Intext Questions

Question 4.1:

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}

Question 4.2:

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

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

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

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

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

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

Question 4.3:

For a reaction, A + B → 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

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 → Y follows second order kinetics.

Therefore, the rate equation for this reaction will be:

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

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

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

$= ka^2$

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.

Question 4.5:

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,

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

$$= \frac{2.303}{1.15\times10^{-3}}\log\frac{5}{3}$$

$$= \frac{2.303}{1.15\times10^{-3}}\times 0.2219$$

= 444.38 s

= 444 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,

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

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

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

Now, substituting these values in the equation:

$$\log\frac{k_2}{k_1} = \frac{E_a}{2.303\ R}\left[\frac{T_2 - T_1}{T_1T_2}\right]$$

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

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

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:

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

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

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

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

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

$$= \text{Anti}\log \overline{19}.1677$$

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