

Chemical Kinetics

Question1

Catalysts are used to increase the rate of a chemical reaction. Because it

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Options:

A.

Increases the activation energy of the reaction

B.

Decreases the activation energy of the reaction

C.

Brings about improper orientation of reactant molecules

D.

Increases the potential energy barrier

Answer: B

Solution:

Positive catalyst decreases the activation energy to increase rate of reaction.

Question2

Half-life of a first order reaction is 20 seconds and initial concentration of reactant is 0.2 M . The concentration of reactant



left after 80 seconds is

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Options:

A. 0.1 M

B. 0.05 M

C. 0.0125 M

D. 0.2 M

Answer: C

Solution:

To find the concentration of a reactant after a certain time in a first-order reaction, we use the concept of half-life.

Given:

The half-life ($t_{1/2}$) of the reaction is 20 seconds.

Initial concentration ($[A_0]$) is 0.2 M.

We want to find the concentration after 80 seconds.

Calculation:

Determine the number of half-lives:

$$\text{Number of half-lives} = \frac{t}{t_{1/2}} = \frac{80}{20} = 4$$

Calculate the concentration after 4 half-lives:

For a first-order reaction, the concentration after n half-lives is given by:

$$[A_t] = [A_0] \left(\frac{1}{2}\right)^n$$

Substituting the values:

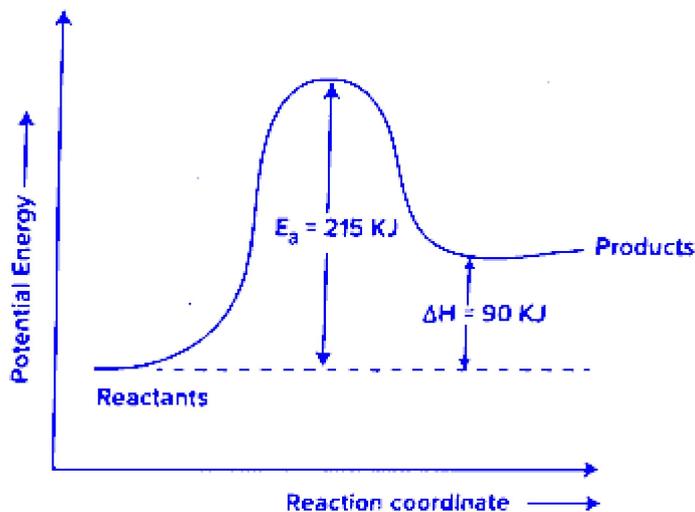
$$[A_t] = 0.2 \times \left(\frac{1}{2}\right)^4 = 0.2 \times \frac{1}{16} = 0.0125 \text{ M}$$

Thus, the concentration of the reactant after 80 seconds is 0.0125 M.



Question3

In the given graph, E_a for the reverse reaction will be



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Options:

- A. 125 KJ
- B. 215 KJ
- C. 90 KJ
- D. 305 KJ

Answer: A

Solution:

To determine the activation energy (E_a) for the reverse reaction from the given graph, we can use the following energy relationship:

$$\Delta H = (E_a)_{\text{forward}} - (E_a)_{\text{backward}}$$

From the graph, we have:

$$\Delta H = 90 \text{ KJ}$$

$$(E_a)_{\text{forward}} = 215 \text{ KJ}$$

Plugging these values into the equation, we get:



$$90 = 215 - (E_a)_{\text{backward}}$$

Solving for $(E_a)_{\text{backward}}$:

$$(E_a)_{\text{backward}} = 215 - 90 = 125 \text{ KJ}$$

Therefore, the activation energy for the reverse reaction is 125 KJ.

Question4

For the reaction $2 \text{N}_2\text{O}_5(\text{g}) \rightarrow 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ initial concentration of N_2O_5 is 2.0molL^{-1} and after 300 min, it is reduced to 1.4molL^{-1} . The rate of production of NO_2 ($\text{in molL}^{-1} \text{min}^{-1}$) is

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Options:

A. 2.5×10^{-4}

B. 4×10^{-4}

C. 2.5×10^{-3}

D. 4×10^{-3}

Answer: D

Solution:

To determine the rate of production of NO_2 for the reaction $2 \text{N}_2\text{O}_5(\text{g}) \rightarrow 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$, we use the fact that the rate of production of a product is related to the rate of disappearance of a reactant.

Given:

Initial concentration of N_2O_5 is 2.0 mol/L .

After 300 minutes, the concentration is reduced to 1.4 mol/L .

Calculate the rate of disappearance of N_2O_5 :

$$\text{Rate of disappearance of } \text{N}_2\text{O}_5 = \frac{-(1.4-2.0)}{300} = \frac{-(-0.6)}{300} = \frac{0.6}{300}$$



As per the stoichiometry of the reaction, the rate of formation of NO_2 is twice the rate of disappearance of N_2O_5 :

$$\text{Rate of formation of } \text{NO}_2 = 2 \times \frac{0.6}{300} = \frac{1.2}{300} = 4 \times 10^{-3} \text{ mol/L/min}$$

Therefore, the rate of production of NO_2 is $4 \times 10^{-3} \text{ mol/L/min}$.

Question 5

For the reaction, $A \rightleftharpoons B$, $E_a = 50 \text{ kJ mol}^{-1}$ and $\Delta H = -20 \text{ kJ mol}^{-1}$. When a catalyst is added E_a decreases by 10 kJ mol^{-1} . What is the E_a for the backward reaction in the presence of catalyst?

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Options:

- A. 60 kJ mol^{-1}
- B. 40 kJ mol^{-1}
- C. 70 kJ mol^{-1}
- D. 20 kJ mol^{-1}

Answer: A

Solution:

For the reaction $A \rightleftharpoons B$, the activation energy (E_a) for the forward reaction is initially 50 kJ/mol and the enthalpy change (ΔH) is -20 kJ/mol . Thus, the activation energy for the backward reaction ($E_{a,\text{backward}}$) is calculated as:

$$E_{a,\text{backward}} = E_a + |\Delta H|$$

Substituting the given values:

$$E_{a,\text{backward}} = 50 \text{ kJ/mol} + 20 \text{ kJ/mol} = 70 \text{ kJ/mol}$$

When a catalyst is added, the activation energy for the forward reaction decreases by 10 kJ/mol :

$$E_{a,\text{forward, catalyst}} = 50 \text{ kJ/mol} - 10 \text{ kJ/mol} = 40 \text{ kJ/mol}$$

A catalyst equally reduces the activation energy for both the forward and backward reactions. Therefore, the new activation energy for the backward reaction in the presence of the catalyst is:

$$E_{a,\text{backward, catalyst}} = 70 \text{ kJ/mol} - 10 \text{ kJ/mol} = 60 \text{ kJ/mol}$$

Thus, the correct option is

Option A

60 kJ/mol

Question 6

For the reaction, $\text{PCl}_5 \longrightarrow \text{PCl}_3 + \text{Cl}_2$, rate and rate constant are $1.02 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$ and $3.4 \times 10^{-5} \text{ s}^{-1}$ respectively at a given instant. The molar concentration of PCl_5 at that instant is

KCET 2024

Options:

- A. 8.0 mol L^{-1}
- B. 3.0 mol L^{-1}
- C. 0.2 mol L^{-1}
- D. 2.0 mol L^{-1}

Answer: B

Solution:

To determine the molar concentration of PCl_5 at a given instant for the reaction $\text{PCl}_5 \rightarrow \text{PCl}_3 + \text{Cl}_2$, we use the rate law expression:

$$\text{Rate} = k [\text{PCl}_5]$$

Given values are:

$$\text{Rate} = 1.02 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

$$\text{Rate constant, } k = 3.4 \times 10^{-5} \text{ s}^{-1}$$

We can find the concentration $[\text{PCl}_5]$ using the formula by rearranging it as follows:

$$\text{Rate} = k [\text{PCl}_5]$$

Substitute the given values into the equation:

$$1.02 \times 10^{-4} = 3.4 \times 10^{-5} [\text{PCl}_5]$$

To solve for $[\text{PCl}_5]$, divide both sides by 3.4×10^{-5} :

$$[\text{PCl}_5] = \frac{1.02 \times 10^{-4}}{3.4 \times 10^{-5}}$$

After calculating the division, we find:

$$[\text{PCl}_5] = 3 \text{ mol L}^{-1}$$

Thus, the molar concentration of PCl_5 at that instant is 3 mol L^{-1} .

Question 7

Which one of the following does not represent Arrhenius equation?

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Options:

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

B. $k = Ae^{-\frac{E_a}{RT}}$

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

D. $k = Ae^{\frac{E_a}{RT}}$

Answer: D

Solution:

Among the given options, option (d) i.e. $K = Ae^{\frac{E_a}{RT}}$ does not represent Arrhenius equation.



Question8

In which one of the following reactions, rate constant has the unit $\text{mol L}^{-1} \text{s}^{-1}$?

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Options:

A. Acid catalysed hydrolysis of $\text{CH}_3\text{COOCH}_3$

B. $\text{CHCl}_3 + \text{Cl}_2 \longrightarrow \text{CCl}_4 + \text{HCl}$

C. $2\text{NO}(g) + \text{O}_2(g) \longrightarrow 2\text{NO}_2(g)$

D. Decomposition of HI on the surface of gold

Answer: D

Solution:

The given unit of the rate constant i.e. $\text{mole L}^{-1} \text{s}^{-1}$ suggests that the reaction is of zero order.

So, among the given options, decomposition of HI on the surface of gold is a zero order reaction.

Question9

For a reaction, the value of rate constant at 300 K is $6.0 \times 10^5 \text{s}^{-1}$. The value of Arrhenius factor A at infinitely high temperature is

KCET 2023

Options:

A. $6 \times 10^5 \times e^{-E_a/300R}$

B. $e^{-E_a/300R}$

C. $\frac{6 \times 10^{-5}}{300}$

D. 6×10^5

Answer: D

Solution:

Given, $k = 6 \times 10^5 \text{ s}^{-1}$

Using the formula of Arrhenius equation,

$$k = e^{E_a/RT}$$

At $T = \alpha$, $\frac{-E_a}{RT}$ will become zero,

So, $k = Ae^0 \Rightarrow k = A$

Thus, $k = A = 6 \times 10^5$

Question10

The rate constant k_1 and k_2 for two different reactions are $10^{16} \times e^{-2000/T}$ and $10^{15} \times e^{-1000/T}$ respectively. The temperature at which $k_1 = k_2$ is

KCET 2023

Options:

A. $\frac{2000}{2.303}$ K

B. 2000 K

C. $\frac{1000}{2.303}$ K

D. 1000 K

Answer: C

Solution:

$$K_1 = 10^{16} \times e^{-\frac{2000}{T}}$$

$$K_2 = 10^{15} \times e^{-\frac{1000}{T}}$$



Now, $K_1 = K_2$

$$10^{16} \times e^{\frac{-2000}{T}} = 10^{15} \times e^{\frac{-1000}{T}}$$
$$10 = e^{\frac{1000}{T}}$$

Taking \ln on both sides,

$$\ln 10 = \frac{1000}{T} \ln e$$

$$2.303 \log 10 = \frac{1000}{T} \Rightarrow T = \frac{1000}{2.303} \text{ K}$$

Question11

For n th order of reaction, half-life period is directly proportional to

KCET 2022

Options:

A. $\frac{1}{a^{1-n}}$

B. a^{n-1}

C. a^{1-n}

D. $\frac{1}{a^{n-1}}$

Answer: C

Solution:

For n th order of reaction, half-life period is directly proportional to a^{1-n} .

Question12

Half-life of a reaction is found to be inversely proportional to the fifth power of its initial concentration, the order of reaction is

KCET 2022

Options:

A. 4

B. 5

C. 6

D. 3

Answer: C

Solution:

The half-life period of n th order reaction is

$$t_{1/2} = \frac{k}{a^{n-1}}$$

a is the initial concentration of reactant $t_{1/2} \propto \frac{1}{a^5}$

$$\therefore t_{1/2} = \frac{k}{a^5} \text{ and } \frac{k}{a^{n-1}} = \frac{k}{a^5}$$

$$\Rightarrow a^5 = a^{n-1}$$

$$n - 1 = 5 = 6$$

\therefore The order of the reaction is 6.

Question13

A first order reaction is half completed in 45 min. How long does it need 99.9% of the reaction to be completed?

KCET 2022

Options:

A. 7.5 Hours

B. 10 Hours

C. 20 Hours



D. 5 Hours

Answer: A

Solution:

$t_{1/2}$ for first order reaction is 45 min

$$t_{1/2} = \frac{0.693}{k} \Rightarrow k = \frac{0.693}{t_{1/2}} = \frac{0.693}{45}$$

$$t = \frac{2.303}{k} \log \frac{a}{a-x} = \frac{2.303}{0.693} \times 45 \log \frac{100}{0.1}$$
$$= 448 \text{ min} \simeq 7.5 \text{ hrs}$$

Question14

The rate of the reaction,

$\text{CH}_3\text{COOC}_2\text{H}_5 + \text{NaOH} \longrightarrow \text{CH}_3\text{COONa} + \text{C}_2\text{H}_5\text{OH}$ is given by the equation, $\text{rate} = k [\text{CH}_3\text{COOC}_2\text{H}_5][\text{NaOH}]$. If concentration is expressed in molL^{-1} , the unit of k is

KCET 2022

Options:

A. $\text{mol L}^{-1} \text{s}^{-1}$

B. $\text{L mol}^{-1} \text{s}^{-1}$

C. s^{-1}

D. $\text{mol}^{-2} \text{L}^2 \text{S}^{-1}$

Answer: B

Solution:

Given that, $\text{rate} = k [\text{CH}_3\text{COOC}_2\text{H}_5][\text{NaOH}]$

That means the reaction is a second order

$$\therefore \frac{dx}{dt} = k[A]^2 \Rightarrow \frac{\text{conc}}{\text{time}} = k[\text{conc.}]^2$$

$$\frac{\text{mol L}^{-1}}{\text{s}} = k \text{ mol L}^{-1} \times \text{mol L}^{-1}$$

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

Question 15

For a reaction, $A + 2B \rightarrow \text{Products}$, when concentration of B alone is increased half-life remains the same. If concentration of A alone is doubled, rate remains the same. The unit of rate constant for the reaction is

KCET 2021

Options:

A. s^{-1}

B. $\text{L mol}^{-1} \text{ s}^{-1}$

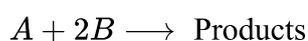
C. $\text{mol L}^{-1} \text{ s}^{-1}$

D. atm^{-1}

Answer: A

Solution:

For the reaction,



Half-life ($t_{1/2}$) remains same as $[B]$ increases for 1st order, $t_{1/2} \propto \frac{1}{[B]}$

\therefore w.r.t. B , reaction is of first order.

Also, as $[A]$ is doubled, rate remains constant.

\therefore It is zero order w.r.t. A .

\therefore Rate law expression is $k[A]^0[B]^1$

\therefore Overall order = $0 + 1 = 1$



\therefore Unit of $k = \text{s}^{-1}$

Question16

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

KCET 2021

Options:

A. $t = \frac{4.606}{k}$

B. $t = \frac{2.303}{k}$

C. $t = \frac{0.693}{k}$

D. $t = \frac{6.909}{k}$

Answer: A

Solution:

Given, for a first order reaction, let initial concentration = $[R_0]$

After 99% completion, concentration

$$\begin{aligned}[R] &= R_0 - 0.99R_0 \\ &= 0.1R_0\end{aligned}$$

Using formula, $t = \frac{2303}{k} \log \frac{[R_0]}{[R]}$

$$= \frac{2303}{k} \log \frac{100}{1} = \frac{4.606}{k}$$

Question17

The rate of a gaseous reaction is given by the expression $k[A][B]^2$. If the volume of vessel is reduced to one half of the initial volume, the reaction rate as compared to original rate is



KCET 2021

Options:

A. $\frac{1}{16}$

B. $\frac{1}{8}$

C. 8

D. 16

Answer: C

Solution:

Given, rate law expression (r) = $k[A][B]^2$

Final volume (V_2) is reduced to half of initial volume (V_1) i.e. $V_2 = 1/2V_1$

\therefore Final concentration $[A]$ and $[B] = [2A]$ and $2[B]$

\therefore Final rate $r' = k[A]^0[B]^2 = k[2A][2B]^2$

$$r' = k \times 2 \times 4[A][B]^2 \Rightarrow r' = 8k[A][B]^2$$

$\therefore r' = 8r$

i.e. rate increases by 8 times.

Question18

Higher order (> 3) reactions are rare due to

KCET 2021

Options:

A. shifting of equilibrium towards reactants due to elastic collisions

B. loss of active species on collision



C. low probability of simultaneous collision of all reacting species

D. increase in entropy as more molecules are involved

Answer: C

Solution:

The rarity of higher order reactions (those involving more than three reacting species simultaneously) is primarily due to Option C: low probability of simultaneous collision of all reacting species.

In a chemical reaction, reactant molecules must collide with the proper orientation and sufficient energy for a reaction to occur. As the number of reactant molecules involved increases, the likelihood of all these molecules meeting at the same time and place with the correct orientation and enough energy decreases dramatically. This is why higher order reactions (those of order greater than three) are rare. The complexity and improbability of such a simultaneous occurrence make these reactions almost nonexistent under normal conditions.

Option A talks about shifting of equilibrium towards reactants due to elastic collisions. This does not directly explain why higher order reactions are rare but rather describes a possible dynamic in reactions where the collisions do not lead to reaction but rather to the dispersal of energy.

Option B involves the loss of active species on collision, which could in principle apply to any reaction order and does not specifically address the rarity of high order reactions.

Option D suggests an increase in entropy as more molecules are involved, which may contribute to the favorability of the overall reaction but does not explain why the reactions of higher order are rare. Entropy increases with more molecules involved, but this does not inherently prevent high order reactions; it is the collision probability that is the key issue.

Therefore, the best explanation for the rarity of higher order reactions is the low probability of a simultaneous collision involving all the reactants, as described in Option C.

Question19

The time required for 60% completion of a first order reaction is 50 min. The time required for 93.6% completion of the same reaction will be

KCET 2020

Options:

A. 100 min

B. 83.8 min



C. 50 min

D. 150 min

Answer: D

Solution:

Time for 60% completion = 50 min for first order reaction.

Time required for 93.6% completion = ?

$$k = \frac{2.303}{50} \log \frac{a}{a - 0.6a} \Rightarrow k = \frac{2.303}{50} \log \frac{a}{0.4a}$$

$$k = \frac{2.303}{50} \log 2.5$$

$$\text{Also, } k = \frac{2.303}{x} \log \frac{a}{a - 0.936a}$$

$$\therefore \frac{2.303 \log 2.5}{50} = \frac{2.303}{x} \log \frac{a}{0.064a}$$

$$x = 50 \times \frac{\log \frac{1000}{64}}{\log 2.5}$$

$$= 50 \times \frac{\log 15.625}{\log 2.5} = \frac{1.19 \times 50}{0.398} = \frac{59.5}{0.398}$$

$$= 150 \text{ min}$$

Question20

For an elementary reaction $2A + 3B \rightarrow 4C + D$ the rate of appearance of C at time ' T ' is $2.8 \times 10^{-3} \text{ mol L}^{-1} \text{ S}^{-1}$. Rate of disappearance of B at ' t ' t will be

KCET 2020

Options:

A. $\frac{4}{3} (2.8 \times 10^{-3}) \text{ mol L}^{-1} \text{ S}^{-1}$

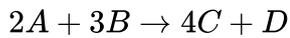
B. $\frac{3}{4} (2.8 \times 10^{-3}) \text{ mol L}^{-1} \text{ S}^{-1}$

C. $2 (2.8 \times 10^{-3}) \text{ mol L}^{-1} \text{ S}^{-1}$

$$D. \frac{1}{4} (2.8 \times 10^{-3}) \text{ mol L}^{-1} \text{ S}^{-1}$$

Answer: B

Solution:



$$\frac{d(C)}{dt} = 2.8 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1} \text{ (at time } t \text{)}$$

$$- \frac{d(B)}{dt} = ?$$

$$- \frac{1}{3} \frac{d(B)}{\Delta t} = + \frac{1}{4} \frac{d[C]}{\Delta t}$$

Rate of disappearance of $B = \frac{3}{4} \times$ rate of appearance of C

$$= \frac{3}{4} \times 2.8 \times 10^{-3}$$

Question21

The rate constant of a reaction is given by $k = PZe^{-E_a/RT}$ under standard notation. In order to speed up the reaction, which of the following factors has to be decreased?

KCET 2020

Options:

A. Z

B. Both Z and T

C. E_a

D. T

Answer: C

Solution:

$$\therefore k = PZe^{-E_a/RT}$$



The above reaction expression is given in terms of exponents. We have to increase the speed of reaction. So, value of k has to be more and as per the given expression.

$$k \propto \frac{1}{e^{E_a/RT}}$$

if E_a decreases, then $[e^{E_a/RT}]$ also decreases and $\left[\frac{1}{e^{E_a/RT}}\right]$ increases.

$\therefore k$ also increases

Question22

The plot of $t_{1/2}$ v/s $[R]_0$ for a reaction is a straight-line parallel to X -axis. The unit for the rate constant of this reaction is

KCET 2019

Options:

- A. $\text{mol L}^{-1} \text{s}$
- B. $\text{mol L}^{-1} \text{s}^{-1}$
- C. $\text{L mol}^{-1} \text{s}^{-1}$
- D. s^{-1}

Answer: D

Solution:

The plot of $t_{1/2}$ vs $[R]_0$ for a reaction is a straight line parallel to X -axis. This type of plot is obtained in first order reaction. The unit for the rate constant of first order reaction is obtained as follows $K = \frac{\text{Rate}}{[A]^x[B]^y}$

where, $x + y = 1$ (for first order reaction)

$$\begin{aligned} \text{So, } K &= \frac{\text{Concentration}}{\text{Time}} \times \frac{1}{(\text{Concentration})^n} \\ &= \frac{\text{mol L}^{-1}}{\text{sec}} \times \frac{1}{(\text{mol L}^{-1})^1} = \text{sec}^{-1} \text{ or } \text{s}^{-1} \end{aligned}$$

Question23

Which is a wrong statement?

KCET 2019

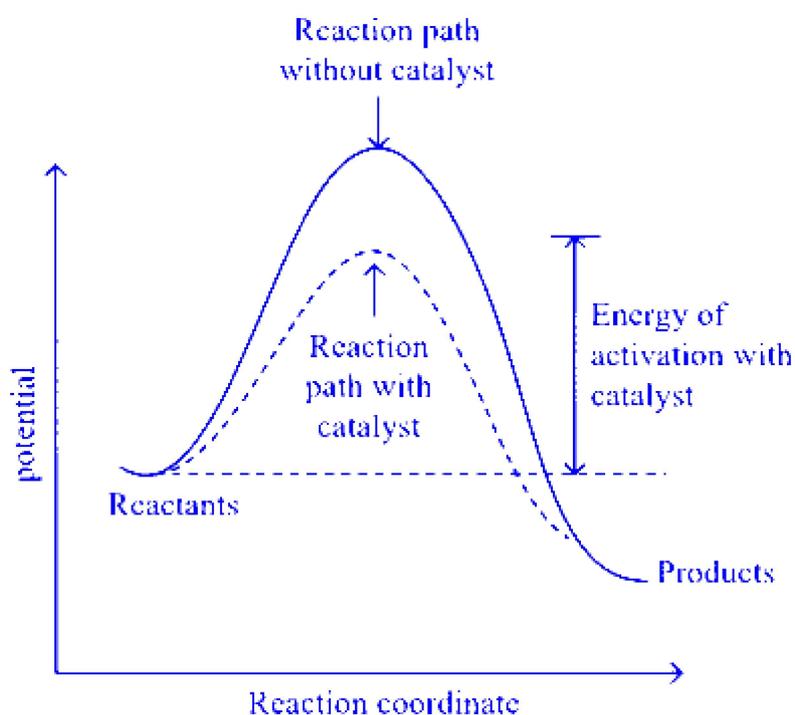
Options:

- A. Rate constant $k =$ Arrhenius constant A : if $E_a = 0$
- B. $e^{-E_a/RT}$ gives the fraction of reactant molecules that are activated at the given temperature
- C. In k vs $\frac{1}{T}$ plot is a straight line
- D. presence of catalyst will not alter the value of E_a

Answer: D

Solution:

Presence of catalyst alter the value of E_a i.e. activation energy. It provides an alternate pathway or reaction mechanism by reducing the activation energy between reactants and products, hence lowering the potential energy barrier as shown in the above figure.



Question24

1L of 2M CH_3COOH is mixed with 1L of 3M $\text{C}_2\text{H}_5\text{OH}$ to form an ester. The rate of the reaction with respect to the initial rate when each solution is diluted with an equal volume of water will be

KCET 2019

Options:

- A. 0.25 times
- B. 2 times
- C. 0.5 times
- D. 4 times

Answer: A

Solution:

When 1 L of 2M CH_3COOH is mixed with 1 L of 3M $\text{C}_2\text{H}_5\text{OH}$, the total volume of the solution is 2 L.

The initial concentrations of the reactants before dilution are:

$$[\text{CH}_3\text{COOH}] = 2 \text{ M}$$

$$[\text{C}_2\text{H}_5\text{OH}] = 3 \text{ M}$$

After diluting each of these solutions with an equal volume of water (1 L), the final volume of each solution becomes 2 L. The new concentrations are:

$$[\text{CH}_3\text{COOH}] = \frac{2\text{M}}{2} = 1 \text{ M}$$

$$[\text{C}_2\text{H}_5\text{OH}] = \frac{3\text{M}}{2} = 1.5 \text{ M}$$

If the reaction is first-order with respect to both acetic acid (CH_3COOH) and ethanol ($\text{C}_2\text{H}_5\text{OH}$), the rate law is given by:

$$\text{Rate} = k \cdot [\text{CH}_3\text{COOH}] \cdot [\text{C}_2\text{H}_5\text{OH}]$$

Let the initial rate be:

$$\text{Initial Rate} = k \cdot 2 \text{ M} \cdot 3 \text{ M} = 6k$$

The rate after dilution is:

$$\text{New Rate} = k \cdot 1 \text{ M} \cdot 1.5 \text{ M} = 1.5k$$

The ratio of the new rate to the initial rate is:

$$\frac{\text{New Rate}}{\text{Initial Rate}} = \frac{1.5k}{6k} = \frac{1.5}{6} = 0.25$$

Therefore, the rate of the reaction with respect to the initial rate when each solution is diluted with an equal volume of water will be:

Option A: 0.25 times

Question 25

For the reaction, $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$, the rate of disappearance of O_2 is $2 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$. The rate of appearance of SO_3 is

KCET 2018

Options:

A. $2 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$

B. $4 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$

C. $1 \times 10^{-1} \text{ mol L}^{-1} \text{ s}^{-1}$

D. $6 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$

Answer: B

Solution:

Let's analyze the reaction step by step:

The balanced chemical equation is:



The rate of reaction is defined in terms of the stoichiometric coefficients. For any species, the rate is given by:

$$\text{Rate} = -\frac{1}{(\text{Coefficient})} \frac{d[\text{Reactant}]}{dt} = \frac{1}{(\text{Coefficient})} \frac{d[\text{Product}]}{dt}$$

Given the rate of disappearance of O_2 is:

$$\frac{d[\text{O}_2]}{dt} = -2 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

The coefficient of O_2 in the reaction is 1, so the rate of the reaction is:

$$\text{Rate} = -\frac{1}{1} \frac{d[\text{O}_2]}{dt} = 2 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$



For SO_3 , the coefficient is 2. The rate of appearance of SO_3 is related to the rate of reaction by:

$$\frac{d[\text{SO}_3]}{dt} = 2 \times \text{Rate}$$

Substitute the rate:

$$\frac{d[\text{SO}_3]}{dt} = 2 \times (2 \times 10^{-4}) = 4 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

Thus, the rate of appearance of SO_3 is $4 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$. This corresponds to Option B.

Question 26

The temperature coefficient of a reaction is 2. When the temperature is increased from 30°C to 90°C , the rate of reaction is increased by

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Options:

- A. 150 times
- B. 410 times
- C. 72 times
- D. 64 times

Answer: D

Solution:

The temperature coefficient tells us that for every 10°C increase in temperature, the reaction rate doubles (i.e., increases by a factor of 2). Here's how we calculate the overall increase when the temperature is raised from 30°C to 90°C :

Determine the temperature difference:

$$\Delta T = 90^\circ\text{C} - 30^\circ\text{C} = 60^\circ\text{C}$$

Calculate the number of 10°C increments:

$$\text{Number of increments} = \frac{60}{10} = 6$$

Since the rate doubles with each 10°C increase, the overall factor is:

$$\text{Overall increase} = 2^6 = 64$$

Therefore, the reaction rate increases by 64 times, which corresponds to Option D.

Question27

The value of rate constant of a pseudo first order reaction

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Options:

- A. depends only on temperature
- B. depends on the concentration of reactants present in small amounts
- C. depends on the concentration of reactants present in excess
- D. is independent of the concentration of reactants

Answer: C

Solution:

In a *pseudo-first-order* reaction, one (or more) of the reactants is taken in large excess so that its concentration remains essentially constant during the reaction. As a result, although the true rate law might be something like

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

if [B] is in large excess and remains nearly constant, it can be lumped into an *effective* or *apparent* rate constant:

$$k_{\text{pseudo}} = k [B]^n \quad (\text{constant in time if [B] is in excess}).$$

Hence, the observed (pseudo-first-order) rate law looks like

$$\text{Rate} = k_{\text{pseudo}} [A]^m,$$

making it *appear* to be first order in [A] alone.

Key Point

Because $k_{\text{pseudo}} = k [B]^n$, it *does* depend on:

The true rate constant k (which depends on temperature via the Arrhenius equation), and

The concentration [B] of the reactant present in large excess.

Of the given options, the *dominant distinguishing factor* in pseudo-first-order kinetics is that the *effective* rate constant includes the *constant* concentration of the reactant in excess. Thus:



The value of the (pseudo-first-order) rate constant depends on the concentration of the reactant in excess.

Conclusion

The correct choice is:

Option C: depends on the concentration of reactants present in excess.

Question28

For a reaction $\frac{1}{2}A \rightarrow 2B$ rate of disappearance of A is related to rate of appearance of B by the expression

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Options:

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

B. $\frac{-d[A]}{dt} = 4 \frac{d[B]}{dt}$

C. $\frac{-d[A]}{dt} = \frac{1}{4} \frac{d[B]}{dt}$

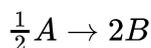
D. $\frac{-d[A]}{dt} = \frac{d[B]}{dt}$

Answer: C

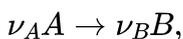
Solution:

Let's analyze the reaction step by step.

Consider the reaction:



For any reaction represented as



the reaction rate is defined as:

$$r = -\frac{1}{\nu_A} \frac{d[A]}{dt} = \frac{1}{\nu_B} \frac{d[B]}{dt}.$$

Here, for our reaction:



The stoichiometric coefficient for A is $\nu_A = \frac{1}{2}$.

The stoichiometric coefficient for B is $\nu_B = 2$.

Step 1: Write the rate expressions for A and B.

For species A:

$$r = -\frac{1}{\frac{1}{2}} \frac{d[A]}{dt} = -2 \frac{d[A]}{dt}.$$

For species B:

$$r = \frac{1}{2} \frac{d[B]}{dt}.$$

Step 2: Equate the two expressions for r since they represent the same reaction rate:

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

Step 3: Solve for $\frac{-d[A]}{dt}$ in terms of $\frac{d[B]}{dt}$.

Multiply both sides of the equation by $\frac{1}{2}$ to isolate $-\frac{d[A]}{dt}$:

$$-\frac{d[A]}{dt} = \frac{1}{4} \frac{d[B]}{dt}.$$

Thus, the rate of disappearance of A is related to the rate of appearance of B by:

$$\frac{-d[A]}{dt} = \frac{1}{4} \frac{d[B]}{dt}.$$

Reviewing the options, we see that this corresponds to Option C.

So, the correct answer is Option C.

Question29

Which of the following statement is incorrect?

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Options:

- A. The rate of law for any reaction cannot be determined experimentally.
- B. Complex reactions have fractional order.
- C. Biomolecular reactions involved simultaneous collision between two species.
- D. Molecularity is only applicable for elementary reaction

Answer: A

Solution:

Option A: "The rate law for any reaction cannot be determined experimentally."

In chemical kinetics, the rate law of a reaction is typically determined by experiments. By varying the concentrations of the reactants and measuring the reaction rate, scientists can determine the order (the exponents in the rate law) with respect to each reactant.

For example, for a reaction with rate law

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

experiments are used to find the values of m and n .

Therefore, the statement in Option A is incorrect.

Option B: "Complex reactions have fractional order."

In many non-elementary (complex) reactions, the overall rate law may show fractional orders. This happens because the mechanism of the reaction involves multiple steps or equilibria, leading to exponents that are not whole numbers.

Thus, Option B can be a valid statement.

Option C: "Bimolecular reactions involved simultaneous collision between two species."

Typically, a bimolecular (often referred to as two-molecule) reaction involves a simultaneous collision between two reacting species.

Even though the term "bimolecular" might be mistaken for reactions involving biomolecules, if we interpret it as bimolecular (elementary two-reactant collisions), then the statement is correct.

Option D: "Molecularity is only applicable for elementary reaction."

Molecularity refers to the number of species colliding in an elementary reaction step. It is not defined for overall complex reactions because these can occur in several steps.

Therefore, Option D is also correct.

Given this analysis, the incorrect statement is:

Option A

Question30

Which of the following statement is in accordance with the Arrhenius equations?

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Options:



- A. Rate of reaction does not change with increase in activation energy
- B. Rate constant decreases exponentially with increase in temperature
- C. Rate of a reaction increases with increases in temperature
- D. Rate of a reaction increases with decreases in activation energy

Answer: A

Solution:

(c, d) The rate of reaction increases with temperature. It becomes double with every 10° change. It increases with decreases in activation energy.

