SHIKSHA CLASSES

Subject : Chemistry **BOARD ANSWER PAPER** **Total Marks : 20**

Class : XII **Topic: 6) Chemical Kinetics**

Section (A)

Q.1 : a) Select and write the most appropriate answer from the following alternatives of each sub question. (05)

i) The rate constant of reaction .....

Ans : a) decreases with increasing E_a

ii) The slope of straight line obtained by plotting rate versus conc. of reactant for a first order reaction is

Ans : d) K

iii) The half life of the first order reaction having rate constant K = 1.7×10^{-5} s^{-1} is

Ans : c) 11.1 hour

iv) The rate constant of zero order reaction has the unit

Ans : b) mol L^{-1} s^{-1}

v) The order of the reaction for which the units of rate constant are mol dm^{-3} s^{-1} is

Ans : c) 0

Q.1 : (b) Very short answer type Question. [2]

1) What is molecularity of the reaction ?

Ans : The molecularity refers to How many reactant molecules involved in reaction

eg. $O_{3(g)} \rightarrow O_{2(g)} + O_{(g)}$

In above reaction there is only one reactant molecule. These are unimolecular reaction or molecularity is one.

2) What is rate law for a reaction.

$NO_{2(g)} + CO_{(g)} \rightarrow NO_{(g)} + CO_{2(g)}$ **If the reaction occurs in the following steps?**

i) $NO_2 + NO_2 \rightarrow NO_3 + NO$ (Slow)

ii) $NO_3 + NO \rightarrow NO_2 + CO_2$ (Fast)

Ans : The rate law for reaction is rate $= -K[NO_2]^2$

NO_3 is reaction intermediate.

Section (B)

Q.2 : Attempt Any Three of the following question. (06)

1) Define rate law. Explain with one example.

Ans : Rate law is defined as an experimentally determined equation that expresses the rate of a chemical reaction in terms of molar concentrations of the reactants.

Ex. : $2H_2O_2 \xrightarrow{\Delta} 2H_2O_{(l)} + O_2(g)$

The rate law for reaction is

$$\text{rate} = -\frac{1}{2}\frac{d[H_2O_2]}{dt} = K[H_2O_2]$$

The reaction is first order reaction

2) Derive the expression for half life of zero order raction.

Ans : The integrated rate law for zero order reaction is $K = \frac{[A]_0 - [A]_t}{t}$

where $[A]_t$ is concentration of A that remains unreacted at time t.

at $t = t_{1/2}, [A]_t = \frac{[A]_0}{2}$

$K = \frac{[A]_0 - [A]_0/2}{t_{1/2}}$ $\quad K = \frac{[A]_0}{2t_{1/2}}$ i.e. $t_{1/2} = \frac{[A]_0}{2K}$

3) Write Arrehenius equation and explain the terms inolved.

Ans : The Arrhenius equation is -

$$\ln k = \ln A - \frac{E_a}{RT}$$

$$\text{or } \log_{10} K = \log_{10} A - \frac{E_a}{2.303RT}$$

when K= rate constant

A= frequency factor or pre exponential factor

E_a = activation energy

T = temperature

4) What is zero order reaction? Derive integrated rate law for zero order reaction.

Ans : The rate of zero order reaction is independent of the reactant concentration

Integrated rate law for zero order reaction A $\rightarrow$ P

The differential law is give by

$$\text{rate} = -\frac{d[A]}{[A]} = K[A]^\circ = K \quad \text{..... (i)}$$

By rearrangement above equ^n given $d[A] = -K\,dt$

Integration between the limits

$[A] = [A]_0$ at $t = 0$ and $[A] = [A]_t$ at time t

$$\int_{[A]_0}^{[A]^t} d[A] = -K\int_0^t dt$$

$$[A]_t - [A_0] = -Kt$$

Hence $Kt = [A]_0 - [A]_t$

Section (C)

Q.3 : Answer the following question (any one). (03)

1) In a first order reaction $A \rightarrow B$, 60% of the given sample of compound decomposes in 45 minutes. What is half life of reaction?

Ans : Given,

$[A]_0 = 100\%$

$\chi = 60\%$

$\therefore [A]_t = [A]_0 - \chi = 100 - 60 = 40\%$

$t = 45$ min

Formula :-

$$K = \frac{2.303}{t}\log_{10}\frac{[A]_0}{[A]_t}\theta$$

$$= \frac{2.303}{45} \times \log_{10}\frac{100}{40}$$

$$= \frac{2.303}{45} \times \log_{10} 2.5$$

$$= \frac{2.303}{45} \times 0.3979$$

$$= 0.0204 \text{ min}^{-1}$$

$$\text{Half life} = t_{1/2} = \frac{0.693}{K}$$

$$= \frac{0.693}{0.0204}$$

$$= 34 \text{ min}$$

ii) For the first order reaction $A_{(g)} \rightarrow 2B_{(g)} + C_{(g)}$ the initial pressure of A is 90 mm Hg and pressure after 10 min is found to be 180 mm Hg. Find rate constant of reaction.

Ans :

	A	$\rightarrow$	2B	+	C
$t = 0$	P_0		0		0
t	$P_0 - x$		2x		x

$$P_T = P_0 + 2x$$

$$x = \frac{P_T - P_0}{2}$$

$$x = \frac{180 - 90}{2} = 45$$

Pressure of A at time 10 min

$$P_0 - x = 90 - 45 = 45 \text{ mm Hg}$$

$$K = \frac{2.303}{t}\log\frac{[P_0]}{[P_t]}$$

$$K = \frac{2.303}{10}\log\frac{90}{45}$$

$$= 0.2303 \times 0.3010$$

$$= 0.0693 \text{ min}^{-1}$$

Section (D)

Q.4 : Answer the following question. (Any one) (04)

1) a) Derive integrated rate law for first order reaction.

Ans : Consider a general first order reaction is.

$A \rightarrow$ products

The differential rate law is $-\frac{d[A]}{dt} = K[A]$

where $[A] = [A]_0$ at time t = 0 and $[A] = [A]_t$ at time t = t

Integrating above equation in limits.

$$\int_{[A]_0}^{[A]_t} \frac{d[A]}{A} = -K\int_0^t dt$$

$$\{\ell n\ [A]\}_{[A]_0}^{[A]_t} = -K(t)_0^t$$

$$\ell n\ [A]_t - \ell n\ [A]_0 = -Kt$$

$$\ell n \frac{[A]_t}{[A]_0} = -Kt$$

$$Kt = \ell n \frac{[A]_0}{[A]_t}$$

$$K = \frac{1}{t} \ell n \frac{[A]_0}{[A]_t}$$

$$K = \frac{1}{t} \ell n \frac{[A]_0}{[A]_t}$$

converting ℓn to $\log_{10}$ the integrated rate law becomes.

$$K = \frac{2.303}{t} \log_{10} \frac{[A]_0}{[A]_t}$$

b) Consider the reaction $2N_2O_{5(g)} \rightarrow 4NO_{2(g)} + O_{2(g)}$ in liquid bromine. At a particular moment during the reaction N_2O_5 disappears at a rate 0.02 M/s. At what rates NO_2 and O_2 are formed? what is rate of reaction.

Ans. : $-\frac{d[N_2O_5]}{dt} = 0.02$ m/s.

The reaction shows that the rate of formation of NO_2 is twice the rate of consumption of N_2O_5.

$$\therefore \frac{d[NO_2]}{dt} = -2\frac{d[N_2O_5]}{dt}$$

$$= -2 \times (-0.02\ M/s) = 0.04\ M/s$$

The rate of consumption of N_2O_5 is twice at the rate of formation of O_2

$$-\frac{d[N_2O_5]}{dt} = 2\frac{d[O_2]}{dt} \text{ or}$$

$$\frac{d[O_2]}{dt} = -\frac{1}{2}\frac{d[N_2O_5]}{dt}$$

Hence $\frac{d[O_2]}{dt} = -\frac{1}{2}$ x (-0.02 M/S)

$= 0.01$ M/s

Rate of reaction =

$$= -\frac{1}{2}\frac{d[N_2O_5]}{dt} = \frac{1}{4}\frac{d[NO_2]}{dt} = \frac{d[O_2]}{dt}$$

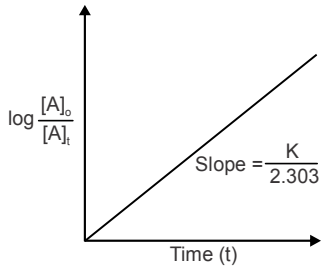
$$= -\frac{1}{2} \text{ x } (-0.02\ M/s)$$

$= -0.01\ M/s.$

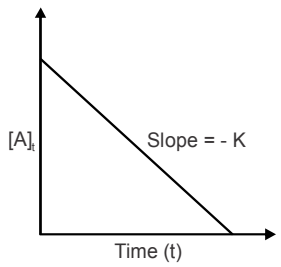
OR

2) a) Draw the following graphs

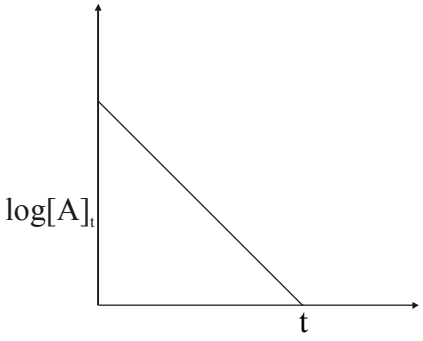
Ans : i) $\log \frac{[A]_o}{[A]_t}$ vs time(t) for first order reaction



ii) $[A]_t$ vs t for zero order reaction.



iii) $\log[A]_t$ vs t for first order



b) The rate of the reaction, A + B $\rightarrow$ **P is 3.6×10^{-2} mol dm^{-3} s^{-1} where**

[A] = 0.2 moles dm^{-3} and [B] = 0.1 moles dm^{-3}.

Calculate the rate constant if the reaction is first order in A and second order in B.

Ans : The reaction is first order in 'A' and second order in 'B'. Hence the rate law gives

rate = K [A][B]^2

$$\text{or } K = \frac{\text{rate}}{[A][B]^2}$$

rate = 3.6×10^{-2} Ms^{-1}

[A] = 0.2 mol dm^{-3} and [B] = 0.1 mol dm^{-3} all value substituting gives

$$K = \frac{3.6\times10^{-2}\text{ mol dm}^{-3}\text{s}^{-1}}{0.2\text{mol dm}^{-3}\times(0.1\text{ mol dm}^{-3})^2}$$

$$= \frac{3.6}{0.2} = 18$$

$$\boxed{K = 18\text{ mol}^{-2}\text{dm}^6\text{ s}^{-1}}$$

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