# NCERT Solutions for Class 12 Chemistry Part 1 Chapter 4

## Chemical Kinetics Class 12

### Chapter 4 Chemical Kinetics Exercise Solutions

**In text :** Solutions of Questions on Page Number : **98**

**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 $= -\frac{\Delta[R]}{\Delta t}$

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

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

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

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

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

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

**Q2 :**

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

**Answer :**

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

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

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

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

= 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:

$Rate_1 = k\,.(a)^2$

$= ka^2$

If the concentration of X is increased to three times, then [X] = 3*a* 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,

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

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

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

= 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 $1^{st}$ 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,

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

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

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

**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 kJ $mol^{-1}$ = 209500 J $mol^{-1}$

$T$ = 581 K

$R$ = 8.314 $JK^{-1}$ $mol^{-1}$ Now, the fraction of molecules of reactants having energy equal to or greater than activation energy is given as:

$x = e^{-E_a/RT} \Rightarrow \ln x = -E$

<< Previous Chapter 3 : ElectrochemistryNext Chapter 5 : Surface Chemistry >>

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

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

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

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

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

Therefore, order of the reaction = 2

Dimension of $k = \frac{\text{Rate}}{[H_2O_2][I^-]}$

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

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

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

Dimension of $k = \frac{\text{Rate}}{[CH_3CHO]^{\frac{3}{2}}}$

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

**(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]$^2$

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

= $8.0 \times 10^{-9}$ mol$^{-2}$L$^2$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]$^2$

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

$$2\,NH_{3(g)} \xrightarrow{Pt} N_{2(g)} + 3\,H_{2(g)}$$

Therefore,

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

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

$$\frac{d[\mathrm{H_2}]}{dt}=3\times2.5\times10^{-4}\ \mathrm{mol\,L^{-1}s^{-1}}$$

= $7.5 \times 10^{-4}$ 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.,**

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

**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 $min^{-1}$

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

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

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

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

**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 = $k\,[A]^2$

$= ka^2$

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

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

$= 4ka^2$

= 4 R

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

$$R' = k(\frac{1}{2}a)^2$$

$$= \frac{1}{4}ka^2$$

$$= \frac{1}{4}R$$

..................................

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

**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:**

**Answer :**

| 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.**

**(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,

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

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, $k = \frac{k_1 + k_2 + k_3}{3}$

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

**Q9 :**

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**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

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

Therefore, the rate of reaction will increase 9 times.

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

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

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 $L^{-1}$ $s^{-1}$ | $5.07 \times 10^{-5}$ | $5.07 \times 10^{-5}$ | $1.43 \times 10^{-4}$ |

**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

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

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

$$\Rightarrow 2.821 = 2^x$$

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

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

= 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 | $6.0 \times 10^{-3}$ |
| II | 0.3 | 0.2 | $7.2 \times 10^{-2}$ |
| III | 0.3 | 0.4 | $2.88 \times 10^{-1}$ |
| IV | 0.4 | 0.1 | $2.40 \times 10^{-2}$ |

**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]$^2$

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

From experiment I, we obtain

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

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

From experiment II, weobtain

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

= 6.0 $L_2 mol_{-2} min_{-1}$

From experiment III, we obtain

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

= 6.0 $L_2 mol_{-2} min_{-1}$

From experiment IV, we obtain

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

= 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 | $2.0 \times 10^{-2}$ |

| | | | |
|---|---|---|---|
| II | -- | 0.2 | $4.0 \times 10^{-2}$ |
| III | 0.4 | 0.4 | -- |
| IV | -- | 0.2 | $2.0 \times 10^{-2}$ |

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

$\Rightarrow$ Rate = $k$ [A]

From experiment I, we obtain

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

$0.2\ min^{-1}$

From experiment II, we obtain

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

$\Rightarrow [A] = 0.2\ mol\ L^{-1}$

From experiment III, we obtain

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

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

From experiment IV, we obtain

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

$\Rightarrow [A] = 0.1\ mol\ L^{-1}$

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

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

= 3.47 ×10 $^{-3}$ s (approximately)

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

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

= 0.35 min (approximately)

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

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

= 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 |

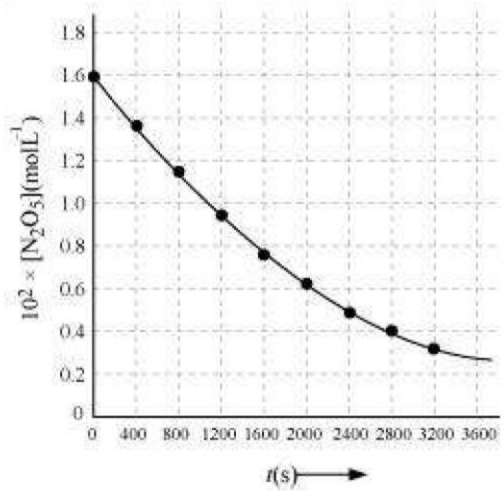
**(i) Plot [$N_2O_5$] against *t*.**
**(ii) Find the half-life period for the reaction.**

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

**(iv) What is the rate law?**

**(v)**

**Answer :**

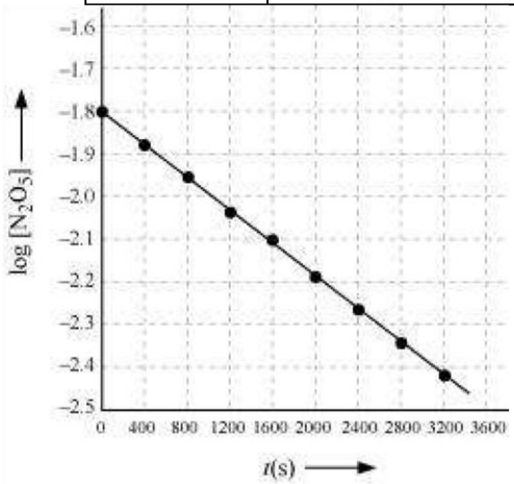



**(ii)** Time corresponding to the concentration, $\frac{1.630\times10^{2}}{2}\,\text{mol}\,\text{L}^{-1} = 81.5\,\text{mol}\,\text{L}^{-1}$, is the half life. From the graph, the half life is obtained as 1450 s.

**(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 1/16$^{th}$ value?**

**Answer :**

It is known that,

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

$$= \frac{2.303}{60\ s^{-1}}\log\frac{1}{1/16}$$

$$= \frac{2.303}{60\ s^{-1}}\log 16$$

$$= 4.6\times10^{-2}\ s\ (\text{approximately})$$

Hence, the required time is $4.6 \times 10^{-2}$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, $k = \frac{0.693}{t_{1/2}} = \frac{0.693}{28.1}\ y^{-1}$

It is known that,

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

$$\Rightarrow 10 = \frac{2.303}{\frac{0.693}{28.1}}\log\frac{1}{[R]}$$
$$\Rightarrow 10 = \frac{2.303}{\frac{0.693}{28.1}}(-\log[R])$$
$$\Rightarrow \log[R] = -\frac{10\times 0.693}{2.303\times 28.1}$$
$$\Rightarrow [R] = \text{antilog}\ (-0.1071)$$
$$= \text{antilog}(\bar{1}.8929)$$
$$= 0.7814\mu g$$

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

Again,

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

$$\Rightarrow 60 = \frac{2.303}{\frac{0.693}{28.1}}\log\frac{1}{[R]}$$
$$\Rightarrow \log[R] = -\frac{60\times 0.693}{2.303\times 28.1}$$
$$\Rightarrow [R] = \text{antilog}\ (-0.6425)$$
$$= \text{antilog}(\bar{1}.3575)$$
$$= 0.2278\mu g$$

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

$$t_1 = \frac{2.303}{k}\log\frac{100}{100-99}$$
$$= \frac{2.303}{k}\log 100$$
$$= 2\times\frac{2.303}{k}$$

$$t_2 = \frac{2.303}{k}\log\frac{100}{100-90}$$
$$= \frac{2.303}{k}\log 10$$
$$= \frac{2.303}{k}$$

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

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,

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

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

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

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

= 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.

$$(CH_3)_2CHN{=}NCH(CH_3)_{2(g)} \longrightarrow N_{2(g)} + C_6H_{14(g)}$$

| | | | |
|---|---|---|---|
| At $t = 0$ | $P_0$ | 0 | 0 |
| At $t = t$ | $P_0 - p$ | $p$ | $p$ |

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

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

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

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

**Q21 :**

**Answer :**

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**The following data were obtained during the first order thermal decomposition of $SO_2Cl_2$ at 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.**

The thermal decomposition of $SO_2Cl_2$ at a constant volume is represented by the following equation.

| | $SO_2Cl_{2(g)}$ | $\longrightarrow$ | $SO_{2(g)}$ | + | $Cl_{2(g)}$ |
|---|---|---|---|---|---|
| At $t = 0$ | $P_0$ | | 0 | | 0 |
| At $t = t$ | $P_0 - p$ | | p | | p |

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, $P_0 - p = P_0 - (P_t - P_0)$

$= 2\ P_0 - P_t$

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

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

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 atm) =
$7.8 \times 10^{-4}$ 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 |

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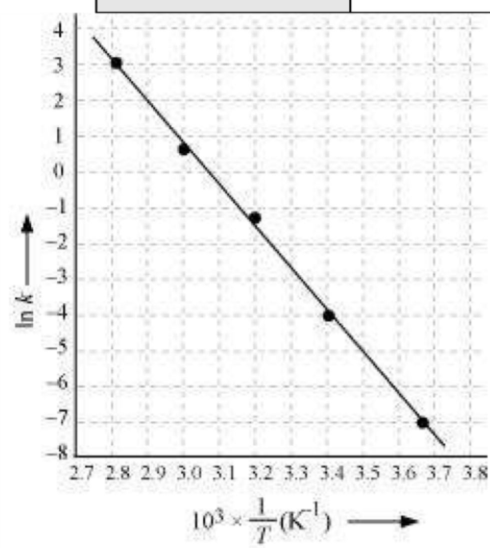
**Draw a graph between ln *k* and 1/*T* and calculate the values of *A* and $E_a$.**

**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 | |
| $\frac{1}{T}$ / $K^{-1}$ | $3.66\times10^{-3}$ | $3.41\times10^{-3}$ | $3.19\times10^{-3}$ | $3.0\times10^{-3}$ | $2.83\times10^{-3}$ | |
| $10^5 \times k/s^{-1}$ | 0.0787 | 1.70 | 25.7 | 178 | 2140 | |
| ln *k* | - 7.147 | - 4.075 | - 1.359 | - 0.577 | 3.063 | |




Slope of the line,

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

According to Arrhenius equation,

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

Again,

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

When $T = 273\,\text{K}$,
$\ln k = -7.147$

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

Therefore, $A = 2.91 \times 10^6$

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

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

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

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

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

**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 × $10^3$J $mol^{-1}$

According to the Arrhenius equation,

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

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

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

$$\Rightarrow \log A = \log k + \frac{E_a}{2.303\ RT}$$

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

= (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 mo$L^{-1}$

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

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

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

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

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

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

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

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

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

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

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

$$= 0.158$$

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

**Q26 :**

**The decomposition of hydrocarbon follows the equation**

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

**$E_a$.**

**Answer :**

The given equation is

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

Arrhenius equation is given by,

$k = Ae^{-E_a/RT}$ (ii)

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

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

$$\Rightarrow E_a = R \times 28000\ K$$

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

$= 232792\ J\ mol^{-1} =$
$232.792\ kJ\ mol^{-1}$

**Q27 :**

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

**log $k$ = 14.34 - 1.25 x $10^4$ 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

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

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

= $1.25 \times 10^4$K × 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,

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

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

= $2.707 \times 10^{-3} min^{-1}$

= $4.51 \times 10^{-5} s^{-1}$

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

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

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

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

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

= 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

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

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 $mol^{-1} = 6.0 \times 10^{4}$ 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 $1.5\times10^{4}\ s^{-1}$ 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 x $10^{10}\ s^{-1}$. 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, $A=4\times10^{10}\,\mathrm{s^{-1}},\ T=318\,\mathrm{K}$

Again, from Arrhenius equation, we obtain

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

$$=\log\left(4\times10^{10}\right)-\frac{76.64\times10^{3}}{2.303\times8.314\times318}$$

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

$$=-1.9855$$

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

$$=1.034\times10^{-2}\,\mathrm{s^{-1}}$$

**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 $\mathrm{kJmol^{-1}}$.