Chemical Kinetics

Question1

$R \longrightarrow P$ is a first order reaction. The concentration of R changed from 0.04 to $0.03 \mathrm{molL}^{-1}$ in 40 min . What is the average velocity of the reaction in $\mathrm{molL}^{-1} \mathrm{~s}^{-1}$?

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Options:

A.

2.5×10^{-4}

B.

4.167×10^{-6}

C.

4.167×10^{6}

D.

2.5×10^{-5}

Answer: B

Solution:

The change in concentration, $\Delta[R]$ is the final concentration minus initial concentration.

$\Delta[R]=0.03-0.04=-0.01 \mathrm{~mol} / \mathrm{L}$

$$\begin{aligned} \text{Average velocity } & =-\frac{\Delta[R]}{\Delta t}=-\left(-\frac{0.01}{40}\right) \\ & =0.00025 \mathrm{~mol}^{-1} \mathrm{~min}^{-1} \end{aligned}$$

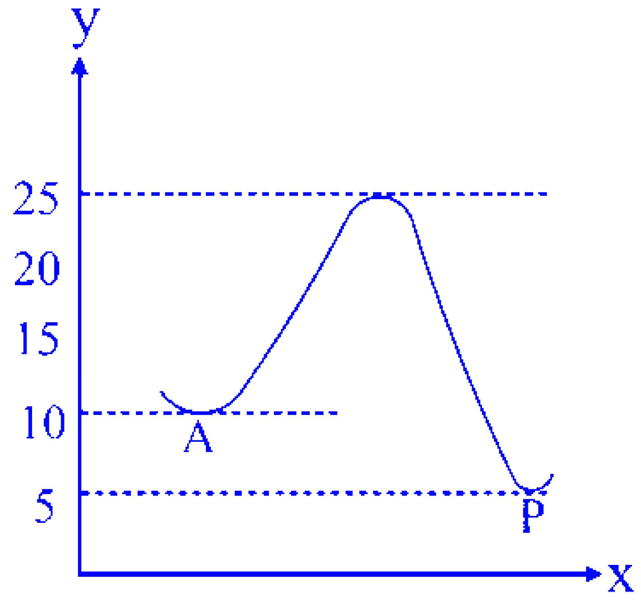
or $2.5 \times 10^{-4} \mathrm{~mol}^{-1} \mathrm{~L}^{-1} \mathrm{~min}^{-1}$

In $\mathrm{Mol^{-1}\,L^{-1}\,s^{-1}} = 4.167 \times 10^{-6}\ \mathrm{mol^{-1}\,L^{-1}\,s^{-1}}$

Question2

The following graph is obtained for a first order reaction $(A \to P)$. The activation energy (E_a in $\mathrm{kJmol^{-1}}$) and heat of reaction ($|\Delta H|$ in $\mathrm{kJmol^{-1}}$) for this reaction are respectively

($x =$ reaction coordinate; $y = E$ in $\mathrm{kJmol^{-1}}$)



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Options:

A.

5,15

B.

15,5

C.

25,5

D.

10,25

Answer: B

Solution:

From the graph,

Energy of reactant $A = 10 \mathrm{~kJ} / \mathrm{mol}$

Energy of product $P = 5 \mathrm{~kJ} / \mathrm{mol}$

Energy of transition state (peak) $= 25 \mathrm{~kJ} / \mathrm{mol}$.

Activation energy $E_a =$ Energy of transition sate - Energy of reactant

$25 - 10 = 15 \mathrm{~kJ} / \mathrm{mol}$

Heat of reaction $\Delta H =$ Energy of product

- Energy of reactant

$\Delta H = 5 - 10 = -5 \mathrm{~kJ} / \mathrm{mol}$

$|\Delta H| = 5 \mathrm{~kJ} / \mathrm{mol}$

Question3

For a first order reaction, the ratio between the time taken to complete $\frac{3}{4}$ th of the reaction and time taken to complete half of the reaction is

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Options:

A.

2

B.

3

C.

1.5

D.

2.5

Answer: A

Solution:

For half completion $[A]_t = \frac{[A]_0}{2}$

$t_{1/2} = \frac{2.303}{K} \log \frac{[A]_0}{[A]_{0/2}}$

Simplify $t_{1/2} = \frac{0.693}{K} \quad \ldots (i)$

for $3/4$ completion

$$[A]_t = \frac{1}{4}[A]_0$$
$$t_{3/4} = 2 \times t_{1/2} \quad \ldots (ii)$$

Taking ratio of Eqs. (i) and (ii),

$\frac{t_{3/4}}{t_{1/2}} = \frac{2 \times t_{1/2}}{t_{1/2}} = 2 : 1$

Question4

The following equation is obtained for a first order reaction at 300 K

$\log_{10} \frac{k}{A} = 0.00174$

What is the activation energy (in Jmol^{-1}) of the reaction?

$\left(R = 8314 \text{ J mol}^{-1} \text{ K}^{-1}\right)$

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Options:

A.

10.0

B.

100.0

C.

0.1

D.

1.0

Answer: A

Solution:

The given equation is

$\log_{10} \frac{k}{A} = 0.00174$

Multiply 2.303 on both side

$$\ln\left(\frac{k}{A}\right) = 2.303 \times 0.00714$$
$$\Rightarrow \quad \ln\left(\frac{k}{A}\right) = 0.00400722 \qquad \ldots (i)$$

The Arrhenius equation is

$k = Ae^{-E_e/RT} \Rightarrow \frac{k}{A} = e^{-Ea/RT}$

Taking natural log

$\ln\left(\frac{k}{A}\right) = -\frac{E_a}{RT} \qquad \ldots (ii)$

Compare Eqs. (i) and (ii)

$-\frac{E_a}{RT} = 0.00400722$

Substituting all values

$E_a \simeq 10.0 \text{ J/mol}$

Question5

$A \rightarrow B$ is a first order reaction. The concentration of A is decreased from x mol L^{-1} to y mol L^{-1} in 100 min . What is the average velocity of the reaction in $\mathrm{molL}^{-1} \mathrm{~min}^{-1}$?

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Options:

A.

$\frac{|x-y|}{100}$

B.

$\frac{|y-x|^2}{100}$

C.

$\frac{100}{|x-y|}$

D.

$\frac{100}{|x+y|}$

Answer: A

Solution:

We want to find the average velocity (rate) for the reaction as A goes to B when the amount of A drops from x to y mol/L in 100 minutes.

The change in the concentration of A is:$\Delta[A] = y - x$ mol/LThis tells us how much A has changed in value.

The average velocity means how much A changes per minute. We can find it by dividing the total change in A by the time taken:Average velocity $= \frac{\Delta[A]}{\Delta t}$where $\Delta t = 100$ minutes.

So,Average velocity $= \frac{y-x}{100}$ mol L^{-1} min^{-1}If you need the answer as a positive value (since rates are usually positive), write it as$\frac{|x-y|}{100}$ mol L^{-1} min^{-1}

Question6

In a first order reaction, the concentration of the reactant is reduced to $1/8$ of the initial concentration in 75 minutes. The $t_{1/2}$ of the reaction (in minutes) is $(\log 2 = 0.30, \log 3 = 0.47, \log 4 = 0.60)$

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Options:

A.

60.2

B.

50.2

C.

25.1

D.

75.1

Answer: C

Solution:

The concentration becomes $\frac{1}{8}$ of what it was at the start.

$\frac{1}{8}=\left(\frac{1}{2}\right)^3$

This means it takes 3 half-lives for the concentration to become $\frac{1}{8}$.

$$t_{1/2}=\frac{75}{3}$$
$t_{1/2}=25$ minutes

Question7

At $T(\mathrm{~K})$ the following equation is obtained for a first order reaction $\log \frac{k}{A}=-\frac{x}{T}$. The activation energy for this reaction is equal to ($R=$ gas constant)

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Options:

A.

$2.303 \times R$

B.

$\frac{2.303R}{x}$

C.

$\frac{x}{2.303R}$

D.

$\frac{1}{2.303 \times R}$

Answer: A

Solution:

Given equation,

$\log\left(\frac{K}{A}\right) = -\frac{x}{T} \qquad \ldots(i)$

Using Arrhenius equation

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

Taking log on both side

$$\log k = \log A - \frac{E_a}{2.303RT}$$
$$\log\left(\frac{k}{A}\right) = -\frac{E_a}{2.303RT} \qquad \ldots(ii)$$

Compare Eqs. (i) and (ii),

$\frac{x}{T} = \frac{E_a}{2.303RT} \Rightarrow E_a = 2.303R \cdot x$

Question8

Consider the reaction given below

$A + 2B \longrightarrow 3C + 2D$

If rate of disappearance of B is $x \times 10^{-2}$ mol L^{-1} s^{-1}, the ratio of rate of reaction and rate of appearance of C is

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Options:

A.

1 : 3

B.

3 : 1

C.

1 : 2

D.

2 : 1

Answer: A

Solution:

Given reaction is,

$A + 2B \rightarrow 3C + 2D$

The rate of reaction is, $r = -\frac{1}{2}\frac{d[B]}{dt}$

Given, $\frac{d[B]}{dt} = -x \times 10^{-2}$

Thus, $r = \frac{1}{2} \times x \times 10^{-2}$ $\qquad \ldots(i)$

Rate of appearance of C

$$= \frac{1}{3}\frac{d[C]}{dt} = -\frac{1}{2}\frac{[dB]}{dt}$$

$$\frac{d[C]}{dt} = \frac{3}{2} \times x \times 10^{-2} \qquad \ldots(ii)$$

Taking the ratio of (i) and (ii)

$\frac{\text{rate of reaction}}{\text{rate of appearance of } C} = 1 : 3$

Question9

Activation energy for the hydrolysis of sucrose by acid is X kJ mol^{-1} whereas activation energy for the hydrolysis of sucrose by sucrase is Y kJ mol^{-1}. X and Y respectively are

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Options:

A.

$6.22, 2.15$

B.

$2.15, 6.22$

C.

$6.22, 6.22$

D.

$2.15, 2.15$

Answer: A

Solution:

The activation energy for acid hydrolysis of sucrose is $6.22 \mathrm{~kJ} / \mathrm{mol}$. The activation energy for sucrose by sucrase is $2.15 \mathrm{~kJ} / \mathrm{mol}$.

Question10

$A \rightarrow$ products, is a first order reaction. The following data is obtained for this reaction at $T(\mathrm{~K})$. The value of $x : y$ is

Rate $\left(\mathrm{molL}^{-1} \mathrm{~min}^{-1}\right)$	$[A]$
0.2	0.02M
0.4	xM
1.0	yM

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Options:

A.

1 : 5

B.

2 : 3

C.

5 : 2

D.

2 : 5

Answer: D

Solution:

For 1st order rate is, Rate $= k[A]$

$\frac{\text{Rate}_1}{\text{Rate}_2} = \frac{[A_1]}{[A_2]}$

For first two data points,

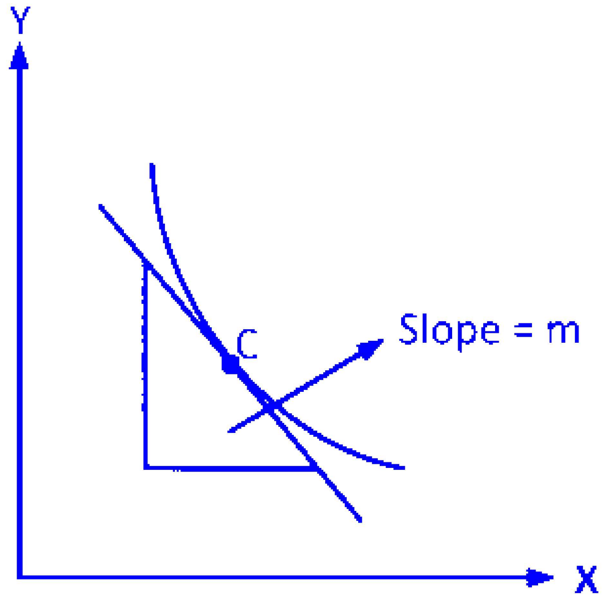
$$\begin{aligned} & \frac{0.2}{0.4} = \frac{0.02}{x} \\ \Rightarrow \quad & x = 0.04 \text{ m} \end{aligned}$$

For second and third data point,

$$\frac{0.4}{1} = \frac{0.04}{y} \Rightarrow y = 0.1\text{M}$$
$$x : y \Rightarrow 0.04 : 0.1 \Rightarrow 2 : 5$$

Question11

$A \rightarrow P$ is a first order reaction. The following graph is obtained for this reaction. (X-axis $=$ time: Y-axis $=$ conc. of A) The instantaneous rate of the reaction at point C is



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Options:

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

B. m

C. 2.303 m

D. $\frac{1}{2.303m}$

Answer: B

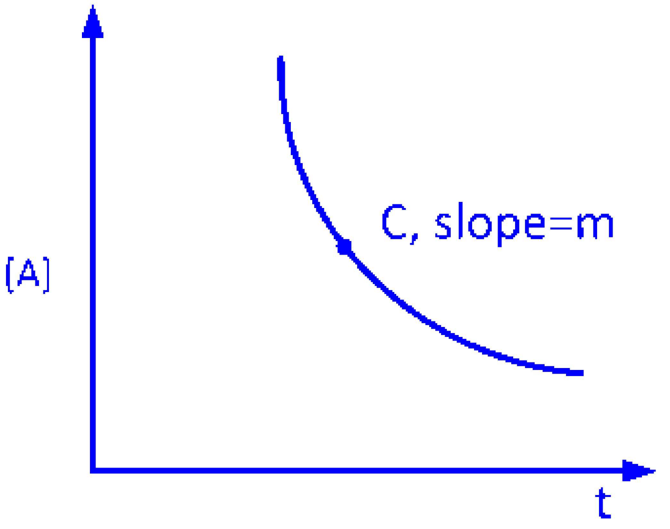
Solution:

The instaneous lst order reaction rate is given as

rate $= \frac{-d[A]}{dt}$ $\quad\quad\quad\quad \ldots(i)$

On X-axis $=$ time, Y-axis $=$ concentration

From graph slope $= -m$



i.e., concentration is decreasing with time.

$-m=\frac{d[A]}{dt}$ from graph $\quad\quad\quad\quad \ldots(ii)$

Equating (i) and (ii)

Rate of reaction $= m$

Question12

The rate constant of a first order reaction was doubled when the temperature was increased from 300 to 310 K . What is its approximate activation energy (in kJmol^{-1})? ($R=8.3 \mathrm{Jmol}^{-1} \mathrm{~K}^{-1}: \log 2=0.3$)

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Options:

A. 5.33

B. 533.3

C. 53333

D. 53.33

Answer: D

Solution:

To find the approximate activation energy (E_a) for a first-order reaction where the rate constant doubles when the temperature changes from 300 K to 310 K, we can use the following relationship:

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

For $T_1 = 300\,\mathrm{K}$, the rate constant is $K_1 = K$.

For $T_2 = 310\,\mathrm{K}$, the rate constant becomes $K_2 = 2K$.

By substituting the known values into the equation:

$\log\left(\frac{2K}{K}\right) = \frac{E_a}{2.303\times8.314}\times\frac{310-300}{310\times300}$

Simplifying and solving the equation:

$\log 2 = \frac{E_a}{2.303\times8.3}\times\frac{10}{310\times300}$

Given that $\log 2 = 0.3$, we rearrange to find E_a:

$E_a = 5333\,\mathrm{J/mol} = 53.33\,\mathrm{kJ/mol}$

Thus, the approximate activation energy is $53.33\,\mathrm{kJ/mol}$.

Question13

Isomerisation of gaseous cyclobutene to butadiene is first order reaction. At $T(\mathrm{~K})$. The rate constant of reaction is $33\times10^{-4}\mathrm{~s}^{-1}$. What is the time required (in min) to complete 90% of this reaction at the temperature? $(\log 2 = 03)$

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Options:

A. 116.67

B. 233.34

C. 58.34

D. 350.0

Answer: A

Solution:

For a first-order reaction, the integrated rate law is:

$k=\frac{1}{t} \ln \frac{[A]_0}{[A]_t}$ ...(i)

Where:

k is the rate constant,

$[A]_0$ is the initial concentration,

$[A]_t$ is the concentration at time t.

Given:

$k=3.3 \times 10^{-4} \mathrm{~s}^{-1}$,

$[A]_0=100$ (assuming initial concentration as 100 for calculation),

$[A]_t=10$ (since 90% of the reaction is complete, only 10% remains).

Substitute these values into equation (i):

$t=\frac{1}{k} \ln \frac{[A]_0}{[A]_t}=\frac{1}{3.3 \times 10^{-4}} \ln \left(\frac{100}{10}\right)$

This simplifies to:

$t=\frac{10^4 \times 2.303}{3.3}=6977.53 \mathrm{~s}$

Converting seconds to minutes:

$\frac{6977.5}{60}=116.67 \mathrm{~min}$

So, the time required to complete 90% of the reaction is 116.67 minutes.

Question14

$A \rightarrow P$ is a zero order reaction. At 298 K the rate constant of the reaction is $1 \times 10^{-3} \mathrm{~mol} \mathrm{~L}^{-1} \mathrm{~s}^{-1}$. Initial concentration of ' A ' is $0.1 \mathrm{~mol} \mathrm{~L}^{-1}$. What is the concentration of ' A after 10 sec ?

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Options:

A. $0.09 \mathrm{~mol} \mathrm{~L}^{-1}$

B. $0.099 \mathrm{~mol} \mathrm{~L}^{-1}$

C. $0.087 \mathrm{~mol} \mathrm{~L}^{-1}$

D. $0.011 \mathrm{~mol} \mathrm{~L}^{-1}$

Answer: A

Solution:

For a zero-order reaction, the rate of reaction is independent of the concentration of the reactant. Given the rate constant $k = 1 \times 10^{-3} \text{ mol L}^{-1}\text{s}^{-1}$, the initial concentration of A is $[A]_0 = 0.1 \text{ mol L}^{-1}$, and the time $t = 10$ seconds, we use the following formula to calculate the concentration of A after 10 seconds:

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

Substituting the known values into the formula, we get:

$1 \times 10^{-3} = \frac{0.1 - [A]_t}{10}$

Solving for $[A]_t$:

$0.01 = 0.1 - [A]_t$

$[A]_t = 0.1 - 0.01$

$[A]_t = 0.09 \text{ mol L}^{-1}$

Therefore, the concentration of A after 10 seconds is 0.09 mol L^{-1}.

Question15

The rate constant of a first order reaction is $3.46 \times 10^{-2} \text{ s}^{-1}$ at 298 K . What is the rate constant of the reaction at 350 K if its activation energy is 50.1 kJ mol^{-1}, $\left(R = 8.314 \text{JK}^{-1} \text{ mol}^{-1}\right)(\log 2 = 0.3010)$

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Options:

A. 0.592 s^{-1}

B. 0.692 s^{-1}

C. 0.792 s^{-1}

D. 0.892 s^{-1}

Answer: B

Solution:

To determine the rate constant of a reaction at a different temperature using the Arrhenius equation, follow this process:

The formula for calculating the change in the rate constant with temperature is:

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

Where:

K_{T_1} is the rate constant at temperature T_1.

K_{T_2} is the rate constant at temperature T_2.

E_a is the activation energy.

R is the universal gas constant.

Given Values:

Initial rate constant, $K_{T_1} = 3.46 \times 10^{-2}\,\mathrm{s}^{-1}$ at $T_1 = 298\,\mathrm{K}$

Activation energy, $E_a = 50.1\,\mathrm{kJ/mol} = 50.1 \times 10^3\,\mathrm{J/mol}$

Universal gas constant, $R = 8.314\,\mathrm{J/K} \cdot \mathrm{mol}$

Temperature, $T_2 = 350\,\mathrm{K}$

Calculation Steps:

Substitute the given values into the formula:

$\log \frac{K_{T_2}}{3.46 \times 10^{-2}} = \frac{50.1 \times 10^3}{2.303 \times 8.314} \times \left(\frac{350 - 298}{298 \times 350} \right)$

Evaluate the expression:

$\log \frac{K_{T_2}}{3\ 46 \times 10^{-2}} = 1.30$

Solve for K_{T_2}:

$\frac{K_{T_2}}{3.46 \times 10^{-2}} = 10^{1.30} = 19.95$

Calculate K_{T_2}:

$K_{T_2} = 3.46 \times 10^{-2} \times 19.95 = 0.6903\,\mathrm{s}^{-1}$

Thus, the rate constant at 350 K is approximately $0.692\,\mathrm{s}^{-1}$.

Question16

At 298 K the value of $-\frac{\Delta[\mathrm{Br}^-]}{\Delta t}$ for the reaction,

$$5\mathrm{Br}^-(aq) + \mathrm{BrO}_3^-(aq) + 6\mathrm{H}^+(aq) \longrightarrow 3\mathrm{Br}_2(aq) + 3\mathrm{H}_2\mathrm{O}(l)$$

is X molLmin^{-1}. What is the rate (in molL^{-1} min^{-1}) of this reaction?

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Options:

A. $5x$

B. x

C. $\frac{x}{5}$

D. $-\frac{x}{5}$

Answer: C

Solution:

For the given reaction:

$5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \longrightarrow 3Br_2(aq) + 3H_2O(l)$

The rate of change for the concentration of Br^- ions is given as:

$-\frac{\Delta[Br^-]}{\Delta t} = x \text{ mol L}^{-1}\text{min}^{-1}$

The overall rate of the reaction can be calculated by considering the stoichiometry of Br^-. Since 5 moles of Br^- are consumed per mole of reaction, the rate of the reaction is:

$\frac{\Delta R}{\Delta t} = -\frac{1}{5}\frac{\Delta[Br^-]}{\Delta t}$

$= -\frac{1}{5} \times (-x) = \frac{x}{5} \text{ mol L}^{-1}\text{min}^{-1}$

Therefore, the rate of the reaction is $\frac{x}{5}$ mol L^{-1}min^{-1}.

Question17

For a first order reaction the concentration of reactant was reduced from 0.03molL^{-1} to 0.02molL^{-1} in 25 min . What is its rate (in molL^{-1} s^{-1})?

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Options:

A. 6.667×10^{-6}

B. 4×10^{-4}

C. 6.667×10^{-4}

D. 4×10^{-6}

Answer: A

Solution:

The average rate of a first-order reaction is calculated using the formula:

$\frac{\Delta R}{\Delta t} = -\left[\frac{R_2 - R_1}{t_2 - t_1}\right]$

In this context:

R_1 is the initial concentration:

$R_1 = 0.03 \, \mathrm{mol/L}$

R_2 is the final concentration:

$R_2 = 0.02 \, \mathrm{mol/L}$

$t_2 - t_1$ is the time interval over which the change occurs, given as 25 minutes. To convert this time into seconds:

$t_2 - t_1 = 25 \times 60 \, \mathrm{seconds}$

By substituting these values into the rate formula, we obtain:

$\frac{\Delta R}{\Delta t} = -\frac{0.02 - 0.03}{25 \times 60}$

Calculating the above expression provides:

$\frac{\Delta R}{\Delta t} = 6.667 \times 10^{-6} \, \mathrm{mol/L/s}$

Question18

The first order reaction, $A(g) \rightarrow B(g) + 2C(g)$ occurs at $25^{\circ}\mathrm{C}$. After 24 minutes the ratio of the concentration of products to the concentration of the reactant is $1 : 3$ What is the half life is of the reaction (in min)? $\log 1.11 = 0.046$

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Options:

A. 150.5

B. 142.2

C. 157.8

D. 15.78

Answer: C

Solution:

For a first-order reaction,

$A(g) \rightarrow B(g) + 2C(g)$

After 24 minutes, the ratio of the concentration of the products to the reactant is $1 : 3$.

The rate constant k for a first-order reaction is given by the formula:

$k = \frac{2.303}{t} \log \left[\frac{a}{a-x}\right]$

Where:

$t = 24$ minutes

$a = 4$ (initial concentration)

$a - x = 3$

Substituting these into the equation, we have:

$k = \frac{2.303}{24} \log 1.11$

Using $\log 1.11 = 0.046$, we get:

$k = \frac{2.303}{24} \times 0.046 \approx 4.40 \times 10^{-3} \text{ min}^{-1}$

The half-life ($t_{1/2}$) for a first-order reaction is calculated using:

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

Substitute $k = 4.40 \times 10^{-3} \text{ min}^{-1}$:

$t_{1/2} = \frac{0.693}{4.40 \times 10^{-3}} \approx 157.8 \text{ minutes}$

Question19

Identify the autocatalytic reaction from the following

Options:

A. $N_2(g) + 3H_2(g) \xrightarrow[\text{Mo(s)}]{\text{Fe(s)}} 2NH_3(g)$

B. $2KClO_3(s) \xrightarrow{MnO_2} 2KCl(s) + 3O_2$

C. $CH_3COOC_2H_5 + H_2O \longrightarrow CH_3COOH + C_2H_5OH$

D. $AgNO_3(aq) + KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$

Answer: C

Solution:

A reaction is described as autocatalytic when one of its products also serves as a catalyst for the same reaction.

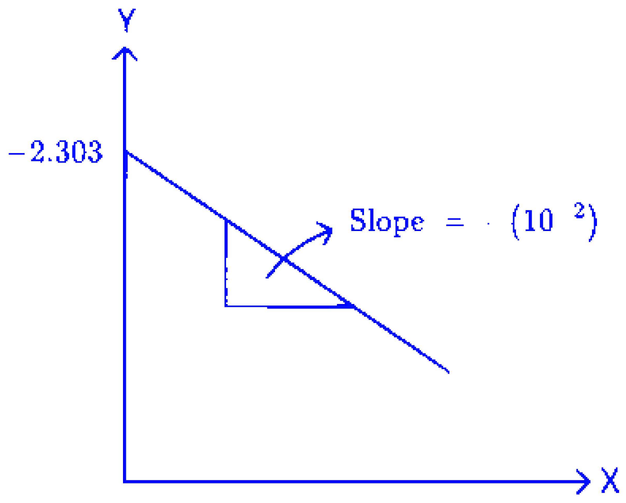
For the given reaction:

$$CH_3COOC_2H_5 + H_2O \xrightarrow{H^+} CH_3COOH + C_2H_5OH$$

Here, CH_3COOH is both a product and a catalyst for the reaction, illustrating the autocatalytic nature.

Question20

At 298 K , for a first order resction $(A \rightarrow P)$ the following graph is obtained. The rate constant (in s $^{-1}$) and initial concentration (in mol L^{-1}) of ' A ' are respectively (Y-axis $= \ln(a-x)$: X-axis $=$ time in sec)



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Options:

A. $2.303, 10^{-1}$

B. $10^{-2}; 2303$

C. $10^{-1}; 10^{-2}$

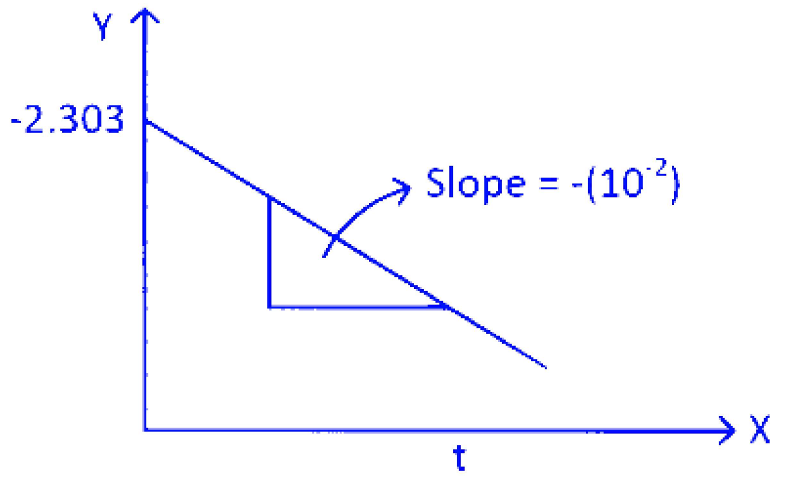
D. $10^{-2} : 10^{-1}$

Answer: D

Solution:

Given, first order reaction graph between $\ln(a-x)$ on Y-axis and t on X-axis.

According to first order reaction, $\ln(a-x)$



Rate constant is given by

$k=\frac{\ln }{t} \frac{[a]}{[a-x]} \quad \ldots$ (i)

$$\text{or}\begin{aligned} & k t=\ln \frac{[a]}{[a-x]} \\ & k t=\ln [a]-\ln [a-x]\end{aligned}$$

or $\ln [a-x]=\ln [a]-k t \quad \ldots$ (ii)

Compare Eq. (ii) by straight line equation.

$Y=\ln [a-x], x=t, m(\text { slope })=-k$
C (intercent $)=\ln$ al

C (intercept) $=\ln [a]$

Therefore, slope $=$ rate constant

$=-\left(-10^{-2}\right)=10^{-2} \quad \ldots$ (iii)

Now initial concentration (time $t=0$) Eq. (ii) becomes

$\ln [a-x]=\ln [a]$

From graph at $t=0, \ln [a-x]=-2.303$

$-2.303=\ln [a]$

Initial concentration, $[a]=10^{-1}$

Question21

The rate constant of a reaction at 500 K and 700 K are $0.02 \mathrm{~s}^{-1}$ and $0.2 \mathrm{~s}^{-1}$ respectively. The activation energy of the reaction (in kJmol^{-1}) is $\left(R=8.3 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}\right)$

Options:

A. 66.90

B. 33.45

C. 22.30

D. 44.45

Answer: B

Solution:

Given,

$k_1 = 0.02\ \mathrm{s}^{-1}$
$k_2 = 0.2\ \mathrm{s}^{-1}$
$E_d = ?$
$R = 8.3$
$T_1 = 500\ \mathrm{K}$
$T_2 = 700\ \mathrm{K}$

$$\text{In } \frac{k_2}{k_1} = -\frac{E_a}{R}\left(\frac{1}{r_2} - \frac{1}{T_1}\right)$$
$$\text{In } \frac{0.2}{0.02} = -\frac{E_a}{8.3}\left[\frac{1}{700} - \frac{1}{500}\right]$$
$$\text{In } 10 = \frac{-E_a}{8.3}\left(\frac{500-700}{700 \times 500}\right)$$
$$2.302 \times 8.3 \times 700 \times 500 = 200E_a$$
$$\Rightarrow E_a = \frac{2.302 \times 8.3 \times 700 \times 500}{200}$$
$$= 33436\ \mathrm{J/mol}$$
$$\approx 33.44\ \mathrm{kJ/mol}$$

Question22

The time required for completion of 93.75% of a first order reaction is x minutes. The half-life of it (in minutes) is

Options:

A. x/8

B. x/2

C. x/4

D. x/3

Answer: C

Solution:

Conc. at time, $t, A=A_0-\frac{93.75}{100} A_0=\frac{6.25}{100} A_0$

$A_0 \rightarrow$ initial conc. $k \rightarrow$ rate constant, time, $t=x$ min Using integrated rate law for lst order reaction,

$$\begin{aligned} k & =\frac{2.303}{t} \log \frac{A_0}{A} \\ & =\frac{2.303}{x} \log \frac{100 A_0}{6.25 A_0} \end{aligned}$$

$$\begin{aligned} \text{So,} \quad k & =\frac{2.303}{x} \log \frac{100}{6.25} \\ & =\frac{2.303 \times 1.20}{x} \\ \Rightarrow \quad k & =\frac{2.303 \times 1.20}{x} \min ^{-1} \end{aligned}$$

We know,

$$\begin{aligned} t_{1/2} & =\frac{0.693}{k} \\ & =\frac{0.693 x}{2.303 \times 1.20} \end{aligned}$$

So, $\quad t_{1/2}=\frac{x}{4} \min$

Question23

. The rate constant for a zero order reaction $A \longrightarrow$ products is $0.0030 \mathrm{~mol~L}^{-1} \mathrm{~S}^{-1}$. How long it will take for

the initial concentration of A to fall from 0.10 M to 0.075 M ?

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Options:

A. 10 s

B. 20 s

C. 8.33 s

D. 1.33 s

Answer: C

Solution:

Given, rate constant, $k_0 = 0.003 \mathrm{molL}^{-1} \mathrm{~s}^{-1}$

Initial conc., $A_0 = 0.1 \mathrm{M}$

Final conc., $A = 0.075 \mathrm{M}$

Using integrated rate law for zeroth order,

$$\begin{aligned} A_0 - A &= kt \\ 0.1 - 0.075 &= 0.003 \times t \\ t &= \frac{0.025}{0.003} = 8.33 \mathrm{~s} \end{aligned}$$

Question24

For a $A + B \rightarrow$ products, the rate of the reaction is given by rate $= k[A][B]^2$. The units of rate constant (k) will be

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Options:

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

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

C. $\mathrm{mol^2\ L^{-2}\ s^{-1}}$

D. $\mathrm{mol^{-2}\ L^{2}\ s^{-1}}$

Answer: D

Solution:

Given:

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

Step 1: Write dimensions of each term

- Rate of reaction has units:

  $\text{Rate} = \mathrm{mol\,L^{-1}\,s^{-1}}$

- $[A]$ and $[B]$ (concentrations) both have units:

  $\mathrm{mol\,L^{-1}}$

Step 2: Substitute into the rate equation

$\mathrm{mol\,L^{-1}\,s^{-1}} = k \times (\mathrm{mol\,L^{-1}}) \times (\mathrm{mol\,L^{-1}})^2$

Simplify right-hand side:

$k \times \mathrm{mol^3\,L^{-3}}$

Step 3: Solve for units of k

$k = \frac{\mathrm{mol\,L^{-1}\,s^{-1}}}{\mathrm{mol^3\,L^{-3}}}$

$k = \mathrm{mol^{-2}\,L^2\,s^{-1}}$

✅ **Final Answer:**

$\boxed{k = \mathrm{mol^{-2}\,L^2\,s^{-1}}}$

Correct option: D

Question25

For an elementary reaction, $X(g) \longrightarrow Y(g) + Z(g)$, the $t_{1/2}$ is 10 min. In what period of time would the

concentration of X be reduced to 10% of its original concentration?

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Options:

A. 20 min

B. 33.2 min

C. 15 min

D. 25.2 min

Answer: B

Solution:

$$\begin{aligned}
&X(g) \longrightarrow Y(g)+Z(g) \\
&t_{1/2}=10 \text{ min} \quad \text{(Given)} \\
&A=A_0 e^{-kt} \quad \text{....... (i)} \\
&[A_0= \text{ Initial concentration }] \\
&k=\frac{0.693}{t_{1/2}}=\frac{0.693}{10}=0.0693
\end{aligned}$$

From eq. (i),

$$\begin{aligned}
0.1 A_0 &= A_0 e^{-0.0693 \times t} \\
\ln 0.1 &= -0.0693 \times t \\
-2.303 &= -0.0693 t \Rightarrow t=33.23 \text{ min.}
\end{aligned}$$

Question26

Which statement among the following is incorrect?

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Options:

A. Unit of rate of reaction is ms^{-1}.

B. Unit of rate of disappearance is ms^{-1}.

C. Unit of rate constant k depends upon order of reaction.

D. Unit of rate constant k for a first order reaction is ms^{-1}.

Answer: D

Solution:

(b) Unit of rate of reaction $\frac{d[A]}{dt}=k[A]$

Unit $=\mathrm{ms}^{-1}$

(a) Rate of disappearance

$$A \rightarrow B$$
$$-\frac{d[A]}{dt} \Rightarrow \mathrm{ms}^{-1}$$

(d) For 1st order

$$\begin{aligned} \frac{-d[A]}{dt} &= k[A] \\ \mathrm{ms}^{-1} &= k\,\mathrm{m} \\ k &= \mathrm{s}^{-1} \end{aligned}$$

For 2nd order

$$\begin{aligned} \frac{-d[A]}{dt} &= k[A]^2 \\ \mathrm{ms}^{-1} &= k[\mathrm{~m}]^2 \\ k &= \mathrm{m}^{-1}\,\mathrm{s}^{-1} \end{aligned}$$

So, k depends on order of reaction and unit of k for 1st order reaction is s^{-1}.

Question27

For zero order reaction, a plot of $t_{1/2}$ versus$[A]_0$ will be

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Options:

A. a straight line passing through the origin and slope $=k$

B. a horizontal line (parallel to x-axis)

C. a straight line with slope $-k$

D. a straight line passing through origin and slope $= \frac{1}{2k}$

Answer: D

Solution:

For zero order reaction,

$[A] = [A_0] - kt$ ..... (i)

$[A] = \frac{[A_0]}{2}$ (half-life) ..... (ii)

From eqs. (i) and (ii)

$\frac{[A_0]}{2} = [A_0] - kt_{1/2} \Rightarrow t_{1/2} = \frac{[A]_0}{2k}$

Plot of $t_{1/2}$ versus $[A_0]$ will be straight line.

Question28

If the rate constant for a first order reaction is 2.303×10^{-3} s^{-1}. Find the time required to reduce 4 g of the reactant to 0.2 g.

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Options:

A. 1.30 hours

B. 21.60 hours

C. 0.36 hours

D. 2.60 hours

Answer: C

Solution:

Initial weight (A_0) or a is 4 g

Final weight at time t or after reduce $(a-x)$ or 0.2 g

First order reaction; $t=\frac{2.303}{k}\log\left(\frac{a}{a-x}\right)$

$$t=\frac{2.303}{2303\times10^{-3}}\log\left(\frac{49}{0.29}\right)$$
$$t=10^3\times\log 20$$
$$t=1301\text{ s or }t=0.36\text{ h}$$
