

CHAIN REACTION

Presented by :

Arvind Singh Heer

MSc-I

(Sem-II)

Inorganic Chemistry

MITHIBAI COLLEGE

CONTENTS

- INTRODUCTION.
- CHAIN REACTION
- REFERENCES

INTRODUCTION

- **Defination:-**A chemical or nuclear reaction which proceed through a sequence (chain) of self repeating steps initiated by a suitable primary process is called a chain reaction.
- **Types of Chain Reactions:**
 - Stationary Chain Reaction
 - Non-stationary Chain Reaction

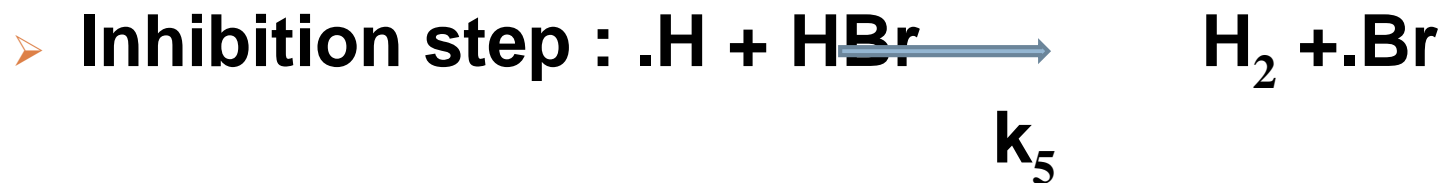
CHAIN REACTION

- Example : Reaction between hydrogen and bromine



The reaction follows simple elementary reaction.

- $$\text{Br}_2 \xrightarrow{k_1} \cdot\text{Br} + \cdot\text{Br}$$
- $$\cdot\text{Br} + \text{H}_2 \xrightarrow{k_2} \text{HBr} + \cdot\text{H}$$
- $$\cdot\text{H} + \text{Br}_2 \xrightarrow{k_3} \text{HBr} + \cdot\text{Br}$$



➤ The steady-state hypothesis applied to the two intermediates Br & H , both of which are present at very low concentrations. The steady-state equation for H is

$$\frac{d[\text{.H}]}{dt} = k_2 [\text{.Br}][\text{H}_2] - k_3 [\text{.H}][\text{Br}_2] - k_4 [\text{.H}][\text{HBr}] = 0$$

.....(1)

$$\frac{d[.Br]}{dt} = k_1 [Br_2] - k_2 [.Br][H_2] + k_3 [.H][Br_2] + k_4 [.H][HBr] - k_5 [.Br]^2 = 0 \dots\dots(2)$$

Adding (1) & (2)

$$k_1 [Br_2] - k_5 [.Br]^2 = 0$$

$$k_1 [Br_2] = k_5 [.Br]^2$$

$$[.Br] = (k_1/k_5 [Br_2])^{1/2} \dots\dots\dots(3)$$

from equation no. (1), we get

$$k_2 [.Br][H_2] = k_3 [.H][Br_2] + k_4 [.H][HBr]$$

$$k_2 [.Br][H_2] = [.H] (k_3 [Br_2] + k_4 [HBr])$$

$$[.H] = \frac{k_2 [.Br][H_2]}{k_3 [Br_2] + k_4 [HBr]}$$

$$[.H] = k_2 (k_1/k_5)^{1/2} \frac{[H_2][Br_2]^{1/2}}{k_3 [Br_2] + k_4 [HBr]} \dots\dots\dots(4)$$

HBr is formed in reaction (2) & disappears in reaction (3). The net rate of formation of HBr gives

$$\frac{d[HBr]}{dt} = k_2 [.Br][H_2] + k_3 [.H][Br_2] - k_4 [.H][HBr] \dots\dots\dots(5)$$

from eq. (1) we know that,

$$k_3 [.H][Br_2] = k_2 [.Br][H_2] - k_4 [.H][HBr]$$

Eq.(5) becomes ,

$$\begin{aligned}
 \frac{d[\text{HBr}]}{dt} &= k_3 [\cdot\text{H}][\text{Br}_2] + k_2 [\cdot\text{Br}][\text{H}_2] - k_4 [\cdot\text{H}][\text{HBr}] \\
 &= k_3 [\cdot\text{H}][\text{Br}_2] + k_3 [\cdot\text{H}][\text{Br}_2] \\
 &= 2k_3 [\cdot\text{H}][\text{Br}_2] \dots\dots\dots(6) \\
 2k_3 [\cdot\text{H}][\text{Br}_2] &= \frac{d[\text{HBr}]}{dt}
 \end{aligned}$$

$$[\cdot\text{H}] = \frac{1}{2k_3 [\text{Br}_2]} \frac{d[\text{HBr}]}{dt}$$

substituting conc.H radical in eq.(4)

$$\frac{1}{2k_3[\text{Br}_2]} \frac{d[\text{HBr}]}{dt} = \frac{k_2(k_1/k_5)^{1/2} [\text{H}_2][\text{Br}_2]^{1/2}}{k_3[\text{Br}_2] + k_4[\text{HBr}]}$$

$$\frac{d[\text{HBr}]}{dt} = \frac{2k_3 k_2(k_1/k_5)^{1/2} [\text{H}_2][\text{Br}_2]^{1/2}}{k_3[\text{Br}_2] + k_4[\text{HBr}]}$$

Dividing both numerator & denominator of above equation by $k_3[\text{Br}_2]$

$$\frac{d[\text{HBr}]}{dt} = \frac{2k_2(k_1/k_5)^{1/2} \cdot [\text{H}_2][\text{Br}_2]^{1/2}}{1 + \frac{k_4[\text{HBr}]}{k_3[\text{Br}_2]}}$$

$$\frac{d[\text{HBr}]}{dt} = \frac{K [\text{H}_2][\text{Br}_2]^{1/2}}{1 + k'[\text{HBr}]/[\text{Br}_2]}$$

$$\text{where } K = 2k_2(k_1/k_5)^{1/2}$$

$$k' = k_4/k_3$$

In the initial state of the reaction order of reaction is 1.5

At the initial stage of the reaction HBr is negligibly small & hence

$$1 + \frac{k'[\text{HBr}]}{[\text{Br}_2]} \approx 1$$

and initial rate becomes,

$$\frac{d[\text{HBr}]}{dt} = K [\text{H}_2][\text{Br}_2]^{1/2}$$

This is the rate law equation for given reaction.

REFERENCES

- LAIDLER K.J AND MEISER J. H. PHYSICAL CHEMISTRY THIRD OXFORD UNIVERSITY PRESS

-THANK YOU