_0 - [A] = [B] = kt = 0.6\times10^{-3}\times20\times60 = 0.72M$

8. 2)Arrhenius equation
$$K = A.e^{-E_a/RT} \Rightarrow \ln K = \ln A - \frac{E_a}{RT}$$
Slope $= \frac{-E_a}{R}$
so, activation energy of reaction can be determined from the slope of ln K vs 1/T

9. 1) $t_{1/2} = \frac{0.693}{K}$ For first order $t_{1/2}$ is independent of initial concentration of reactant

10. 3)For first order reaction
$k = \frac{1}{t}\ln\left[\frac{A_0}{A_t}\right]$ For 99% completion, $[A]_0$=100, $[A]_t$=1
$$k = \frac{1}{t}\ln\left[\frac{100}{1}\right]\ ;\ k = \frac{2.303\log_{10}100}{t}$$
$$k = \frac{2.303\times2}{t}\ ;\ k = \frac{4.606}{t} \Rightarrow t = \frac{4.606}{k}$$

11. 3)

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$$-\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = +\frac{1}{2}\frac{d[NH_3]}{dt}$$

12. 4)

For a first order reaction; $t_{1/2} = \frac{0.693}{K}$

$$t_{1/2} = \frac{0.693}{2.303 \times 10^{-3}} = 301 \text{ s}$$

The time required for 40 g of reactant to reduce to 10g

$$t_{75\%} = 2 \times t_{1/2}$$

$$t_{75\%} = 2 \times 301 = \mathbf{602 \text{ s}}$$

13. 2)

$$\log\left(\frac{K_2}{K_1}\right) = \frac{E_a}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$E_a = 0$$

$$\log\left(\frac{K_2}{K_1}\right) = 0$$

$$\frac{K_2}{K_1} = 10^{\circ} = 1 \Rightarrow K_2 = K_1$$

$\mathbf{K_2 = 1.6 \times 10^6 \ s^{-1}}$ at 400K

14. (1)

$$(t_{1/2})_{zero} = \frac{[A]_0}{2K}$$

$$100s = \frac{0.02M}{2K}$$

$$K = \frac{0.02M}{2 \times 100} = 1 \times 10^4 \text{ mol L}^{-1} \text{ s}^{-1}$$

15. (1) Z_{AB} = Collision frequency

16. 1)As concentration increases, no of reactant molecules present per unit volume increases, hence collision frequency increases

17. 1) $\Delta H_{rxn} = (E_a)_f - (E_a)_b$

$$-4.2 = (E_a)_f - (E_a)_b$$

$$-4.2 = 9.6 - (E_a)_b$$

$$(E_a)_b = 9.6 + 4.2 = 13.8 \text{kJmol}^{-1}$$

Since reaction is exothermic, so possible graph is (1) only.

Also $(E_a)_f < (E_a)_h$, so answer is option (1).

18. 4) $\ln k = \ln A - \frac{E_a}{R} \cdot \frac{1}{T}$

$$\text{Slope} = -\frac{E_a}{R} = -5 \times 10^3$$

$$E_a = 5 \times 10^3 \times 8.314 = 41.5 \times 10^3 J = 41.5 kJ$$

19. For zero order rate is independent of conc. Of reactants also $t^{1/2}$ and [C] are independent for first order kinetics

20. $$PV = \frac{w}{m-w} \times R \times T$$

$$P = \frac{w}{m-w} \times \frac{RT}{V} = \frac{64}{32} \times \frac{0.0831 \times 300k}{10} = 4.9$$

21 $$K = \frac{2.303}{t} \log\left(\frac{a}{a-x}\right) = \frac{2.303}{5} \log\left(\frac{0.1}{0.001}\right) = 0.9212$$