Chemical Kinetics

Chapter 4 Chemical Kinetics Exercise Solutions

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Q1 :

For the reaction R → P, the concentration of a reactant changes from 0.03 M to 0.02 M in 25 minutes. Calculate the average rate of reaction using units of time both in minutes and seconds.

Answer :

Average rate of reaction $= -\frac{\Delta[\mathrm{R}]}{\Delta t}$

$$= -\frac{[\mathrm{R}]_2 - [\mathrm{R}]_1}{t_2 - t_1}$$

$$= -\frac{0.02 - 0.03}{25}\ \mathrm{M\ min^{-1}}$$

$$= -\frac{-0.01}{25}\mathrm{M\ min^{-1}}$$

$$= 4 \times 10^{-4}\ \mathrm{M\ min^{-1}}$$

$$= \frac{4 \times 10^{-4}}{60}\ \mathrm{M\ s^{-1}}$$

$= 6.67 \times 10^{-6}\ \mathrm{M\ s^{-1}}$

Q2 :

In a reaction, 2A → Products, the concentration of A decreases from 0.5 mol $\mathrm{L^{-1}}$ to 0.4 mol $\mathrm{L^{-1}}$ in 10 minutes. Calculate the rate during this interval?

Answer :

Average rate $= -\frac{1}{2}\frac{\Delta[\mathrm{A}]}{\Delta t}$

$$= -\frac{1}{2}\frac{[A]_2 - [A]_1}{t_2 - t_1}$$

$$= -\frac{1}{2}\frac{0.4 - 0.5}{10}$$

$$= -\frac{1}{2}\frac{-0.1}{10}$$

= 0.005 mol L^{-1} min^{-1}

= 5 ×10^{-3}M min^{-1}

Q3 :

For a reaction, A + B Ã¢â€ ’Product; the rate law is given by, $r = k[A]^{1/2}[B]^2$. What is the order of the reaction?

Answer :

The order of the reaction $= \frac{1}{2} + 2$

$= 2\frac{1}{2}$

= 2.5

Q4 :

The conversion of molecules X to Y follows second order kinetics. If concentration of X is increased to three times how will it affect the rate of formation of Y?

Answer :

The reaction X → Y follows second order kinetics.

Therefore, the rate equation for this reaction will be:

Rate = $k[X]^2$ (1)

Let [X] = a mol L^{-1}, then equation (1) can be written as:

$Rate_1 = k \cdot (a)^2$

$= ka^2$

If the concentration of X is increased to three times, then [X] = $3a$ mol L^{-1}

Now, the rate equation will be:

Rate = $k\ (3a)^2$

= $9(ka^2)$

Hence, the rate of formation will increase by 9 times.

Q5 :

A first order reaction has a rate constant $1.15\ 10^{-3} s^{-1}$. How long will 5 g of this reactant take to reduce to 3 g?

Answer :

From the question, we can write down the following information:

Initial amount = 5 g

Final concentration = 3 g

Rate constant = $1.15\ 10^{-3} s^{-1}$

We know that for a 1st order reaction,

$$t = \frac{2.303}{k} \log \frac{[\mathrm{R}]_0}{[\mathrm{R}]}$$

$$= \frac{2.303}{1.15 \times 10^{-3}} \log \frac{5}{3}$$

$$= \frac{2.303}{1.15 \times 10^{-3}} \times 0.2219$$

= 444.38 s

= 444 s (approx)

Q6 :

Time required to decompose SO_2Cl_2 to half of its initial amount is 60 minutes. If the decomposition is a first order reaction, calculate the rate constant of the reaction.

Answer :

We know that for a 1st order reaction,

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

It is given that $t_{1/2}$ = 60 min

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

$$= \frac{0.693}{60}$$

$$= 0.01155 \text{ min}^{-1}$$
$$= 1.155 \text{ min}^{-1}$$

$$\text{Or } k = 1.925 \times 10^{-4} \text{ s}^{-1}$$

Q7 :

What will be the effect of temperature on rate constant?

Answer :

The rate constant of a reaction is nearly doubled with a 10°rise in temperature. However, the exact dependence of the rate of a chemical reaction on temperature is given by Arrhenius equation,

$$k = \text{Ae}^{-Ea/RT}$$

Where,

A is the Arrhenius factor or the frequency factor

$T$is the temperature

$R$is the gas constant

$E_a$is the activation energy

Q8 :

The rate of the chemical reaction doubles for an increase of 10 K in absolute temperature from 298 K. Calculate E_a.

Answer :

It is given that T_1 = 298 K

$\therefore T_2$ = (298 + 10) K

= 308 K

We also know that the rate of the reaction doubles when temperature is increased by 10°.

Therefore, let us take the value of $k_1 = k$ and that of $k_2 = 2k$

Also, $R = 8.314\ \text{J K}^{-1}\ \text{mol}^{-1}$

Now, substituting these values in the equation:

$$\log \frac{k_2}{k_1} = \frac{E_a}{2.303\ R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

We get:

$$\log \frac{2k}{k} = \frac{E_a}{2.303 \times 8.314}\left[\frac{10}{298 \times 308}\right]$$

$$\Rightarrow \log 2 = \frac{E_a}{2.303 \times 8.314}\left[\frac{10}{298 \times 308}\right]$$

$$\Rightarrow E_a = \frac{2.303 \times 8.314 \times 298 \times 308 \times \log 2}{10}$$

$= 52897.78\ \text{J mol}^{-1}$

$= 52.9\ \text{kJ mol}^{-1}$

Note: There is a slight variation in this answer and the one given in the NCERT textbook.

Q9 :

The activation energy for the reaction

$2HI_{(g)} \rightarrow H_2 + I_{2(g)}$

is 209.5 kJ mol^{-1} at 581K. Calculate the fraction of molecules of reactants having energy equal to or greater than activation energy?

Answer :

In the given case:

$E_a = 209.5\ \text{kJ mol}^{-1} = 209500\ \text{J mol}^{-1}$

$T = 581$ K

$R = 8.314\ \text{JK}^{-1}\ \text{mol}^{-1}$

Now, the fraction of molecules of reactants having energy equal to or greater than activation energy is given as:
$x = e^{-E_a/RT} \Rightarrow \text{In}\ x = -E$

Exercise : Solutions of Questions on Page Number : **117**

Q1 :

From the rate expression for the following reactions, determine their order of reaction and the dimensions of the rate constants.

(i) 3 NO(g) Ã¢â€ ’ N_2O(g) Rate = *k*[NO]^2

(ii) H_2O_2 (aq) + 3 I^-(aq) + 2 H^+Ã¢â€ ’ 2 H_2O (l) + I_3^- Rate = *k*[H_2O_2][I^-]

(iii) CH_3CHO(g) Ã¢â€ ’ CH_4(g) + CO(g) Rate = *k* [CH_3CHO]$^{3/2}$

(iv) C_2H_5Cl(g) Ã¢â€ ’ C_2H_4(g) + HCl(g) Rate = *k* [C_2H_5Cl]

Answer :

(i) Given rate = *k* [NO]^2

Therefore, order of the reaction = 2

Dimension of $k = \frac{\text{Rate}}{[\text{NO}]^2}$

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

(ii) Given rate = *k* [H_2O_2] [I^-]

Therefore, order of the reaction = 2

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

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

(iii) Given rate = *k* [CH_3CHO]$^{3/2}$

Therefore, order of reaction = $\frac{3}{2}$

Dimension of $k = \frac{\text{Rate}}{[\text{CH}_3\text{CHO}]^{\frac{3}{2}}}$

$$= \frac{\text{mol L}^{-1}\ \text{s}^{-1}}{\left(\text{mol L}^{-1}\right)^{\frac{3}{2}}}$$

$$= \frac{\text{mol L}^{-1}\ \text{s}^{-1}}{\text{mol}^{\frac{3}{2}}\ \text{L}^{-\frac{3}{2}}}$$

$$= \text{L}^{\frac{1}{2}}\ \text{mol}^{-\frac{1}{2}}\ \text{s}^{-1}$$

(iv) Given rate = $k\,[C_2H_5Cl]$

Therefore,order of the reaction = 1

Dimension of $k = \frac{\text{Rate}}{[C_2H_5Cl]}$

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

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

Q2 :

For the reaction:

$2A + B \rightarrow A_2B$

the rate = $k[A][B]^2$ with $k = 2.0 \times 10^{-6}\,\text{mol}^{-2}\text{L}^2\text{s}^{-1}$. Calculate the initial rate of the reaction when [A] = 0.1 mol L^{-1}, [B] = 0.2 mol L^{-1}. Calculate the rate of reaction after [A] is reduced to 0.06 mol L^{-1}.

Answer :

The initial rate of the reactionis

Rate = $k\,[A][B]^2$

$= (2.0 \times 10^{-6}\,\text{mol}^{-2}\text{L}^2\text{s}^{-1})\ (0.1\ \text{mol L}^{-1})\ (0.2\ \text{mol L}^{-1})^2$

$= 8.0 \times 10^{-9}\,\text{mol}^{-2}\text{L}^2\text{s}^{-1}$

When [A] is reduced from 0.1 mol $L^{-1}$to 0.06 mol^{-1}, the concentration of A reacted = (0.1 - 0.06) mol L^{-1} = 0.04 mol L^{-1}

Therefore, concentration of B reacted $= \frac{1}{2} \times 0.04\ \text{mol L}^{-1}$ = 0.02 mol L^{-1}

Then, concentration of B available, [B] = (0.2 - 0.02) mol L^{-1}

= 0.18 mol L^{-1}

After [A] is reduced to 0.06 mol L^{-1}, the rate of the reaction is given by,

Rate = $k\,[A][B]^2$

$= (2.0 \times 10^{-6}\,\text{mol}^{-2}\text{L}^2\text{s}^{-1})\ (0.06\ \text{mol L}^{-1})\ (0.18\ \text{mol L}^{-1})^2$

= 3.89 mol L^{-1}s^{-1}

Q3 :

The decomposition of NH_3 on platinum surface is zero order reaction. What are the rates of production of N_2 and H_2 if *k* = 2.5 x 10^{-4}mol^{-1}L s^{-1}?

Answer :

The decomposition of NH_3 on platinum surface is represented by the following equation.

$$2\,\mathrm{NH}_{3(g)} \xrightarrow{\mathrm{Pt}} \mathrm{N}_{2(g)} + 3\,\mathrm{H}_{2(g)}$$

Therefore,

$$\text{Rate} = -\frac{1}{2}\frac{d[\mathrm{NH}_3]}{dt} = \frac{d[\mathrm{N}_2]}{dt} = \frac{1}{3}\frac{d[\mathrm{H}_2]}{dt}$$

However,it is given that the reaction is of zero order.

Therefore,

$$-\frac{1}{2}\frac{d[\mathrm{NH}_3]}{dt} = \frac{d[\mathrm{N}_2]}{dt} = \frac{1}{3}\frac{d[\mathrm{H}_2]}{dt} = k$$
$$= 2.5\times10^{-4}\ \mathrm{mol\ L^{-1}\ s^{-1}}$$

Therefore, the rate of production of N_2 is

$$\frac{d[\mathrm{N}_2]}{dt} = 2.5\times10^{-4}\ \mathrm{mol\,L^{-1}s^{-1}}$$

And, the rate of production of H_2 is

$$\frac{d[\mathrm{H}_2]}{dt} = 3\times2.5\times10^{-4}\ \mathrm{mol\,L^{-1}s^{-1}}$$

= 7.5 × 10^{-4}mol L^{-1}s^{-1}

Q4 :

The decomposition of dimethyl ether leads to the formation of CH_4, H_2 and CO and the reaction rate is given by

Rate = *k* $[CH_3OCH_3]^{3/2}$

The rate of reaction is followed by increase in pressure in a closed vessel, so the rate can also be expressed in terms of the partial pressure of dimethyl ether, i.e.,

$$\text{Rate} = k\left(p_{\mathrm{CH_3OCH_3}}\right)^{3/2}$$

If the pressure is measured in bar andtime in minutes, then what are the units of rate and rate constants?

Answer :

If pressure is measured in bar and time in minutes, then

Unit of rate = bar min $^{-1}$

$$\text{Rate} = k\left(p_{\text{CH}_3\text{OCH}_3}\right)^{3/2}$$

$$\Rightarrow k = \frac{\text{Rate}}{\left(p_{\text{CH}_3\text{OCH}_3}\right)^{3/2}}$$

Therefore, unit of rate constants $(k) = \dfrac{\text{bar min}^{-1}}{\text{bar}^{3/2}}$

$$= \text{bar}^{-1/2} \ \ \text{min}^{-1}$$

Q5 :

Mention the factors that affect the rate of a chemical reaction.

Answer :

The factors that affect the rate of a reaction areas follows.

(i) Concentration of reactants (pressure in case of gases)

(ii) Temperature

(iii) Presence of a catalyst

Q6 :

A reaction is second order with respect to a reactant. How is the rate of reaction affected if the concentration of the reactant is

(i) doubled (ii) reduced to half?

Answer :

Letthe concentration of the reactant be [A] = *a*

Rate of reaction, R = $k\,[\text{A}]^2$

$= ka^2$

(i) If the concentration of the reactant is doubled, i.e. [A] = 2*a*, then the rate of the reaction would be

$R' = k(2a)^2$

$= 4ka^2$

$= 4R$

Therefore, the rate of the reaction would increase by 4 times.

(ii) If the concentration of the reactant is reduced to half, i.e. $[A] = \frac{1}{2}a$, then the rate of the reaction would be

$$R' = k\left(\frac{1}{2}a\right)^2$$
$$= \frac{1}{4}ka^2$$
$$= \frac{1}{4}R$$

Therefore, the rate of the reaction would be reduced to $\frac{1}{4}^{\text{th}}$.

Q7 :

What is the effect of temperature on the rate constant of a reaction? How can this temperature effect on rate constant be represented quantitatively?

Answer :

The rate constant is nearly doubled with a rise in temperature by 10° for a chemical reaction.

The temperature effect on the rate constant can be represented quantitatively by Arrhenius equation,

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

where, k is the rate constant,

A is the Arrhenius factor or the frequency factor,

R is the gas constant,

T is the temperature, and

E_a is the energy of activation for the reaction

Q8 :

In a pseudo first order hydrolysis of ester in water, the following results were obtained:

t/s	0	30	60	90
[Ester]mol L^{-1}	0.55	0.31	0.17	0.085

(i) Calculate the average rate of reaction between the time interval 30 to 60 seconds.

(ii) Calculate the pseudo first order rate constant for the hydrolysis of ester.

Answer :

(i) Average rate of reaction between the time interval, 30 to 60 seconds, $= \frac{d[\text{Ester}]}{dt}$

$$= \frac{0.31 - 0.17}{60 - 30}$$
$$= \frac{0.14}{30}$$

= 4.67 × 10^{-3}mol $L^{-1}s^{-1}$

(ii) For a pseudo first order reaction,

$$k = \frac{2.303}{t}\log\frac{[\text{R}]_0}{[\text{R}]}$$

For t= 30 s, $k_1 = \frac{2.303}{30}\log\frac{0.55}{0.31}$

= 1.911 × $10^{-2}s^{-1}$

For t= 60 s, $k_2 = \frac{2.303}{60}\log\frac{0.55}{0.17}$

= 1.957 × $10^{-2}s^{-1}$

For t= 90 s, $k_3 = \frac{2.303}{90}\log\frac{0.55}{0.085}$

= 2.075 × $10^{-2}s^{-1}$

Then, average rate constant, $k = \frac{k_1 + k_2 + k_3}{3}$

$$= \frac{\left(1.911\times10^{-2}\right) + \left(1.957\times10^{-2}\right) + \left(2.075\times10^{-2}\right)}{3}$$
$$= 1.98\times10^{-2}\ \text{s}^{-1}$$

Q9 :

A reaction is first order in A and second order in B.

(i) Write the differential rate equation.

(ii) How is the rate affected on increasing the concentration of B three times?

(iii) How is the rate affected when the concentrations of both A and B are doubled?

Answer :

(i) The differential rate equation will be

$$-\frac{d[\mathrm{R}]}{dt}=k[\mathrm{A}][\mathrm{B}]^2$$

(ii) If the concentration of B is increased three times, then

$$\begin{aligned}-\frac{d[\mathrm{R}]}{dt}&=k[\mathrm{A}][3\mathrm{B}]^2\\&=9\cdot k[\mathrm{A}][\mathrm{B}]^2\end{aligned}$$

Therefore, the rate of reaction will increase 9 times.

(iii) When the concentrations of both A and B are doubled,

$$\begin{aligned}-\frac{d[\mathrm{R}]}{dt}&=k[\mathrm{A}][\mathrm{B}]^2\\&=k[2\mathrm{A}][2\mathrm{B}]^2\\&=8\cdot k[\mathrm{A}][\mathrm{B}]^2\end{aligned}$$

Therefore, the rate of reaction will increase 8 times.

Q10 :

In a reaction between A and B, the initial rate of reaction (r_0) was measured for different initial concentrations of A and B as given below:

A/ mol L^{-1}	0.20	0.20	0.40
B/ mol L^{-1}	0.30	0.10	0.05
r_0/ mol L^{-1} s^{-1}	5.07×10^{-5}	5.07×10^{-5}	1.43×10^{-4}

What is the order of the reaction with respect to A and B?

Answer :

Let the order of the reaction with respect to A be *x*and with respect to B be *y*.

Therefore,

$$r_0 = k[A]^x[B]^y$$

$$5.07\times10^{-5} = k[0.20]^x[0.30]^y \qquad (i)$$

$$5.07\times10^{-5} = k[0.20]^x[0.10]^y \qquad (ii)$$

$$1.43\times10^{-4} = k[0.40]^x[0.05]^y \qquad (iii)$$

Dividing equation (i) by (ii), we obtain

$$\frac{5.07\times10^{-5}}{5.07\times10^{-5}} = \frac{k[0.20]^x[0.30]^y}{k[0.20]^x[0.10]^y}$$

$$\Rightarrow 1 = \frac{[0.30]^y}{[0.10]^y}$$

$$\Rightarrow \left(\frac{0.30}{0.10}\right)^0 = \left(\frac{0.30}{0.10}\right)^y$$

$$\Rightarrow y = 0$$

Dividing equation (iii) by (ii), we obtain

$$\frac{1.43\times10^{-4}}{5.07\times10^{-5}} = \frac{k[0.40]^x[0.05]^y}{k[0.20]^x[0.30]^y}$$

$$\Rightarrow \frac{1.43\times10^{-4}}{5.07\times10^{-5}} = \frac{[0.40]^x}{[0.20]^x} \qquad \begin{bmatrix}\text{Since } y = 0,\\ [0.05]^y = [0.30]^y = 1\end{bmatrix}$$

$$\Rightarrow 2.821 = 2^x$$

$$\Rightarrow \log 2.821 = x\log 2 \qquad (\text{Taking log on both sides})$$

$$\Rightarrow x = \frac{\log 2.821}{\log 2}$$

= 1.496

= 1.5 (approximately)

Hence, the order of the reaction with respect to A is 1.5 and with respect to B is zero.

Q11 :

The following results have been obtained during the kinetic studies of the reaction:

2A + B Ã¢â€ ’ C + D

Experiment	A/ mol L^{-1}	B/ mol L^{-1}	Initial rate of formation of D/mol L^{-1} min^{-1}
I	0.1	0.1	6.0×10^{-3}

II	0.3	0.2	7.2×10^{-2}
III	0.3	0.4	2.88×10^{-1}
IV	0.4	0.1	2.40×10^{-2}

Determine the rate law and the rate constant for the reaction.

Answer :

Let the order of the reaction with respect to A be *x*and with respect to B be *y*.

Therefore, rate of the reaction is given by,

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

According to the question,

$$6.0\times10^{-3} = k[0.1]^x[0.1]^y \qquad \text{(i)}$$

$$7.2\times10^{-2} = k[0.3]^x[0.2]^y \qquad \text{(ii)}$$

$$2.88\times10^{-1} = k[0.3]^x[0.4]^y \qquad \text{(iii)}$$

$$2.40\times10^{-2} = k[0.4]^x[0.1]^y \qquad \text{(iv)}$$

Dividing equation (iv) by (i), we obtain

$$\frac{2.40\times10^{-2}}{6.0\times10^{-3}} = \frac{k[0.4]^x[0.1]^y}{k[0.1]^x[0.1]^y}$$

$$\Rightarrow 4 = \frac{[0.4]^x}{[0.1]^x}$$

$$\Rightarrow 4 = \left(\frac{0.4}{0.1}\right)^x$$

$$\Rightarrow (4)^1 = 4^x$$

$$\Rightarrow x = 1$$

Dividing equation (iii) by (ii), we obtain

Therefore, the rate law is

Rate = k [A] $[B]^2$

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

From experiment I, we obtain

$$k = \frac{6.0 \times 10^{-3} \text{ mol L}^{-1} \text{ min}^{-1}}{\left(0.1 \text{ mol L}^{-1}\right)\left(0.1 \text{ mol L}^{-1}\right)^2}$$

= 6.0 $L^2 mol^{-2} min^{-1}$

From experiment II, weobtain

$$k = \frac{7.2 \times 10^{-2} \text{ mol L}^{-1} \text{ min}^{-1}}{\left(0.3 \text{ mol L}^{-1}\right)\left(0.2 \text{ mol L}^{-1}\right)^2}$$

= 6.0 $L^2 mol^{-2} min^{-1}$

From experiment III, we obtain

$$k = \frac{2.88 \times 10^{-1} \text{ mol L}^{-1} \text{ min}^{-1}}{\left(0.3 \text{ mol L}^{-1}\right)\left(0.4 \text{ mol L}^{-1}\right)^2}$$

= 6.0 $L^2 mol^{-2} min^{-1}$

From experiment IV, we obtain

$$k = \frac{2.40 \times 10^{-2} \text{ mol L}^{-1} \text{ min}^{-1}}{\left(0.4 \text{ mol L}^{-1}\right)\left(0.1 \text{ mol L}^{-1}\right)^2}$$

= 6.0 $L^2 mol^{-2} min^{-1}$

Therefore, rate constant, k = 6.0 $L^2 mol^{-2} min^{-1}$

Q12 :

The reaction between A and B is first order with respect to A and zero order with respect to B. Fill in the blanks in the following table:

Experiment	A/ mol L^{-1}	B/ mol L^{-1}	Initial rate/mol L^{-1} min^{-1}
I	0.1	0.1	2.0×10^{-2}
II	--	0.2	4.0×10^{-2}
III	0.4	0.4	--
IV	--	0.2	2.0×10^{-2}

Answer :

The given reaction is of the first order with respect to A and of zero order with respect to B.

Therefore, the rate of the reaction is given by,

Rate = $k\,[A]^1[B]^0$

⇒ Rate = $k\,[A]$

From experiment I, we obtain

$2.0 \times 10^{-2}\,\text{mol L}^{-1}\text{min}^{-1} = k\,(0.1\,\text{mol L}^{-1})$

⇒ $k = 0.2\,\text{min}^{-1}$

From experiment II, we obtain

$4.0 \times 10^{-2}\,\text{mol L}^{-1}\text{min}^{-1} = 0.2\,\text{min}^{-1}[A]$

⇒ $[A] = 0.2\,\text{mol L}^{-1}$

From experiment III, we obtain

Rate = $0.2\,\text{min}^{-1} \times 0.4\,\text{mol L}^{-1}$

$= 0.08\,\text{mol L}^{-1}\text{min}^{-1}$

From experiment IV, we obtain

$2.0 \times 10^{-2}\,\text{mol L}^{-1}\text{min}^{-1} = 0.2\,\text{min}^{-1}[A]$

⇒ $[A] = 0.1\,\text{mol L}^{-1}$

Q13 :

Calculate the half-life of a first order reaction from their rate constants given below:

(i) 200 s^{-1} (ii) 2 min^{-1} (iii) 4 years^{-1}

Answer :

(i) Half life, $t_{1/2} = \frac{0.693}{k}$

$= \frac{0.693}{200\ \text{s}^{-1}}$

$= 3.47 \times 10^{-3}$ s (approximately)

(ii) Half life, $t_{1/2} = \frac{0.693}{k}$

$= \frac{0.693}{2\ \text{min}^{-1}}$

= 0.35 min (approximately)

(iii) Half life, $t_{1/2} = \frac{0.693}{k}$

$$= \frac{0.693}{4 \text{ years}^{-1}}$$

= 0.173 years (approximately)

Q14 :

The half-life for radioactive decay of ^{14}C is 5730 years. An archaeological artifact containing wood had only 80% of the ^{14}C found in a living tree. Estimate the age of the sample.

Answer :

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

$$= \frac{0.693}{5730} \text{ years}^{-1}$$

It is known that,

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

$$= \frac{2.303}{\frac{0.693}{5730}} \log \frac{100}{80}$$

= 1845 years (approximately)

Hence, the age of the sample is 1845 years.

Q15 :

The experimental data for decomposition of N_2O_5

$$\left[2N_2O_5 \longrightarrow 4NO_2 + O_2\right]$$

in gas phase at 318K are given below:

$t(s)$	0	400	800	1200	1600	2000	2400	2800	3200
$10^2 \times [N_2O_5] \text{ mol L}^{-1}$	1.63	1.36	1.14	0.93	0.78	0.64	0.53	0.43	0.35

(i) Plot $[N_2O_5]$ against *t*.

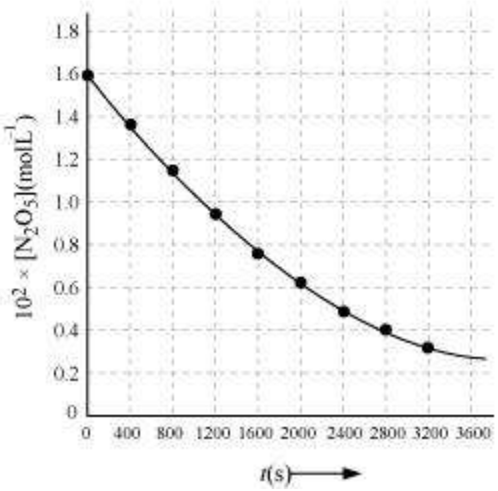
(ii) Find the half-life period for the reaction.

(iii) Draw a graph between log [N_2O_5] and *t*.

(iv) What is the rate law?

(v)

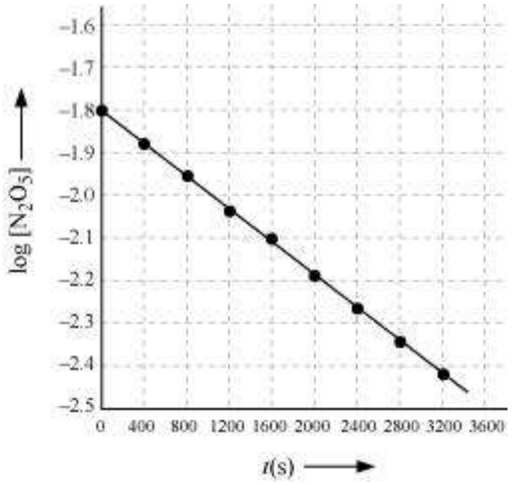
Answer :



(ii) Time corresponding to the concentration, $\frac{1.630\times10^{2}}{2}\,\mathrm{mol\,L^{-1}} = 81.5\,\mathrm{mol\,L^{-1}},$ is the half life. From the graph, the half life is obtained as 1450 s.

(iii)

t(s)	$\mathbf{10^2 \times [N_2O_5] / mol\,L^{-1}}$	$\mathbf{\log[N_2O_5]}$
0	1.63	- 1.79
400	1.36	- 1.87
800	1.14	- 1.94
1200	0.93	- 2.03
1600	0.78	- 2.11
2000	0.64	- 2.19
2400	0.53	- 2.28
2800	0.43	- 2.37
3200	0.35	- 2.46



(iv) The given reaction is of the first order as the plot, $\log[N_2O_5]$ v/s *t*, is a straight line. Therefore, the rate law of the reaction is

$$\text{Rate} = k[N_2O_5]$$

(v) From the plot, $\log[N_2O_5]$ v/s *t*, we obtain

$$\text{Slope} = \frac{-2.46-(-1.79)}{3200-0}$$
$$= \frac{-0.67}{3200}$$

Again, slope of the line of the plot $\log[N_2O_5]$ v/s *t* is given by

$$-\frac{k}{2.303}$$.

Therefore, we obtain,

$$-\frac{k}{2.303} = -\frac{0.67}{3200}$$

Q16 :

The rate constant for a first order reaction is 60 s^{-1}. How much time will it take to reduce the initial concentration of the reactant to its 1/16thvalue?

Answer :

It is known that,

$$
\begin{aligned}
t &= \frac{2.303}{k}\log\frac{[\mathrm{R}]_0}{[\mathrm{R}]}\\
&= \frac{2.303}{60\ \mathrm{s}^{-1}}\log\frac{1}{1/16}\\
&= \frac{2.303}{60\ \mathrm{s}^{-1}}\log 16\\
&= 4.6\times10^{-2}\ \mathrm{s}\ (\text{approximately})
\end{aligned}
$$

Hence, the required time is 4.6×10^{-2}s.

Q17 :

During nuclear explosion, one of the products is ^{90}Sr with half-life of 28.1 years. If 1ÃŽÂ¼g of ^{90}Sr was absorbed in the bones of a newly born baby instead of calcium, how much of it will remain after 10 years and 60 years if it is not lost metabolically.

Answer :

Here, $k = \frac{0.693}{t_{1/2}} = \frac{0.693}{28.1}\ \mathrm{y}^{-1}$

It is known that,

$$t = \frac{2.303}{k}\log\frac{[\mathrm{R}]_0}{[\mathrm{R}]}$$

$$
\begin{aligned}
&\Rightarrow 10 = \frac{2.303}{\frac{0.693}{28.1}}\ \log\frac{1}{[\mathrm{R}]}\\
&\Rightarrow 10 = \frac{2.303}{\frac{0.693}{28.1}}\left(-\log[\mathrm{R}]\right)\\
&\Rightarrow \log[\mathrm{R}] = -\frac{10\times0.693}{2.303\times28.1}\\
&\Rightarrow [\mathrm{R}] = \text{antilog}\ (-0.1071)\\
&\qquad\quad = \text{antilog}\left(\bar{1}.8929\right)\\
&\qquad\quad = 0.7814\mu\text{g}
\end{aligned}
$$

Therefore, 0.7814 ÃŽÂ¼g of ^{90}Sr will remain after 10 years.

Again,

$$t=\frac{2.303}{k}\log\frac{[\mathrm{R}]_0}{[\mathrm{R}]}$$

$$\begin{aligned}&\Rightarrow 60=\frac{2.303}{\frac{0.693}{28.1}}\ \log\frac{1}{[\mathrm{R}]}\\&\Rightarrow \log[\mathrm{R}]=-\frac{60\times0.693}{2.303\times28.1}\\&\Rightarrow[\mathrm{R}]=\text{antilog }(-0.6425)\\&\qquad\quad=\text{antilog}(\bar{1}.3575)\\&\qquad\quad=0.2278\mu\mathrm{g}\end{aligned}$$

Therefore, 0.2278 ÃŽÂ¼g of ^{90}Sr will remain after 60 years.

Q18 :

For a first order reaction, show that time required for 99% completion is twice the time required for the completion of 90% of reaction.

Answer :

For a first order reaction, the time required for 99% completionis

$$\begin{aligned}t_1&=\frac{2.303}{k}\log\frac{100}{100-99}\\&=\frac{2.303}{k}\log100\\&=2\times\frac{2.303}{k}\end{aligned}$$

For a first order reaction, the time required for 90% completion is

$$\begin{aligned}t_2&=\frac{2.303}{k}\log\frac{100}{100-90}\\&=\frac{2.303}{k}\log10\\&=\frac{2.303}{k}\end{aligned}$$

Therefore, $t_1= 2t_2$

Hence, the time required for 99% completion of a first order reaction is twice the time required for the completion of 90% of the reaction.

Q19 :

A first order reaction takes 40 min for 30% decomposition. Calculate $t_{1/2}$.

Answer :

For a first order reaction,

$$t = \frac{2.303}{k}\log\frac{[\mathrm{R}]_0}{[\mathrm{R}]}$$

$$\begin{aligned} k &= \frac{2.303}{40\ \text{min}}\log\frac{100}{100-30} \\ &= \frac{2.303}{40\ \text{min}}\log\frac{10}{7} \\ &= 8.918\times10^{-3}\ \text{min}^{-1} \end{aligned}$$

Therefore, $t_{1/2}$ of the decomposition reaction is

$$\begin{aligned} t_{1/2} &= \frac{0.693}{k} \\ &= \frac{0.693}{8.918\times10^{-3}}\ \text{min} \end{aligned}$$

= 77.7 min (approximately)

Q20 :

For the decomposition of azoisopropane to hexane and nitrogen at 543 K, the following data are obtained.

t (sec)	P(mm of Hg)
0	35.0
360	54.0
720	63.0

Calculate the rate constant.

Answer :

The decomposition of azoisopropane to hexane and nitrogen at 543 K is represented by the following equation.

$$\begin{array}{lcccc} & (\mathrm{CH_3})_2\mathrm{CHN{=}NCH}(\mathrm{CH_3})_{2(g)} & \longrightarrow & \mathrm{N}_{2(g)} & +\ \mathrm{C_6H}_{14(g)} \\ \text{At } t=0 & \mathrm{P}_0 & & 0 & 0 \\ \text{At } t=t & \mathrm{P}_0-p & & p & p \end{array}$$

After time, *t*, total pressure, $P_t = (P_0 - p) + p + p$

$$\Rightarrow P_t = P_0 + p$$

$$\Rightarrow p = P_t - P_0$$

$$\text{Therefore, } P_o - p = P_o - (P_t - P_o)$$

$= 2P_0 - P_t$

For a first order reaction,

$$k = \frac{2.303}{t}\log\frac{P_0}{P_0 - p}$$

$$= \frac{2.303}{t}\log\frac{P_0}{2P_0 - P_t}$$

When $t = 360$ s, $k = \frac{2.303}{360\text{ s}}\log\frac{35.0}{2\times 35.0 - 54.0}$

$= 2.175 \times 10^{-3}\ s^{-1}$

When $t = 720$ s, $k = \frac{2.303}{720\text{ s}}\log\frac{35.0}{2\times 35.0 - 63.0}$

$= 2.235 \times 10^{-3}\ s^{-1}$

Hence, the average value of rate constant is

$$k = \frac{(2.175\times 10^{-3}) + (2.235\times 10^{-3})}{2}\,s^{-1}$$

$= 2.21 \times 10^{-3}\ s^{-1}$

Note: There is a slight variation in this answer and the one given in the NCERT textbook.

Q21 :

The following data were obtained during the first order thermal decomposition of SO_2Cl_2at a constant volume.

$$SO_2Cl_2(g) \longrightarrow SO_2(g) + Cl_2(g)$$

Experiment	Time/s^{-1}	Total pressure/atm
1	0	0.5
2	100	0.6

Calculate the rate of the reaction when total pressure is 0.65 atm.

Answer :

The thermal decomposition of SO_2Cl_2at a constant volume is represented by the following equation.

$$\begin{array}{llccc} & SO_2Cl_{2(g)} & \longrightarrow & SO_{2(g)} & + \; Cl_{2(g)} \\ \text{At } t=0 & P_0 & & 0 & 0 \\ \text{At } t=t & P_0 - p & & p & p \end{array}$$

After time, *t*, total pressure, $P_t = (P_0 - p) + p + p$

$$\Rightarrow P_t = P_0 + p$$
$$\Rightarrow p = P_t - P_0$$

Therefore, $P_o - p = P_o - (P_t - P_o)$

$= 2\,P_0 - P_t$

For a first order reaction,

$$k = \frac{2.303}{t}\log\frac{P_0}{P_0 - p}$$
$$= \frac{2.303}{t}\log\frac{P_0}{2P_0 - P_t}$$

When *t*= 100 s, $k = \frac{2.303}{100\text{ s}}\log\frac{0.5}{2\times 0.5 - 0.6}$

$= 2.231 \times 10^{-3} s^{-1}$

When P_t= 0.65 atm,

$P_0 + p = 0.65$

$\Rightarrow p = 0.65 - P_0$

= 0.65 - 0.5

= 0.15 atm

Therefore, when the total pressure is 0.65 atm, pressure of $SOCl_2$is

$p_{SOCl_2} = P_0 - p$

= 0.5 - 0.15

= 0.35 atm

Therefore, the rate of equation, when total pressure is 0.65 atm, is given by,

Rate = $k(p_{SOCl_2})$

$= (2.23 \times 10^{-3} s^{-1})\,(0.35\text{ atm})$

$= 7.8 \times 10^{-4}\text{atm } s^{-1}$

Q22 :

The rate constant for the decomposition of N_2O_5 at various temperatures is given below:

T/°C	0	20	40	60	80
$10^5 \times k / s^{-1}$	0.0787	1.70	25.7	178	2140

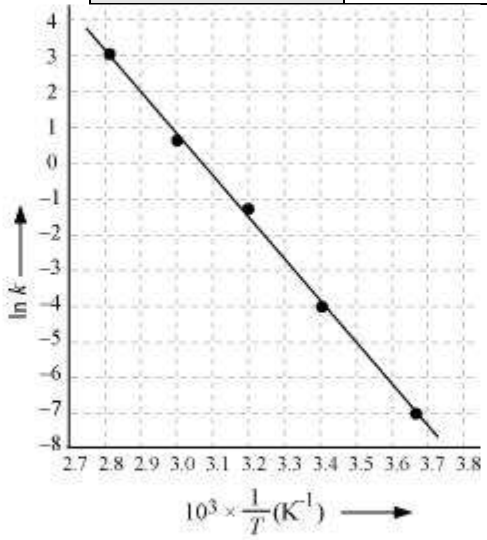
Draw a graph between ln *k* and 1/*T* and calculate the values of *A* and E_a.

Predict the rate constant at 30 º and 50 ºC.

Answer :

From the given data, we obtain

T/°C	0	20	40	60	80
T/K	273	293	313	333	353
$\frac{1}{T}$ / K^{-1}	3.66×10^{-3}	3.41×10^{-3}	3.19×10^{-3}	3.0×10^{-3}	2.83×10^{-3}
$10^5 \times k / s^{-1}$	0.0787	1.70	25.7	178	2140
ln k	- 7.147	- 4.075	- 1.359	- 0.577	3.063



Slope of the line,

$$\frac{y_2 - y_1}{x_2 - x_1} = -12.301\,\mathrm{K}$$

According to Arrhenius equation,

$$\text{Slope} = -\frac{E_a}{\mathrm{R}}$$
$$\Rightarrow E_a = -\text{Slope} \times \mathrm{R}$$
$$= -(-12.301\,\mathrm{K}) \times (8.314\ \mathrm{J\,K^{-1}mol^{-1}})$$
$$= 102.27\ \mathrm{kJ\,mol^{-1}}$$

Again,

$$\ln k = \ln A - \frac{E_a}{\mathrm{R}T}$$
$$\ln A = \ln k + \frac{E_a}{\mathrm{R}T}$$

When $T = 273\,\mathrm{K}$,

$$\ln k = -7.147$$
$$\text{Then, } \ln A = -7.147 + \frac{102.27 \times 10^3}{8.314 \times 273}$$
$$= 37.911$$

$$\text{Therefore, } A = 2.91 \times 10^6$$

When $T = 30 + 273\,\mathrm{K} = 303\,\mathrm{K}$,

$$\frac{1}{T} = 0.0033\,\mathrm{K} = 3.3 \times 10^{-3}\,\mathrm{K}$$

Then, $\text{at } \frac{1}{T} = 3.3 \times 10^{-3}\,\mathrm{K},$

$$\ln k = -2.8$$
$$\text{Therefore, } k = 6.08 \times 10^{-2}\,\mathrm{s^{-1}}$$

Again, when $T = 50 + 273\,\mathrm{K} = 323\,\mathrm{K}$,

Q23 :

The rate constant for the decomposition of hydrocarbons is 2.418 x 10^{-5} s^{-1}at 546 K. If the energy of activation is 179.9 kJ/mol, what will be the value of pre-exponential factor.

Answer :

k= 2.418 × $10^{-5}$$s^{-1}$

T= 546 K

E_a= 179.9 kJ mol^{-1}= 179.9 × 10^3J mol^{-1}

According to the Arrhenius equation,

$$k = \mathrm{A}e^{-E_a/\mathrm{R}T}$$

$$\Rightarrow \ln k = \ln \mathrm{A} - \frac{E_a}{\mathrm{R}T}$$

$$\Rightarrow \log k = \log \mathrm{A} - \frac{E_a}{2.303\ \mathrm{R}T}$$

$$\Rightarrow \log \mathrm{A} = \log k + \frac{E_a}{2.303\ \mathrm{R}T}$$

$$= \log\left(2.418 \times 10^{-5}\ \mathrm{s}^{-1}\right) + \frac{179.9 \times 10^3\ \mathrm{J\ mol^{-1}}}{2.303 \times 8.314\ \mathrm{Jk^{-1}\ mol^{-1}} \times 546\ \mathrm{K}}$$

= (0.3835 - 5) + 17.2082

= 12.5917

Therefore, A = antilog (12.5917)

= 3.9 × $10^{12}$$s^{-1}$(approximately)

Q24 :

Consider a certain reaction A → Products with *k* = 2.0 x 10^{-2} s^{-1}. Calculate the concentration of *A* remaining after 100 s if the initial concentration of *A* is 1.0 mol L^{-1}.

Answer :

k= 2.0 × $10^{-2}$$s^{-1}$

T= 100 s

$[A]_o$= 1.0 moL^{-1}

Sincethe unit of *k*is s^{-1}, the given reaction is a first order reaction.

Therefore, $k = \frac{2.303}{t} \log \frac{[\mathrm{A}]_0}{[\mathrm{A}]}$

$$\Rightarrow 2.0\times10^{-2}\ \text{s}^{-1} = \frac{2.303}{100\ \text{s}}\log\frac{1.0}{[\text{A}]}$$

$$\Rightarrow 2.0\times10^{-2}\ \text{s}^{-1} = \frac{2.303}{100\,\text{s}}\left(-\log[\text{A}]\right)$$

$$\Rightarrow -\log[\text{A}] = \frac{2.0\times10^{-2}\times100}{2.303}$$

$$\Rightarrow [\text{A}] = \text{anti}\log\left(-\frac{2.0\times10^{-2}\times100}{2.303}\right)$$

= 0.135 mol L^{-1}(approximately)

Hence, the remaining concentration of A is 0.135 mol L^{-1}.

Q25 :

Sucrose decomposes in acid solution into glucose and fructose according to the first order rate law, with$t_{1/2}$ = 3.00 hours. What fraction of sample of sucrose remains after 8 hours?

Answer :

For a first order reaction,

$$k = \frac{2.303}{t}\log\frac{[\text{R}]_0}{[\text{R}]}$$

It is given that, $t_{1/2}$= 3.00 hours

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

$$= \frac{0.693}{3}\ \text{h}^{-1}$$

= 0.231 h^{-1}

Then, 0.231 h^{-1} $= \frac{2.303}{8\ \text{h}}\log\frac{[\text{R}]_0}{[\text{R}]}$

$$\Rightarrow \log\frac{[R]_0}{[R]} = \frac{0.231\ \text{h}^{-1} \times 8\ \text{h}}{2.303}$$

$$\Rightarrow \frac{[R]_0}{[R]} = \text{antilog}(0.8024)$$

$$\Rightarrow \frac{[R]_0}{[R]} = 6.3445$$

$$\Rightarrow \frac{[R]}{[R]_0} = 0.1576\ \text{(approx)}$$

$$= 0.158$$

Hence, the fraction of sample of sucrose that remains after 8 hours is 0.158.

Q26 :

The decomposition of hydrocarbon follows the equation

$k = (4.5 \times 10^{11}\ \text{s}^{-1})\ e^{-28000\ \text{K}/T}$

Calculate E_a.

Answer :

The given equation is

$k = (4.5 \times 10^{11}\ \text{s}^{-1})\ e^{-28000\ \text{K}/T}$ (i)

Arrhenius equation is given by,

$k = \text{A}e^{-E_a/RT}$ (ii)

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

$$\frac{E_a}{RT} = \frac{28000\,\text{K}}{T}$$

$$\Rightarrow E_a = \text{R} \times 28000\ \text{K}$$

$= 8.314\ \text{J K}^{-1}\text{mol}^{-1} \times 28000\ \text{K}$

$= 232792\ \text{J mol}^{-1}$

$= 232.792\ \text{kJ mol}^{-1}$

Q27 :

The rate constant for the first order decomposition of H_2O_2 is given by the following equation:

$\log k = 14.34 - 1.25 \times 10^4\ \text{K}/T$

Calculate E_a for this reaction and at what temperature will its half-period be 256 minutes?

Answer :

Arrhenius equation is given by,

$$k = \mathrm{A}e^{-E_a/\mathrm{R}T}$$

$$\Rightarrow \ln k = \ln \mathrm{A} - \frac{E_a}{\mathrm{R}T}$$

$$\Rightarrow \ln k = \log \mathrm{A} - \frac{E_a}{\mathrm{R}T}$$

$$\Rightarrow \log k = \log \mathrm{A} - \frac{E_a}{2.303\ \mathrm{R}T} \qquad \text{(i)}$$

The given equation is

$$\log k = 14.34 - 1.25 \times 10^4\ \mathrm{K}/T \qquad \text{(ii)}$$

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

$$\frac{E_a}{2.303\ \mathrm{R}T} = \frac{1.25 \times 10^4\ \mathrm{K}}{T}$$

$$\Rightarrow E_a = 1.25 \times 10^4 \mathrm{K} \times 2.303 \times \mathrm{R}$$

= 1.25×10^4K × 2.303 × 8.314 J K^{-1}mol^{-1}

= 239339.3 J mol^{-}1 (approximately)

= 239.34 kJ mol^{-1}

Also, when $t_{1/2}$= 256 minutes,

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

$$= \frac{0.693}{256}$$

= 2.707×10^{-3}min^{-1}

= 4.51×10^{-5}s^{-1}

It is also given that, log k= 14.34 - 1.25 × 10^4K/T

$$\Rightarrow \log\left(4.51\times10^{-5}\right)=14.34-\frac{1.25\times10^{4}\ \text{K}}{T}$$

$$\Rightarrow \log\left(0.654-05\right)=14.34-\frac{1.25\times10^{4}\ \text{K}}{T}$$

$$\Rightarrow \frac{1.25\times10^{4}\ \text{K}}{T}=18.686$$

$$\Rightarrow T=\frac{1.25\times10^{4}\ \text{K}}{18.686}$$

= 668.95 K

= 669 K (approximately)

Q28 :

The decomposition of A into product has value of *k* as 4.5 x 10^3 s^{-1} at 10°C and energy of activation 60 kJ mol^{-1}. At what temperature would *k* be 1.5 x 10^4 s^{-1}?

Answer :

From Arrhenius equation, we obtain

$$\log\frac{k_2}{k_1}=\frac{E_a}{2.303\,\text{R}}\left(\frac{T_2-T_1}{T_1T_2}\right)$$

Also, $k_1 = 4.5 \times 10^3\ s^{-1}$

$T_1 = 273 + 10 = 283$ K

$k_2 = 1.5 \times 10^4\ s^{-1}$

$E_a = 60$ kJ mol^{-1} = 6.0×10^4 J mol^{-1}

Then,

$$\log\frac{1.5\times10^{4}}{4.5\times10^{3}}=\frac{6.0\times10^{4}\ \text{J mol}^{-1}}{2.303\times8.314\ \text{J K}^{-1}\ \text{mol}^{-1}}\left(\frac{T_2-283}{283T_2}\right)$$

$$\Rightarrow 0.5229=3133.627\left(\frac{T_2-283}{283T_2}\right)$$

$$\Rightarrow \frac{0.5229\times283\ T_2}{3133.627}=T_2-283$$

$$\Rightarrow 0.0472T_2=T_2-283$$

$$\Rightarrow 0.9528\ T_2=283$$

$$\Rightarrow T_2=297.019\ \text{K (approximately)}$$

= 297 K

= 24°C

Hence, *k* would be $1.5 \times 10^4 \text{ s}^{-1}$ at 24°C.

Note: There is a slight variation in this answer and the one given in the NCERT textbook.

Q29 :

The time required for 10% completion of a first order reaction at 298 K is equal to that required for its 25% completion at 308 K. If the value of *A* is $4 \times 10^{10} \text{ s}^{-1}$. Calculate *k* at 318 K and E_a.

Answer :

For a first order reaction,

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

At 298 K, $t = \frac{2.303}{k}\log\frac{100}{90}$

$$= \frac{0.1054}{k}$$

At 308 K, $t' = \frac{2.303}{k'}\log\frac{100}{75}$

$$= \frac{2.2877}{k'}$$

According to the question,

$$t = t'$$

$$\Rightarrow \frac{0.1054}{k} = \frac{0.2877}{k'}$$

$$\Rightarrow \frac{k'}{k} = 2.7296$$

From Arrhenius equation, we obtain

$$\log\frac{k'}{k}=\frac{E_a}{2.303\,\mathrm{R}}\left(\frac{T'-T}{TT'}\right)$$

$$\log(2.7296)=\frac{E_a}{2.303\times 8.314}\left(\frac{308-298}{298\times 308}\right)$$

$$E_a=\frac{2.303\times 8.314\times 298\times 308\times\log(2.7296)}{308-298}$$

$$=76640.096\ \mathrm{J\,mol^{-1}}$$

$$=76.64\ \mathrm{kJ\,mol^{-1}}$$

To calculate *k* at 318 K,

It is given that, $A=4\times 10^{10}\,\mathrm{s^{-1}},\ T=318\,\mathrm{K}$

Again, from Arrhenius equation, we obtain

$$\log k=\log A-\frac{E_a}{2.303\,\mathrm{R}T}$$

$$=\log\left(4\times 10^{10}\right)-\frac{76.64\times 10^{3}}{2.303\times 8.314\times 318}$$

$$=(0.6021+10)-12.5876$$

$$=-1.9855$$

$$\text{Therefore, } k=\text{Antilog}(-1.9855)$$

$$=1.034\times 10^{-2}\,\mathrm{s^{-1}}$$

Q30 :

The rate of a reaction quadruples when the temperature changes from 293 K to 313 K. Calculate the energy of activation of the reaction assuming that it does not change with temperature.

Answer :

From Arrhenius equation, we obtain

$$\log\frac{k_2}{k_1}=\frac{E_a}{2.303\,\mathrm{R}}\left(\frac{T_2-T_1}{T_1T_2}\right)$$

It is given that, $k_2 = 4k_1$

$T_1 = 293\,\mathrm{K}$

$T_2 = 313\,\mathrm{K}$

Therefore, $\log\frac{4k_1}{k_2}=\frac{E_a}{2.303\times 8.314}\left(\frac{313-293}{293\times 313}\right)$

$$\Rightarrow 0.6021=\frac{20\times E_a}{2.303\times 8.314\times 293\times 313}$$

$$\Rightarrow E_a=\frac{0.6021\times 2.303\times 8.314\times 293\times 313}{20}$$

$$=52863.33\ \mathrm{J\,mol^{-1}}$$

$$=52.86\ \mathrm{kJ\,mol^{-1}}$$

Hence, the required energy of activation is 52.86 $\mathrm{kJmol^{-1}}$.