

NCERT Solution for Class 12

Chemistry

Chapter 3 - Chemical Kinetics

Intext Exercise

1. For the reaction $R \rightarrow P$, 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.

Ans: The average rate of the reaction can be calculated by dividing the change in the rate of decreasing the rate of reactant by the time taken. This is given below:

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

This can be written as:

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

$$R_1 = 0.03 \text{ M}$$

$$R_2 = 0.02 \text{ M}$$

$$t_2 - t_1 = 25 \text{ min}$$

Putting the values, we get:

$$\text{Average rate} = - \frac{0.02 - 0.03}{25}$$

$$\text{Average rate} = - \frac{-0.01}{25} = 4 \times 10^{-4} \text{ M min}^{-1}$$

So, the average rate in minutes will be $4 \times 10^{-4} \text{ M min}^{-1}$

Now, to find the average rate in seconds we have to divide the above answer by 60.

So, the answer will be:

$$\text{Average rate} = - \frac{4 \times 10^{-4}}{60} = 6.66 \times 10^{-6} \text{ M s}^{-1}$$

Therefore, the average rate in seconds will be $6.66 \times 10^{-6} \text{ M s}^{-1}$

2. In a reaction, $2A \rightarrow \text{Products}$, 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.

Ans: The average rate of the reaction can be calculated by dividing the change in the rate of decreasing the rate of reactant by the time taken. This is given below:

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

This is due the fact that the reaction given is:



So, the average rate will be written as:

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

$$A_1 = 0.5 \text{ M}$$

$$A_2 = 0.4 \text{ M}$$

$$t_2 - t_1 = 10 \text{ min}$$

Putting the values, we get:

$$\text{Average rate} = - \frac{1}{2} \times \frac{0.4 - 0.5}{10}$$

$$\text{Average rate} = - \frac{1}{2} \times \frac{-0.1}{10} = 5 \times 10^{-3} \text{ M min}^{-1}$$

So, the average rate will be $5 \times 10^{-3} \text{ M min}^{-1}$

3. For a reaction, $A + B \rightarrow \text{Product}$, the rate law is given by: $r = k [A]^{1/2} [B]^2$.

What is the order of the reaction?

Ans: The order of the reaction can be calculated by adding the stoichiometry coefficients of the reactants in the given rate of the reaction.

Given the rate is $r = k [A]^{1/2} [B]^2$

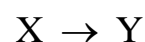
So, the order will be:

$$\text{Order} = 2 + \frac{1}{2} = 2.5$$

So, the order of the reaction is 2.5

4. The conversion of the 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?

Ans: The reaction will be:



As the question says that this reaction follows the second order reaction, we can write the rate law equation as:

$$\text{Rate} = k[X]^2 = ka^2$$

$$\text{If } [X] = a \text{ mol/L}$$

It is said that the concentration of X increases by three times, so we can write:

$$[X] = 3a \text{ mol/L}$$

Therefore, the rate of reaction will be:

$$\text{Rate} = k(3a)^2 = 9ka^2$$

Thus, the rate of the reaction will increase by 9 times or the rate formation will increase by 9 times.

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

Ans: The initial amount of the reactant is given as 5 g. We can write:

$$[R]_0 = 5 \text{ g}$$

The final amount of the reactant is given as 3 g. We can write:

$$[R] = 3 \text{ g}$$

We are also given the value of rate constant as:

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

We know that the reaction is a 1st order reaction, the time can be calculated by:

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

Putting the values in the above formula, we get:

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

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

$$t = 444 \text{ seconds}$$

So, the time taken will be 444 seconds.

6. Time required to decompose SO_2Cl_2 to half its initial amount is 60 minutes.

If the decomposition is a first order reaction, calculate the rate constant of the reaction.

Ans: We are given that the decomposition of SO_2Cl_2 is a first order reaction. So, we can write:

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

It is mentioned that the time required for the initial amount to become half is 60 minutes.

$$t_{1/2} = 60 \text{ min}$$

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

Putting the value, we get:

$$k = \frac{0.693}{60 \times 60} = 1.925 \times 10^{-4} \text{ s}^{-1}$$

Thus, the rate constant is $1.925 \times 10^{-4} \text{ s}^{-1}$.

7. What will be the effect of temperature on rate constant?

Ans: An increase of 10 degrees in temperature causes a reaction's rate constant to almost double in size. In any case, the Arrhenius equation gives the exact temperature dependency of a chemical reaction rate.

The Arrhenius equation is given below:

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

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.

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

Ans: The formula that can be used to solve the question is:

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

T_1 temperature given is absolute temperature and it is equal to 298 K.

$$T_2 = 298 + 10 = 308 \text{ K}$$

It is given that the rate of the chemical reaction double for an increase of 10 K, therefore, we can write the values of rate constant as:

$$k_1 = x$$

$$k_2 = 2x$$

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

Now, putting all the values in the formula, we get:

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

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

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

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

$$E_a = 52.89 \text{ kJ mol}^{-1}$$

9. The activation energy for the reaction $2\text{HI}_{(g)} \rightarrow \text{H}_{2(g)} + \text{I}_{2(g)}$ is $209.5 \text{ kJ mol}^{-1}$ at 581 K. Calculate the fraction of molecules of molecules having energy equal to or greater than activation energy.

Ans: We are given the activation energy as $209.5 \text{ kJ mol}^{-1}$

$$T = 581 \text{ K}$$

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

Now, the fraction of molecules of reactants having energy equal to or greater than activation energy is given as:

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

$$\ln x = \frac{-E_a}{RT}$$

$$\log x = \frac{-E_a}{2.303 RT}$$

$$\log x = \frac{209500}{2.303 \times 8.314 \times 581} = 18.8323$$

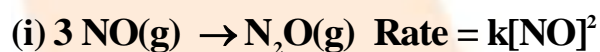
Now, taking the antilog:

$$x = \text{Antilog}(18.8323)$$

$$x = 1.471 \times 10^{-19}$$

NCERT Exercise

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



Ans: We are given:

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

From this we can see that the order of the reaction = 2

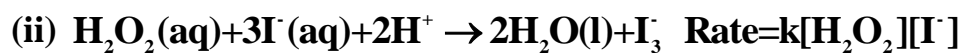
$$k = \frac{\text{rate}}{[\text{No}]^2}$$

Dimensions will be:

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

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

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



Ans: Rate = $k[\text{H}_2\text{O}_2][\text{I}^-]$

From this we can see that the order of the reaction = 2

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

Dimensions will be:

$$\begin{aligned} k &= \frac{\text{mol L}^{-1} \text{s}^{-1}}{(\text{mol L}^{-1})^2} \\ &= \frac{\text{mol L}^{-1} \text{s}^{-1}}{\text{mol}^2 \text{L}^2} \\ &= \text{L mol}^{-1} \text{s}^{-1} \end{aligned}$$



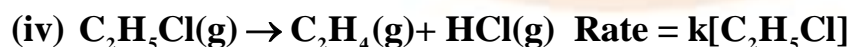
Ans: Rate = $k[\text{CH}_3\text{CHO}]^{3/2}$

From this we can see that the order of the reaction = $\frac{3}{2}$

$$k = \frac{\text{rate}}{[\text{CH}_3\text{CHO}]^{3/2}}$$

Dimensions will be:

$$\begin{aligned} k &= \frac{\text{mol L}^{-1} \text{s}^{-1}}{(\text{mol L}^{-1})^{3/2}} \\ &= \frac{\text{mol L}^{-1} \text{s}^{-1}}{\text{mol}^{3/2} \text{L}^{3/2}} \\ &= \text{L}^{1/2} \text{mol}^{-1/2} \text{s}^{-1} \end{aligned}$$



Ans: Rate = $k[\text{C}_2\text{H}_5\text{Cl}]$

From this we can see that the order of the reaction = 1

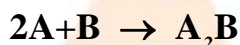
$$k = \frac{\text{rate}}{[\text{C}_2\text{H}_5\text{Cl}]}$$

Dimensions will be:

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

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

2. For a reaction:



The rate = $k[\text{A}][\text{B}]^2$ with $k = 2.0 \times 10^{-6} \text{ mol}^{-2} \text{ L}^2 \text{ s}^{-1}$.

Calculate the initial rate of the reaction when $[\text{A}] = 0.1 \text{ mol L}^{-1}$, $[\text{B}] = 0.2 \text{ mol L}^{-1}$. Calculate the rate of reaction after $[\text{A}]$ is reduced to 0.06 mol L^{-1} .

Ans: We are given the rate of the reaction as:

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

Putting the values in this, we get the rate as:

$$\text{rate} = 2.0 \times 10^{-6} \times 0.1 \times (0.2)^2$$

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

When $[\text{A}]$ is reduced from 0.10 mol L^{-1} to 0.06 mol L^{-1} . So, the amount of $[\text{A}]$ reacted will be:

$$= 0.10 - 0.06 = 0.04 \text{ mol L}^{-1}$$

Therefore, the concentration of B reacted will be:

$$= \frac{1}{2} \times 0.04 = 0.02 \text{ mol L}^{-1}$$

Hence, new $[\text{B}] = 0.2 - 0.02 = 0.18 \text{ mol L}^{-1}$

Now, the new rate of the reaction will be:

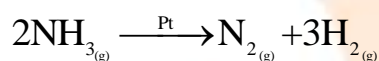
$$\text{rate} = 2.0 \times 10^{-6} \times 0.06 \times (0.18)^2$$

$$\text{rate} = 3.89 \times 10^{-9} \text{ mol L}^{-1} \text{ s}^{-1}$$

Therefore, the rate of the reaction is $3.89 \times 10^{-9} \text{ mol L}^{-1} \text{ s}^{-1}$.

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

Ans: The following equation represents the breakdown of Ammonia on the platinum surface.



Therefore, we can write the rate of the reaction as:

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

But we are given that the reaction is a zero order reaction.

$$\text{So, Rate} = -\frac{1}{2} \frac{d[\text{NH}_3]}{dt} = \frac{d[\text{N}_2]}{dt} = \frac{1}{3} \frac{d[\text{H}_2]}{dt} = 2.5 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

Thus, the rate of production of N_2 will be:

$$\frac{d[\text{N}_2]}{dt} = 2.5 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

And the rate of production of H_2 will be:

$$\frac{d[\text{H}_2]}{dt} = 3 \times 2.5 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1} = 7.5 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

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

$$\text{Rate} = k [\text{CH}_3\text{OCH}_3]^{3/2}$$

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.,

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

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

Ans: If the pressure is measured in bar and time in minutes, then unit of rate will be:
= bar min^{-1}

We are given the rate of the reaction as: $\text{Rate} = k [\text{CH}_3\text{OCH}_3]^{3/2}$

$$\text{Therefore, } k = \frac{\text{Rate}}{[\text{CH}_3\text{OCH}_3]^{3/2}}$$

So, we can write the units of rate constant as:

$$k = \frac{\text{bar min}^{-1}}{\text{bar}^{3/2}} = \text{bar}^{-1/2} \text{ min}^{-1}$$

So, the units are $\text{bar}^{-1/2} \text{ min}^{-1}$.

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

Ans: Factors that influence a reaction's speed include.

- (i) Reactant nature: The rate of reaction is affected by the kind of reactant. For example, ionic compound reactions are quicker than covalent compound reactions.
- (ii) The state of the reactants: Solid reactions are sluggish, liquid reactions are rapid, and gas reactions are very quick.
- (iii) Temperature: In addition, temperature has a significant impact on response rate. Every 10 degrees Celsius increase in temperature doubles the pace of reaction.
- (iv) Presence of catalyst: The rate of reaction is also affected by the presence of a catalyst in the reaction. It enhances the pace of reaction by increasing reaction surface area, by generating unstable intermediates with the substrate, and by offering a lower activation energy alternative path.

6. 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

Ans: Let us assume that the concentration of the reactant be $[A] = a$

Rate of the reaction will be:

$$R = k[A]^2 = ka^2$$

It is said that the concentration of A increases by two times, so we can write:

$$[A] = 2a \text{ mol/L}$$

Therefore, the rate of reaction will be:

$$\text{Rate} = k (2a)^2 = 4 ka^2$$

Thus, the rate of the reaction will increase by 4 times.

(ii) Reduced to half

It is said that the concentration of A reduced to half, so we can write:

$$[X] = \frac{1}{2} \text{ a mol/ L}$$

Therefore, the rate of reaction will be:

$$\text{Rate} = k \left[\frac{1}{2} \text{ a} \right]^2 = \frac{1}{4} \text{ ka}^2$$

Thus, the rate of the reaction will reduce by $\frac{1}{4}$ times.

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

Ans: An increase of 10 degrees in temperature causes a reaction's rate constant to almost double in size. In any case, the Arrhenius equation gives the exact temperature dependency of a chemical reaction rate.

The Arrhenius equation is given below:

$$k = A e^{-E_a/Rt}$$

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.

The formula can also be written as:

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

Where k_2 is the rate constant at temperature T_2

k_1 is the rate constant at temperature T_1

8. 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.

Ans: Between the time interval of 30 to 60 sec the average rate of reaction will be calculated as:

$$= \frac{d[\text{Ester}]}{dt}$$

Putting the values from the data given in the question, we can write:

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

$$= 4.67 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}$$

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

Ans: The rate law formula for pseudo first order reaction will be:

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

Now, when $t = 30$ s, then we rate constant will be:

$$k_1 = \frac{2.303}{30} \log \frac{0.55}{0.31} = 1.91 \times 10^{-2} \text{ s}^{-1}$$

Now, when $t = 60$ s, then we rate constant will be:

$$k_2 = \frac{2.303}{60} \log \frac{0.55}{0.17} = 1.96 \times 10^{-2} \text{ s}^{-1}$$

Now, when $t = 90$ s, then we rate constant will be:

$$k_3 = \frac{2.303}{90} \log \frac{0.55}{0.085} = 2.075 \times 10^{-2} \text{ s}^{-1}$$

So, we can calculate the average rate constant as:

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

$$= 1.98 \times 10^{-2} \text{ s}^{-1}$$

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

(i) Write the differential rate equation

Ans: The differential rate equation can be written as:

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

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

Ans: The concentration of B is increased by 3 times, then $B = 3B$

Therefore, the rate will be:

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

Therefore, the rate will increase by 9 times.

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

Ans: The concentration of A is doubled, then $A = 2A$

The concentration of B is doubled, then $B = 2B$

Therefore, the rate will be:

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

Therefore, the rate will increase by 8 times.

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

A/mol L⁻¹	0.20	0.20	0.04
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?

Ans: Let us assume that the order of the reaction with respect to A be x and with respect to B be y.

Therefore, we can write:

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

$$5.07 \times 10^{-5} = k [0.20]^x [0.30]^y \quad \text{.....(i)}$$

$$5.07 \times 10^{-5} = k [0.20]^x [0.10]^y \quad \text{.....(ii)}$$

$$1.43 \times 10^{-5} = k [0.40]^x [0.05]^y \quad \text{.....(iii)}$$

Let us divide (i) by (ii), we get:

$$\frac{5.07 \times 10^{-5}}{5.07 \times 10^{-5}} = \frac{k [0.20]^x [0.30]^y}{k [0.20]^x [0.10]^y}$$

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

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

Therefore, $y = 0$

Now, dividing (iii) by (i), we get:

$$\frac{1.43 \times 10^{-4}}{5.07 \times 10^{-5}} = \frac{k [0.40]^x [0.05]^y}{k [0.20]^x [0.30]^y}$$

$$\frac{1.43 \times 10^{-4}}{5.07 \times 10^{-5}} = \frac{k [0.40]^x}{k [0.20]^x}$$

Because $y = 0$

$$= 2.821 = 2^x$$

Now, taking log on both the sides.

$$\log 2.821 = x \log 2$$

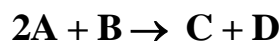
$$x = \frac{\log 2.821}{\log 2}$$

$$x = 1.496$$

$$x = 1.5$$

Thus, the order of the reaction according to A is 1.5 and according to B is 0.

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



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	7.3×10^{-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.

Ans: Let us assume that the order of the reaction with respect to A be x and with respect to B be y.

Therefore, we can write:

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

$$6.0 \times 10^{-3} = k [0.1]^x [0.1]^y \quad \text{.....(i)}$$

$$7.2 \times 10^{-2} = k [0.3]^x [0.2]^y \quad \text{.....(ii)}$$

$$2.88 \times 10^{-1} = k [0.3]^x [0.4]^y \quad \text{.....(iii)}$$

$$2.40 \times 10^{-2} = k [0.4]^x [0.1]^y \quad \text{.....(iv)}$$

Let us divide (iv) by (i), we get:

$$\frac{2.40 \times 10^{-2}}{6.0 \times 10^{-3}} = \frac{k [0.4]^x [0.1]^y}{k [0.1]^x [0.1]^y}$$

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

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

Therefore, x = 1

Now, dividing (iii) by (ii), we get:

$$\frac{2.88 \times 10^{-1}}{7.2 \times 10^{-2}} = \frac{k [0.3]^x [0.4]^y}{k [0.3]^x [0.2]^y}$$

$$= 4 = \left(\frac{0.4}{0.2}\right)^y$$

$$= 4 = 2^y$$

$$= 2^2 = 2^y$$

$$y = 2$$

Thus, the order of the reaction according to A is 1 and according to B is 2.

So, the rate law is:

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

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

Now, putting the values for each experiment, we get:

From experiment I:

$$k = \frac{6.0 \times 10^{-3}}{[0.1][0.1]^2}$$

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

From experiment II:

$$k = \frac{7.3 \times 10^{-2}}{[0.3][0.2]^2}$$

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

From experiment III:

$$k = \frac{2.88 \times 10^{-1}}{[0.3][0.4]^2}$$

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

From experiment IV:

$$k = \frac{2.40 \times 10^{-2}}{[0.4][0.1]^2}$$

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

Therefore, the rate constant will be $k = 6.0 \text{ L}^2 \text{ mol}^{-2} \text{ min}^{-1}$

12. 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 D/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}

Ans: We are given that, the reaction between A and B is first order with respect to A and zero order with respect to B.

Therefore, the rate of the reaction is given by:

$$\text{Rate} = k[A]^1[B]^0$$

So, we can write:

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

According to the experiment I we can write:

$$2.0 \times 10^{-2} = k(0.1)$$

$$k = 0.2 \text{ min}^{-1}$$

According to the experiment II we can write:

$$4.0 \times 10^{-2} = (0.2) [A]$$

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

According to the experiment III we can write:

$$\text{Rate} = 0.2 \times 0.4$$

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

According to the experiment III we can write:

$$2.0 \times 10^{-2} = (0.2) [A]$$

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

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

(i) 200 s^{-1}

Ans: Half-life of the reaction can be related with the rate constant of the reaction as:

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

Putting the value of time, we get:

$$k = \frac{0.693}{200} = 3.46 \times 10^{-3} \text{ s}^{-1}$$

So, the rate of the reaction is $3.46 \times 10^{-3} \text{ s}^{-1}$

(ii) 2 min^{-1}

Ans: Half-life of the reaction can be related with the rate constant of the reaction as:

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

Putting the value of time, we get:

$$k = \frac{0.693}{2} = 0.346 \text{ min}^{-1}$$

So, the rate of the reaction is 0.346 min^{-1} .

(iii) 4 years^{-1}

Ans: Half-life of the reaction can be related with the rate constant of the reaction as:

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

Putting the value of time, we get:

$$k = \frac{0.693}{4} = 0.173 \text{ years}^{-1}$$

So, the rate of the reaction is 0.173 years^{-1} .

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

Ans: The given reaction in the above question is radioactive decay and the radioactive decay follows the first order kinetics. Therefore, the decay constant:

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

We are given the half-life time as 5730 years.

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

To find the age of the sample, we can write:

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

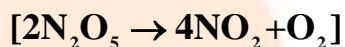
80% of the wood is found so, the initial amount can be taken as 100 and the final amount as 80. Putting the values, we get:

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

$$t = 1845 \text{ years}$$

Therefore, the age of the sample is 1845 years.

15. The experiment data for decomposition of N_2O_5

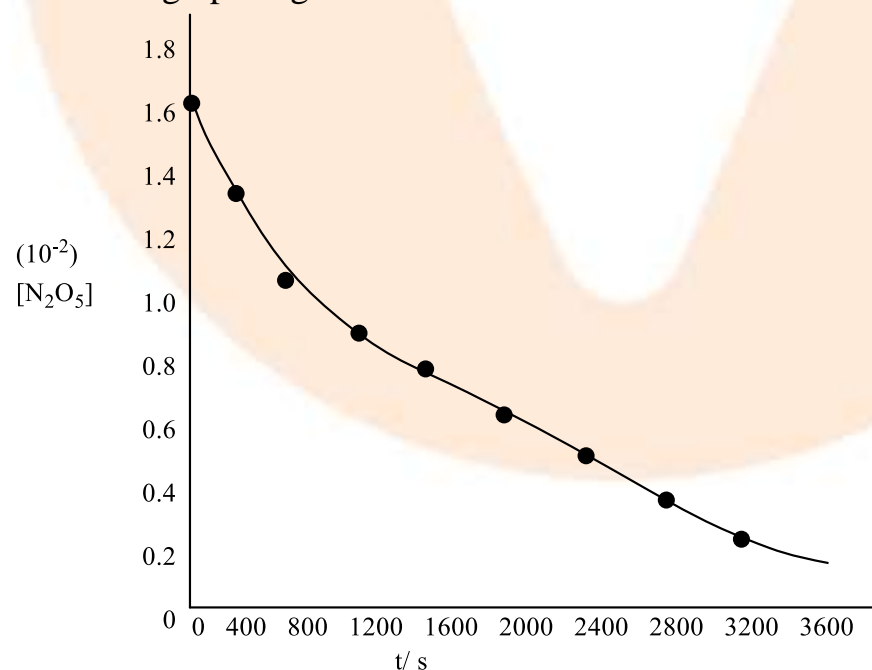


In gas phase at 318 K are given below:

t/s	0	400	800	1200	1600	2000	2400	1800	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

Ans: The graph is given below:



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

Ans: Time corresponding to the concentration,

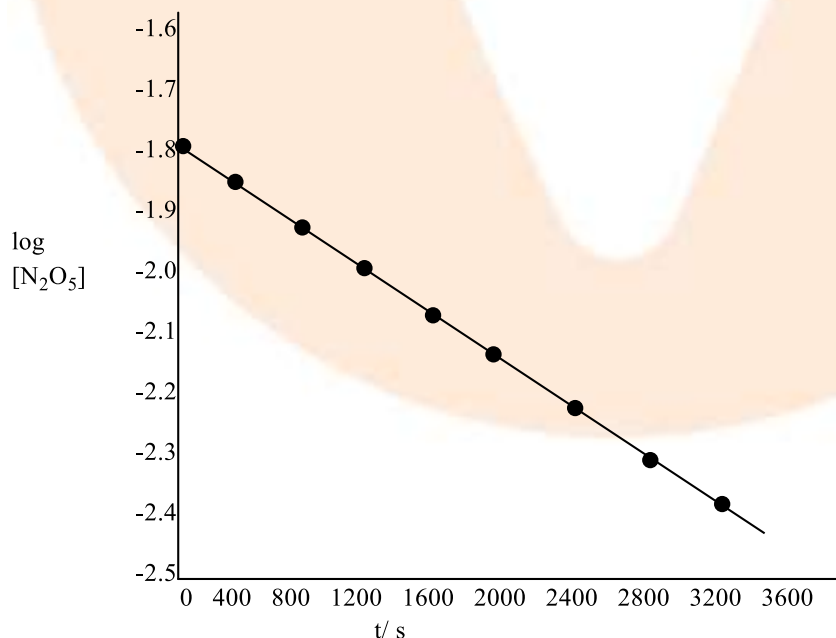
$$\frac{1.630 \times 10^{-2}}{2} \text{ mol L}^{-1} = 0.815 \times 10^{-2} \text{ mol L}^{-1} \text{ is the half-life.}$$

From the graph, the half-life obtained is 1440 s.

(iii) Draw the graph between $\log[\text{N}_2\text{O}_5]$ and t

Ans:

T(s)	$10^{-2} \times [\text{N}_2\text{O}_5]/\text{mol L}^{-1}$	$\log[\text{N}_2\text{O}_5]$
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) What is the rate law?

Ans: The rate law of the reaction will be:

$$\text{Rate} = k[\text{N}_2\text{O}_5]$$

(v) Calculate the rate constant.

Ans: From the plot $[\text{N}_2\text{O}_5]$ v/s t , is given by:

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

Therefore, we obtain:

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

$$= 4.82 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$$

(vi) Calculate the half-life period from k and compare it with (ii)

Ans: Half-life is given by:

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{4.82 \times 10^{-4}} \text{ s}$$

$$= 1438 \text{ seconds.}$$

The value of half-life calculated from the k is very close to that obtained from the graph.

16. 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 $1/16^{\text{th}}$ value?

Ans: As we know that:

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

The initial value of the reactant has become $1/16^{\text{th}}$. Now, putting the values, we get:

$$t = \frac{2.303}{60} \log \frac{16}{1}$$

$$t = \frac{2.303}{60} \log 16 = 4.62 \times 10^{-2} \text{ s}$$

Therefore, the time required will be 4.62×10^{-2} s.

17. During nuclear explosion, one of the products is ^{90}Sr with a half-life of 28.1 years. If 1 μg 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.

Ans: As radioactive disintegration follows first order kinetics.

$$\text{Decay constant of } ^{90}\text{Sr} (k) = \frac{0.693}{t_{1/2}} = \frac{0.693}{28.1} = 2.466 \times 10^{-2} \text{ y}^{-1}$$

To calculate the amount left after 10 years.

$$a = 1 \mu\text{g}$$

$$t = 10 \text{ years}$$

$$k = 2.466 \times 10^{-2} \text{ y}^{-1}$$

$$(a-x) = ?$$

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

$$2.466 \times 10^{-2} = \frac{2.303}{10} \log \frac{1}{a-x}$$

$$\log(a-x) = -0.1071$$

$$(a-x) = \text{Antilog } -0.1071 = 0.7814 \mu\text{g}$$

To calculate the amount left after 60 years.

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

$$2.466 \times 10^{-2} = \frac{2.303}{60} \log \frac{1}{a-x}$$

$$\log(a-x) = -0.6425$$

$$(a-x) = \text{Antilog } -0.6425 = 0.2278 \mu\text{g}$$

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

Ans: For first order reaction, we can write:

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

99% completion means that $x = 99\%$ of $a = 0.99 a$

So, we can write:

$$t_{99\%} = \frac{2.303}{k} \log \frac{a}{a-0.99a}$$

$$t_{99\%} = \frac{2.303}{k} \log 10^2$$

$$t_{99\%} = 2 \times \frac{2.303}{k}$$

90% completion means that $x = 90\%$ of $a = 0.90 a$

$$t_{90\%} = \frac{2.303}{k} \log \frac{a}{a-0.90a}$$

$$t_{90\%} = \frac{2.303}{k} \log 10$$

$$t_{90\%} = \frac{2.303}{k}$$

Now, we can take the ratio as given below:

$$\frac{t_{99\%}}{t_{90\%}} = \frac{\left(\frac{2 \times 2.303}{k} \right)}{\left(\frac{2.303}{k} \right)}$$

$$t_{99\%} = 2 \times t_{90\%}$$

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.

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

Ans: 30% decomposition means that $x = 30\%$ of $a = 0.30 a$

Since, the reaction is of 1st order, we can write:

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

Time is given as 40 min. So, putting the values, we get:

$$k = \frac{2.303}{40} \log \frac{a}{a-0.30a}$$

$$k = \frac{2.303}{40} \log \frac{10}{7} \text{ min}^{-1}$$

$$k = \frac{2.303}{40} \times 0.1549 \text{ min}^{-1} = 8.918 \times 10^{-3} \text{ min}^{-1}$$

Now, we can calculate the half-life period as we have the rate constant value.

We can write:

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{8.918 \times 10^{-3}} = 7.77 \text{ min}$$

So, the half-life is 7.77 min.

20. 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.

Ans: The decomposition of azoisopropane to hexane and nitrogen at 54.3 k is represent by the following equation.



Total pressure after time t, we will be:

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

$$P_t = P_0 + p$$

$$p = P_t - P_0$$

Now, we can substitute the value of p for the pressure of reactant at time t

$$=P_o - p$$

$$=P_o - (P_t - P_o)$$

$$= 2P_o - P_t$$

Now, we can apply the rate constant formula of 1st order reaction.

$$k = \frac{2.303}{t} \log \frac{P}{2P_o - P_t}$$

When $t = 360$ s,

Putting the values, we get:

$$k = \frac{2.303}{360} \log \frac{35.0}{2 \times 35 - 54} = 2.175 \times 10^{-3} \text{ s}^{-1}$$

When $t = 720$ s,

Putting the values, we get:

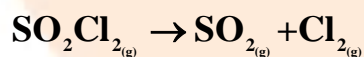
$$k = \frac{2.303}{720} \log \frac{35.0}{2 \times 35 - 63} = 2.235 \times 10^{-3} \text{ s}^{-1}$$

Now, we can find the average value:

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

$$k = 2.20 \times 10^{-3} \text{ s}^{-1}$$

21. The following data were obtained during the first order thermal decomposition of SO_2Cl_2 at a constant volume.



Experiment	Time/ s	Total pressure/ atm
1	0	0.5
2	100	0.6

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

Ans: The given reaction shows the thermal decomposition of SO_2Cl_2 at constant volume.



Total pressure after time t, we will be:

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

$$P_t = P_0 + p$$

$$p = P_t - P_0$$

Now, we can substitute the value of p for the pressure of reactant at time t

$$= P_0 - p$$

$$= P_0 - (P_t - P_0)$$

$$= 2P_0 - P_t$$

Now, we can apply the rate constant formula of 1st order reaction.

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

When the t = 100 s

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

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

When $P_t = 0.65 \text{ atm}$

Therefore, pressure of SO_2Cl_2 at time t total pressure is 0.65 atm, is

$$P_{\text{SO}_2\text{Cl}_2} = 2P_0 - P_t$$

$$= 2 \times 0.50 - 0.65$$

$$= 0.35 \text{ atm}$$

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

$$\text{Rate} = k(P_{\text{SO}_2\text{Cl}_2})$$

$$\text{Rate} = (2.33 \times 10^{-3})(0.354) = 7.8 \times 10^{-4} \text{ atm s}^{-1}$$

22. The rate constant for the decomposition of N_2O_5 at various temperature is

given below:

T/°c	0	20	40	60	80
$10^5 \times k/s$	0.0787	1.70	25.7	178	2140

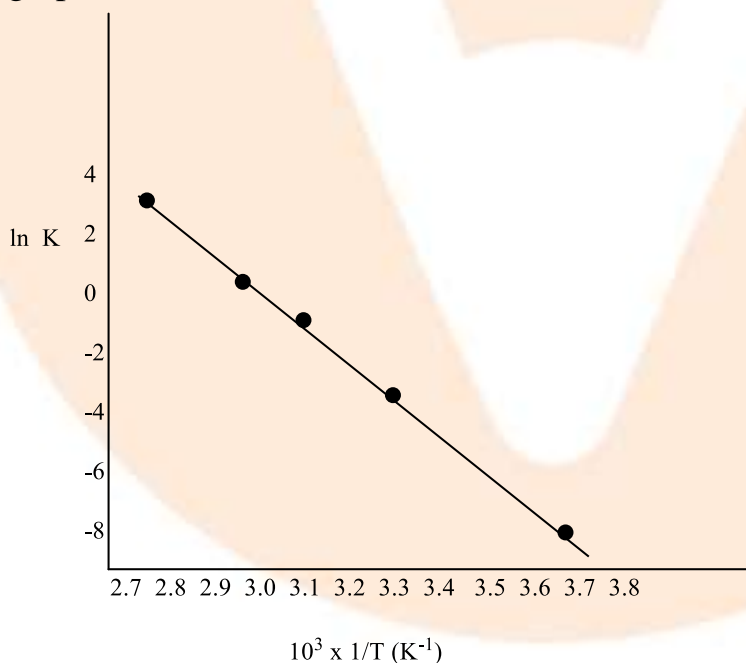
Draw a graph between $\ln K$ and $1/T$ and calculate the values of A and E_a .

Predict the rate constant at 30° and 50° C.

Ans: As the data is given we can write:

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}
$10^5 \times k/s$	0.0787	4.075	25.7	178	2140
$\ln K$	-7.147	-4.075	-1.359	-0.577	3.063

The graph is draw below:



Slope of the line, will be given as:

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

According to the Arrhenius equation,

$$\begin{aligned}\text{Slope} &= -\frac{E_a}{R} \\ &= E_a = -\text{slope} \times R \\ &= (-12.301) (8.314) \\ &= 102.27 \text{ KJ mol}^{-1}\end{aligned}$$

Again,

$$\begin{aligned}\ln k &= \ln A - \frac{E_a}{RT} \\ \ln A &= \ln k + \frac{E_a}{RT}\end{aligned}$$

As $T = 273 \text{ K}$ and $\ln k = -7.147$

Applying this in the formula, we get:

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

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

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

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

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

$$\ln k = -2.8$$

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

When $T = 50 + 273 \text{ K} = 323 \text{ K}$

$$\frac{1}{T} = 0.0031 \text{ K}^{-1} = 3.1 \times 10^{-3} \text{ K}^{-1}$$

Now, at $\frac{1}{T} = 0.0031 \text{ K}^{-1} = 3.1 \times 10^{-3} \text{ K}^{-1}$

$$\ln k = -0.5$$

Therefore, $k = 0.607 \text{ s}^{-1}$

23. The rate constant for the decomposition of hydrocarbons is $2.418 \times 10^{-5} \text{ s}^{-1}$

at 546 K. If the energy of activation is 179.9 kJ /mol, what will be the value of the pre-exponential factor?

Ans: We are given some values as:

$$K = 2.418 \times 10^{-5} \text{ s}^{-1}$$

$$T = 546 \text{ K}$$

$$E_a = 179.9 \text{ kJ mol}^{-1} = 179.9 \times 10^3 \text{ J mol}^{-1}$$

We use the Arrhenius equation is:

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

In the log form, this can be written as:

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

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

$$\log k = \log (2.418 \times 10^{-5}) - \frac{179.9 \times 10^3}{2.303 \times 8.314 \times 546}$$

$$= (0.3835 - 5) + 17.2082 = 12.5917$$

Therefore, $A = \text{antilog}(12.5917)$

$$A = 3.912 \times 10^{12} \text{ s}^{-1}$$

24. Consider a certain reaction $A \rightarrow \text{Products}$ with $k = 2.0 \times 10^{-2} \text{ s}^{-1}$. Calculate the concentration of A remaining after 100s if the initial concentration of A is 1.0 mol L^{-1} .

Ans: We are given some values, that are given below:

$$k = 2.0 \times 10^{-2} \text{ s}^{-1}$$

$$t = 100 \text{ s}$$

$$[A]_0 = 1.0 \text{ mol L}^{-1}$$

As we can see that the units of k is given in s^{-1} , this means that the reaction is a first order reaction.

Therefore, we can write:

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

Putting the values, we get:

$$2.0 \times 10^{-2} = \frac{2.303}{100} \log \frac{1.0}{[A]}$$

$$2.0 \times 10^{-2} = \frac{2.303}{100} (-\log[A])$$

$$(-\log[A]) = \frac{2.0 \times 10^{-2} \times 100}{2.303}$$

$$[A] = \text{antilog} \left(\frac{2.0 \times 10^{-2} \times 100}{2.303} \right)$$

$$= 0.135 \text{ mol L}^{-1}$$

Therefore, the remaining amount of A is 0.135 mol L^{-1} .

25. 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 a sample of sucrose remains after 8 hours?

Ans: The given reaction is a first order reaction. So, we can write:

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

We are given a half-life of 3 hours. Therefore, we can write:

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

So, putting the values in this, we get:

$$k = \frac{0.693}{3} = 0.231 \text{ h}^{-1}$$

Now, we can put this value of rate constant in the first order reaction formula.

$$0.231 = \frac{2.303}{8} \log \frac{[R]_0}{[R]}$$

$$\log \frac{[R]_0}{[R]} = \frac{0.231 \times 8}{2.303}$$

$$\log \frac{[R]_0}{[R]} = 0.8024$$

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

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

Or we can write:

$$\frac{[R]}{[R]_0} = 0.158$$

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

26. The decomposition of hydrocarbon follows the equation

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

Ans: According to the Arrhenius equation,

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

We are given the equation as:

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

Therefore, the formula can be written as:

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

This can be written as:

$$E_a = 28000 \times R$$

$$E_a = 28000 \times 8.314 = 232.79 \text{ kJ mol}^{-1}$$

Therefore, the value of E_a is $232.79 \text{ kJ mol}^{-1}$

27. 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?

Ans: According to the Arrhenius equation,

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

This can be written as:

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

In the log form it can be written as:

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

We are given:

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

Comparing these two, we get:

$$\frac{E_a}{2.303 RT} = \frac{1.25 \times 10^4 \text{ K}}{T}$$

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

$$E_a = 2.303 \times 8.314 \times 1.25 \times 10^4 \text{ K}$$

$$E_a = 239.34 \text{ kJ mol}^{-1}$$

We are given half-life time as 256 minutes.

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

$$k = \frac{0.693}{256 \times 60} = 4.51 \times 10^{-5} \text{ s}^{-1}$$

Now, we have the value of rate constant, we can put in the equation:

$$\log(4.51 \times 10^{-5}) = 14.34 - \frac{1.25 \times 10^4}{T}$$

$$T = 669 \text{ K}$$

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

Ans: We are some information:

$$k_1 = 4.5 \times 10^3$$

$$T_1 = 10 + 273 = 283 \text{ K}$$

$$k_2 = 1.5 \times 10^4$$

$$T_2 = ?$$

$$E_a = 60 \text{ kJ mol}^{-1}$$

Applying Arrhenius equation:

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

Putting the values, we can write:

$$\log \frac{1.5 \times 10^4}{4.5 \times 10^3} = \frac{60}{2.303 \times 8.314} \left(\frac{T_2 - 283}{283 T_2} \right)$$

$$\log 3.333 = 3133.63 \left(\frac{T_2 - 283}{283 T_2} \right)$$

$$\frac{0.5228}{3133.63} = \left(\frac{T_2 - 283}{283 T_2} \right)$$

$$0.0472 T_2 = T_2 - 283$$

$$T_2 = 297 \text{ K}$$

Or we can write:

$$T_2 = 24^\circ \text{C}$$

29. 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 .

Ans: There are two cases in this question. As the reaction given is first order reaction, we can use:

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

For case 1:

$$k_{298\text{K}} = \frac{2.303}{t_1} \log \frac{a}{a - 0.10a}$$

$$k_{298K} = \frac{2.303}{t_1} \log \frac{10}{9}$$

$$k_{298K} = \frac{2.303}{t_1} \times 0.0458$$

$$t_1 = \frac{0.1055}{k_{298K}}$$

For case 2:

$$k_{308K} = \frac{2.303}{t_2} \log \frac{a}{a - 0.25a}$$

$$k_{308K} = \frac{2.303}{t_2} \log \frac{4}{3}$$

$$k_{308K} = \frac{2.303}{t_2} \times 0.125$$

$$t_2 = \frac{0.2879}{k_{308K}}$$

But $t_1 = t_2$

Hence,

$$\frac{0.1055}{k_{298K}} = \frac{0.2879}{k_{308K}}$$

$$\frac{k_{308K}}{k_{298K}} = 2.7289$$

Now, applying the Arrhenius equation,

$$\log \frac{k_{308K}}{k_{298K}} = \frac{E_a}{2.303R} \left(\frac{T_2 - T_1}{T_1 T_2} \right)$$

$$\log(2.7289) = \frac{E_a}{2.303 \times 8.314} \left(\frac{308 - 298}{298 \times 308} \right)$$

$$E_a = 76.623 \text{ kJ mol}^{-1}$$

Now, the calculation of k at 318 K

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

$$\log k = \log (4 \times 10^{10}) - \frac{7623}{2.303 \times 8.314 \times 318}$$

$$\log k = 10.6021 - 12.5843 = -1.9822$$

$$k = \text{Antilog} (-1.9822) = \text{antilog} (\bar{2}.0178) = 1.042 \times 10^{-2} \text{ s}^{-1}$$

Therefore, k is $1.042 \times 10^{-2} \text{ s}^{-1}$.

30. 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.

Ans: We are given that:

$$k_2 = 4 k_1$$

$$T_1 = 293\text{K}$$

$$T_2 = 313\text{K}$$

According to the Arrhenius equation, we get:

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

Putting the values, we get:

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

$$0.6021 = \frac{E_a}{2.303 \times 8.314} \left(\frac{313 - 293}{293 \times 313} \right)$$

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

$$E_a = 52863.00 \text{ J mol}^{-1}$$

$$E_a = 52.863 \text{ kJ mol}^{-1}$$

Therefore, the required activation energy is $52.863 \text{ kJ mol}^{-1}$.