

## Intext Questions

## 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

Average rate of reaction

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

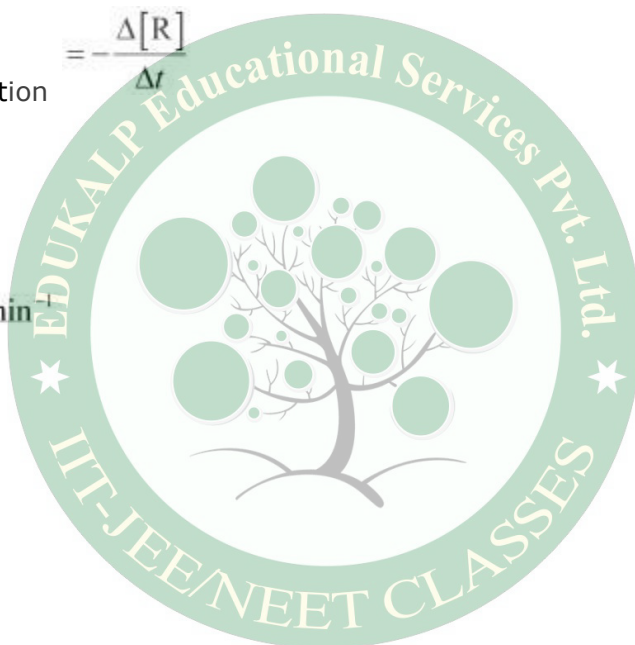
$$= -\frac{0.02 - 0.03}{25} \text{ M min}^{-1}$$

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

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

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

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



## Question 4.2:

In a reaction,  $2A \rightarrow \text{Products}$ , the concentration of A decreases from 0.5 mol L<sup>-1</sup> to 0.4 mol L<sup>-1</sup> 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 1<sup>st</sup> 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  $\text{SO}_2\text{Cl}_2$  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<sup>st</sup> 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

$$\begin{aligned}\therefore T_2 &= (298 + 10) \text{ K} \\ &= 308 \text{ K}\end{aligned}$$

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:

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

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

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

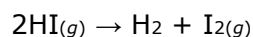
$$= 52897.78 \text{ J mol}^{-1}$$

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

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

#### Question 4.9:

The activation energy for the reaction



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

Answer

In the given case:

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

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

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

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

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

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

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

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

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

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

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

