First order Reaction

A reaction is said to be of first order if its rate is determined by the change of one concentration term only

A $\longrightarrow$ Products

	Time	Concentration
Initial	T =0	a
Final	T = t	(a -x)

$\therefore \quad \frac{dx}{dt} = K(a - x)$

On integrating $\int \frac{dx}{a-x} = K \int dt$ we get $-\log_e (a-x) = Kt + C$

$\because$ at t = 0 x = 0 $\quad \therefore$ C = $-\log_e a$

$\therefore -\log_e (a - x) = Kt - \log_e a$

or $Kt = log_e \frac{a}{a-x}$

Rate equation for first order reaction: $\boldsymbol{Kt = 2.303\ log\frac{a}{a-x}}$

Here $[A]_0 = a\ (Initial\ Concentration$

$[A]_0 = (a - x)\ (Remaning\ concentration)$

By rearranging the equation, we get

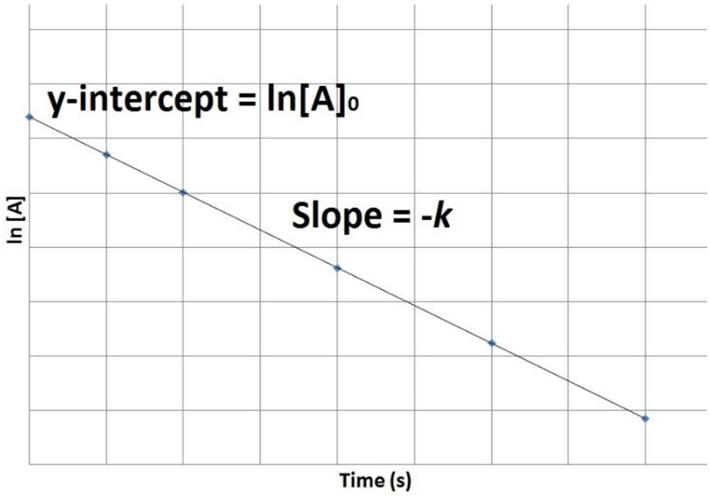
$$\frac{kt}{2.303} = log\frac{[A]_0}{[A]_0}$$

$$\text{or } \frac{kt}{2.303} = log[A]_0 - log[A]$$

$$\text{or} \qquad log[A] = \frac{-kt}{2.303} + log[A]_o$$

The graph between log[A] and time ‘t’ is a straight line, and the slope of this line is given by

Slope $= -\frac{k}{2.303}$.



Half-life of a reaction: Half-life of first order reaction does not depend upon initial concentration of the reactants.

For 1st order reaction: k = $\frac{2.303}{tk} log \frac{[A]_o}{[A]}$

From the above expression we get: t = $\frac{2.303}{k} log \frac{[A]_0}{[A]}$

when t = $t_{1/2}$ then [A] = $\frac{1}{2}[A]_o$

therefore $t_{1/2} = \frac{2.303}{k} log \frac{[A]_o}{\frac{1}{2}[A]_o} = \frac{2.303}{k} log\, 2$

or $t_{1/2} = \frac{0.693}{k}$

Units of rate constant for reactions of different orders

(i) For zero order reactions

Rate =$\frac{dx}{dt} = k[A]^o$; k = mol L^{-1} sec^{-1}.

(ii) For first order reactions

Rate = k[A] =$\frac{dx}{dt}$; hence k.[mol L^{-1}] = $\left(\frac{mol}{L-sec}\right)$

∴ k = (sec^{-1}).

(iii) For second order reactions

Rate =$k[A]^2 = \frac{dx}{dt}$; $k\left(\frac{mol}{L}\right)^2 = \left(\frac{mol}{L-sec}()\right)$

Hence k = L $mol^{-1}sec^{-1}$.