

Chapter 3: Chemical Kinetics

NCERT · Class 12 · Chemistry

Textbook exercises

Question 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.

Solution

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:

$$\text{Average rate} = - \frac{\Delta[R]}{\Delta t}$$

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 $6.66 \times 10^{-6} \text{ M s}^{-1}$

Question 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.

Solution

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

Question 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?

Solution

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:

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

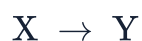
So, the order of the reaction is 2.5.

Question 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?

Solution

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:

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

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.

Question 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?

Solution

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.

Question 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.

Solution

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 $1.925 \times 10^{-4} \text{ s}^{-1}$.

Question 7

What will be the effect of temperature on rate constant?

Solution

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.

Question 8

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

Solution

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:

$$k_1 = x$$

$$k_2 = 2x$$

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

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

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

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

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

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

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

Question 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.

Solution

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:

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

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