

Kinetics

Chapter 12

KINETIC ENERGY

The energy an object has because of its motion, involving the movement of the object over a specific period of time.

$$KE = \frac{1}{2} mV^2$$

m = mass (kg)

V = Velocity (m/s)

RATE EXPRESSION

the change in concentration of reactants or products over time in a chemical reaction.

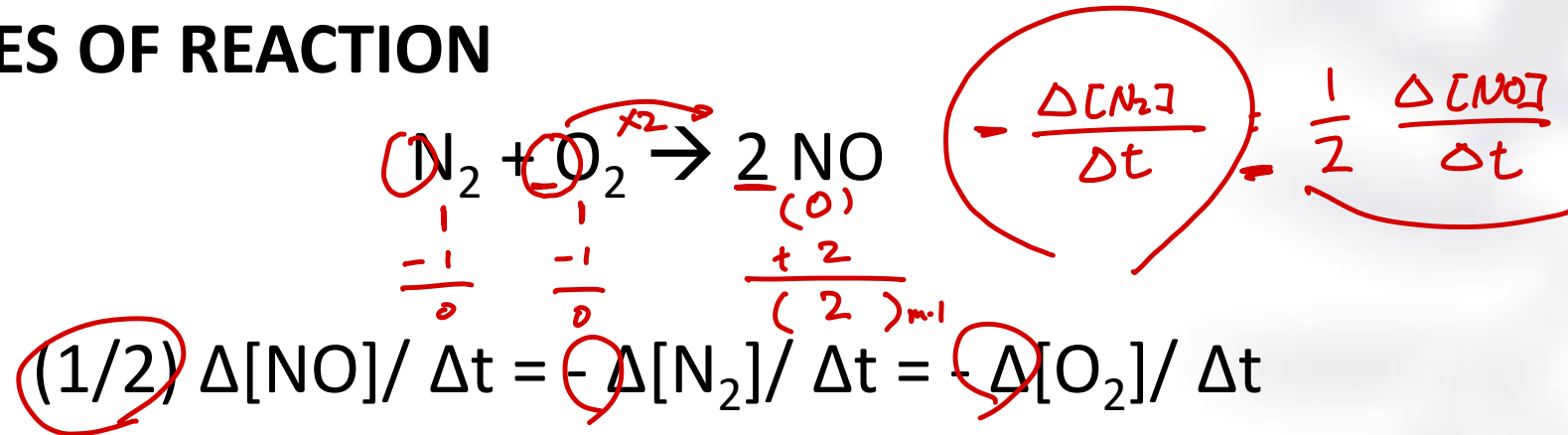


$$\begin{aligned}\text{Rate of decomposition of H}_2\text{O}_2 &= - \frac{\text{change in concentration of reactant}}{\text{time interval}} \\ &= - \frac{[\text{H}_2\text{O}_2]_{t_2} - [\text{H}_2\text{O}_2]_{t_1}}{t_2 - t_1} \\ &= - \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t}\end{aligned}$$

Δ = “change in” or “final – initial”

The bracket [] indicates molarity (mol/L)

RELATIVE RATES OF REACTION



Due to the coefficient (mol ratio), every 1 mol of the reactant (N_2 and O_2) spent, only $\frac{1}{2}$ of the product (NO) is produced

$$\frac{1}{\text{Coefl.}} \frac{\Delta[\]}{\Delta t}$$

The concentration of N_2 and O_2 decrease at half the rate at which the concentration of NO increases.

Ex) Expressions for Relative Reaction Rates

$$\frac{\Delta [\text{NH}_3]}{\Delta t} = -10 \text{ M/s}$$

The first step in the production of nitric acid is the combustion of ammonia:



Write the equations that relate the rates of consumption of the reactants and the rates of formation of the products.

$$\frac{1}{4} \frac{\Delta [\text{NH}_3]}{\Delta t} = -\frac{1}{5} \frac{\Delta [\text{O}_2]}{\Delta t} = \frac{1}{4} \frac{\Delta [\text{NO}]}{\Delta t} = \frac{1}{6} \frac{\Delta [\text{H}_2\text{O}]}{\Delta t}$$

$$5 \times \frac{10}{4} = \frac{4}{5} \times$$

$$\frac{50}{4} = \underline{12.5}$$

$$-\frac{\Delta [\text{NH}_3]}{\Delta t} = -\frac{4}{5} \frac{\Delta [\text{O}_2]}{\Delta t} = \frac{\Delta [\text{NO}]}{\Delta t} = \frac{4}{6} \frac{\Delta [\text{H}_2\text{O}]}{\Delta t}$$

10 = 10



$$-\frac{1}{2} \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = \frac{1}{1} \frac{\Delta[\text{O}_2]}{\Delta t}$$

The instantaneous rate of decomposition of H_2O_2 at $t = 11.1$ h is determined to be $3.20 \times 10^{-2} \frac{\text{mol}}{\text{L} \cdot \text{h}}$. That is:

$$\text{H}_2\text{O} = 3.2 \times 10^{-2} \text{ M/h}$$

$$\text{O}_2 = 0.016 \text{ M/h}$$

$$-\frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} = 3.20 \times 10^{-2} \frac{\text{mol}}{\text{L} \cdot \text{h}}$$

What is the instantaneous rate of production of H_2O and O_2 ?

$$-\frac{1}{2} (3.20 \times 10^{-2} \text{ M/h}) = \frac{1}{2} (\text{H}_2\text{O rate})$$

$$-\frac{1}{2} \left(\frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} \right) = 1 \frac{\Delta[\text{O}_2]}{\Delta t}$$

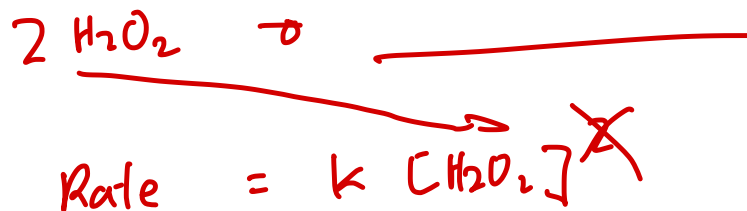
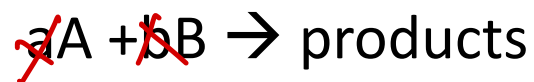
$$-\frac{1}{2} (3.2 \times 10^{-2} \text{ M/h}) = \text{rate of } \text{O}_2$$

rate

RATE LAW

Or differential rate laws or rate equations

Mathematical expressions that describe the relationship between the rate of a chemical reaction and the concentration of its reactants.



$$\text{Rate} = k[\text{A}]^m[\text{B}]^n$$

K = rate constant

m,n = reaction order

Note:

Unlike the exponents in a chemical equilibrium equation, the reaction orders of the reactants are not related to their coefficients in a balanced chemical equation. It must be determined experimentally.

METHOD OF RATE LAW DETERMINATION

- Method of initial Rates

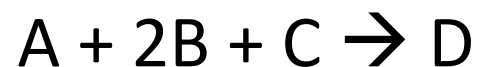
Determines the rate law by comparing initial reaction rates at different reactant concentrations to see how changes in concentration affect the rate.

- Using integrated Rate Laws

Determines the rate law by analyzing how the concentration of reactants changes over time, fitting this data to integrated rate equations to find the reaction order and rate constant.

METHOD OF INITIAL RATES

Ex)



Experiment	Initial Concentration of Reactants (M)			Initial Rates of Formation of D (M/s)
	[A]	[B]	[C]	
1	0.10	0.10	0.10	0.01
2	0.10	0.10	0.20	0.01
3	0.10	0.20	0.10	0.02
4	0.20	0.20	0.10	0.08

$$\text{Rate} = k[A]^x[B]^y[C]^z$$

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Let's look at [A]

$$(2)^x = 4 \text{ or } x = 2$$

$$\text{rate} = k[A]^x[B]^y[C]^z$$

$\times 4$ $\times 2$ $\times 2$

Experiment 3 & 4

Why? [A] changes while the other reactant concentrations remain constant. For this reason, the rate change is solely based on [A]

$$\text{Rate} = k[A]^2[B]^y[C]^z$$

When the concentration doubles
(0.10 → 0.20),
the rate quadruples (0.02 → 0.08)