# Class 12 Chemistry

## NCERT Solutions for Chapter 4 – Chemical Kinetics
## Intext Questions

**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:

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

This can be written as:

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

$R_1 = 0.03 \, M$

$R_2 = 0.02 \, M$

$t_2 - t_1 = 25 \, min$

Putting the values, we get:

Average rate = $- \dfrac{0.02 - 0.03}{25}$

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

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

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

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

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

**2. In a reaction, 2A $\rightarrow$ 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:

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

This is due the fact that the reaction given is:

2A $\rightarrow$ Products

So, the average rate will be written as:

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

$A_1 = 0.5$ M

$A_2 = 0.4$ M

$t_2 - t_1 = 10$ min

Putting the values, we get:

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

Average rate = $-\frac{1}{2}\times\frac{-0.1}{10}=5\times10^{-3}$ M min$^{-1}$

So, the average rate will be $5\times10^{-3}$ M min$^{-1}$

**3. For a reaction,** A + B $\rightarrow$ 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:

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:

$X \rightarrow Y X \rightarrow Y$

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

Rate = k [ X ] 2 = ka2Rate = k [ X ] 2 = ka2

If X◆ = a mol/ L

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

X◆ = 3a mol/ L

Therefore, the rate of reaction will be:

Rate = k (3a)2 = 9 ka2Rate = k (3a)2 = 9 ka2

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** $1.15 \times 10\text{-}3 \ s\text{-}11.15 \times 10\text{-}3 \ s\text{-}1$**. 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 g [ R ] 0 = 5 g

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

R◆ = 3 g

We are also given the value of rate constant as:

Rate constant = 1.15 × 10-3 s-1Rate constant = 1.15 × 10-3 s-1

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

t = 2.303k log [ R ] 0 [ R ] t = 2.303k log [ R ] 0 [ R ]

Putting the values in the above formula, we get:

t = 2.3031.15 x 10−3 log53t = 2.3031.15 x 10−3 log53

t = 2.3031.15 × 10−3 × 0.2219t = 2.3031.15 × 10−3 × 0.2219

t = 444 seconds

So, the time taken will be 444 seconds.


**6. Time required to decompose** $SO_2Cl_2SO_2Cl_2$ **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 $SO_2Cl_2SO_2Cl_2$ is a first order reaction. So, we can write:

t1/2 = 0.693kt1/2 = 0.693k

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

t1/2 = 60 mint1/2 = 60 min

Therefore, k = 0.693t1/2k = 0.693t1/2

Putting the value, we get:

k = 0.69360 × 60 = 1.925   × 10-4 s-1k = 0.69360 × 60 = 1.925   × 10-4 s-1

Thus, the rate constant is 1.925   × 10-4 s-11.925   × 10-4 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-Ea/Rtk = 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.


**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 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, $R = 8.314$ J K-1 mol-1

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

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

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

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

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

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


**9. The activation energy for the reaction $2HI(g) \rightarrow H_2(g) + I_2(g)$ is 209.5 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 kJ mol-1

T = 581 K

R = 8.314 J K-1 mol-1R = 8.314 J K-1 mol-1

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

x = eEa/RTx = eEa/RT

ln x =-EaRTln x =-EaRT

log x =-Ea2.303 RTlog x =-Ea2.303 RT

log x =2095002.303 × 8.314 × 581 = 18.8323log x =2095002.303 × 8.314 × 581 = 18.8323

Now, taking the antilog:

x = Antilog (18.8323)

x = 1.471 × 10-19x = 1.471 × 10-19