

CHAPTER - 4

CHEMICAL KINETICS INTEXT SOLUTIONS

Question 4.1:

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.

Answer

$$\begin{aligned}\text{Average rate of reaction} &= -\frac{\Delta[R]}{\Delta t} \\ &= -\frac{[R]_2 - [R]_1}{t_2 - t_1} \\ &= -\frac{0.02 - 0.03}{25} \text{ M min}^{-1} \\ &= -\frac{-0.01}{25} \text{ M min}^{-1} \\ &= 4 \times 10^{-4} \text{ M min}^{-1} \\ &= \frac{4 \times 10^{-4}}{60} \text{ M s}^{-1} \\ &= 6.67 \times 10^{-6} \text{ M s}^{-1}\end{aligned}$$

Question 4.2:

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?

Answer

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

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

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

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

$$= 0.005 \text{ mol L}^{-1} \text{ min}^{-1}$$

$$= 5 \times 10^{-3} \text{ M min}^{-1}$$

Question 4.3:

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?

Answer

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

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

$$= 2.5$$

Question 4.4:

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 $X \rightarrow Y$ follows second order kinetics.

Therefore, the rate equation for this reaction will be:

$$\text{Rate} = k[X]^2 \quad (1)$$

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

$$\text{Rate}_1 = k \cdot (a)^2$$

$$= ka^2$$

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

Now, the rate equation will be:

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

$$= 9(ka^2)$$

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

Question 4.5:

A first order reaction has a rate constant $1.15 \times 10^{-3} \text{ s}^{-1}$. 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 = $1.15 \times 10^{-3} \text{ s}^{-1}$

We know that for a 1st order reaction,

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

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

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

$$= 444.38 \text{ s}$$

$$= 444 \text{ s (approx)}$$

Question 4.6:

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

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

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

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

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

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

Question 4.7:

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,

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

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

Question 4.8:

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, $R = 8.314$ J K⁻¹ 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:

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

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

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

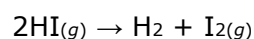
$= 52897.78$ J mol⁻¹

$= 52.9$ kJ mol⁻¹

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

Question 4.9:

The activation energy for the reaction



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

Answer

In the given case:

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

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

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

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

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

$$\Rightarrow \ln x = -E_a / RT$$

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

$$\Rightarrow \log x = \frac{209500 \text{ J mol}^{-1}}{2.303 \times 8.314 \text{ JK}^{-1} \text{ mol}^{-1} \times 581} = 18.8323$$

$$\text{Now, } x = \text{Anti log}(18.8323)$$

$$= \text{Anti log } \overline{19.1677}$$

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