MASTERMIND SCHOLARS EDUCATIONAL ALLIANCE

CHEMISTRY

RATE OF CHEMICAL REACTION

Rate of chemical reaction is the amount of reactant consumed per unit time.

Or

Rate of chemical reaction is the amount of product formed per unit time.

Or

Rate is the change in concentration of a substance (reactants and products) with time.

Expression to show rates

Consider a reaction

$$A \longrightarrow B$$

A is reactants (it is consumed during the reaction)

B is the product (it is formed during the reaction)

Rate of disappearance of A

$$R = -\frac{d[A]}{dt}$$

(The negative sign indicates that the concentration of the reactant is decreasing/consumed as the reaction proceeds.)

Rate of appearance of B

$$R = +\frac{d[B]}{dt}$$

(The positive sign shows that the concentrations of the products are increasing as the reaction proceeds.)

Note that:

All the rates are equal in magnitude :

$$R = -\frac{d[A]}{dt} \quad = + \frac{d[B]}{dt}$$

This means the reaction occurs at the same overall rate, even though the reactants decrease and the products increase as the reaction continues.

Expressing Reaction Rates Using Unitary Mole Ratios

When all coefficients in a balanced chemical equation are **unity (1)**, the rates of consumption and formation of substances can be directly compared.
In this case, each rate is technically **divided by 1**, but since dividing by 1 does not change the value, there is **no need for further division or scaling**.

$$A + B \longrightarrow C + D$$

A and B are the reactants (they are consumed during the reaction)

C and D are the products (they are formed during the reaction)

Rate of Reactant A

$R = -\frac{d[A]}{dt}$

Rate of Reactant B

$R = -\frac{d[B]}{dt}$

(The negative sign indicates that the concentrations of the reactants are decreasing/consumed as the reaction proceeds.)

Rate of Product C

Rate of Product D

$R = \frac{d[C]}{dt}$ $\qquad\qquad$ $R = \frac{d[D]}{dt}$

(The positive sign shows that the concentrations of the products are increasing as the reaction proceeds.)

Note that :

$$R = -\frac{d[A]}{dt} = -\frac{d[B]}{dt} = \frac{d[C]}{dt} = \frac{d[D]}{dt}$$

This means the reaction proceeds at the same overall rate, even though the reactants decrease and the products increase as the reaction continues.

Relating Reaction Rates to Stoichiometric Coefficients

When expressing the rate of a reaction, the coefficients in the balanced chemical equation are used to relate the rates of consumption of reactants and formation of products.

$$2A + 3B \longrightarrow 5C + 6D$$

A and B are the reactants (they are consumed during the reaction)

C and D are the product (they are formed during the reaction)

To compare the rates of change of each substance, divide the rate of change of concentration by its respective stoichiometric coefficient (moles).

Rates of reactants

Rate of Reactant A

$R = -\frac{d[A]}{dt} \div 2$

$R = -\frac{d[A]}{dt} \times \frac{1}{2}$

Rate of Reactant B

$R = -\frac{d[B]}{dt} \div 3$

$R = -\frac{d[B]}{dt} \times \frac{1}{3}$

$R = -\frac{1}{2}\frac{d[A]}{dt}$

$R = -\frac{1}{3}\frac{d[B]}{dt}$

(The negative sign indicates that the concentrations of the reactants are decreasing as the reaction proceeds.)

Rates of products

Rate of Product C

$R = \frac{d[C]}{dt} \div 5$

$R = \frac{d[C]}{dt} \times \frac{1}{5}$

$R = \frac{1}{5}\frac{d[C]}{dt}$

Rate of Product D

$R = \frac{d[D]}{dt} \div 6$

$R = \frac{d[D]}{dt} \times \frac{1}{6}$

$R = \frac{1}{6}\frac{d[D]}{dt}$

(The positive sign shows that the concentrations of the products are increasing as the reaction proceeds.)

Hence:

All the rates are equal in magnitude when properly expressed with their coefficients:

$$R = -\frac{1}{2}\frac{d[A]}{dt} = -\frac{1}{3}\frac{d[B]}{dt} = \frac{1d[C]}{5dt} = \frac{1d[D]}{6dt}$$

This means the reaction proceeds at the same overall rate, even though the reactants decrease and the products increase as the reaction continues.

Relating Reaction Rates Using Fractional (Scaled) Coefficients

To express the rate of reaction for each substance in a balanced chemical equation with fractional coefficients, divide the rate of change of its concentration by its fractional stoichiometric coefficient.

$$\frac{1}{2}\text{A} + \frac{1}{3}\text{B} \longrightarrow \frac{1}{4}\text{C} + \frac{1}{5}\text{D}$$

A and B are the reactants (they are consumed during the reaction)
C and D are the product (they are formed during the reaction)

Rate of Reactant A

$$R = -\frac{d[A]}{dt} \div \frac{1}{2}$$

$$R = -\frac{d[A]}{dt} \times 2$$

$$R = -2\frac{d[A]}{2dt}$$

Rate of Reactant B

$$R = -\frac{d[B]}{dt} \div \frac{1}{3}$$

$$R = -\frac{d[B]}{dt} \times 3$$

$$R = -3\frac{d[B]}{dt}$$

(The negative sign indicates that the concentrations of the reactants are decreasing as the reaction proceeds.)

Rate of Product C

$$R = \frac{d[A]}{dt} \div \frac{1}{4}$$

$$R = \frac{d[A]}{dt} \times 4$$

$$R = 4\frac{d[A]}{2dt}$$

Rate of Product D

$$R = \frac{d[B]}{dt} \div \frac{1}{5}$$

$$R = \frac{d[B]}{dt} \times 5$$

$$R = 5\frac{d[B]}{dt}$$

(The positive sign shows that the concentrations of the products are increasing as the reaction proceeds.)

Hence

$$R = -2\frac{d(A)}{2dt} = -3\frac{d(B)}{dt} = 4\frac{d(A)}{2dt} = 5\frac{d(B)}{dt}$$

This means the reaction proceeds at the same overall rate, even though the reactants decrease and the products increase as the reaction continues.

EXAMPLE 1

Consider the following chemical reactions

a. $2N_2O_{5\,(g)} \longrightarrow 4NO_{2\,(g)} + O_{2\,(g)}$

b. $3A + \frac{1}{2}B \longrightarrow 2C$

c. $\frac{1}{2}N_2 + \frac{2}{3}H_2 \longrightarrow NH_3$

1. Write down all the rate of reaction for each of the reactants and products

SOLUTION

a. $2N_2O_{5\,(g)} \longrightarrow 4NO_{2\,(g)} + O_{2\,(g)}$

Rate of reactant (dinitrogen pentaoxide)

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

Rate of Product ($NO_{2\,(g)}$)

$$R = \frac{1}{4}\frac{d[NO_2]}{dt}$$

Rate of product

$$R = \frac{\mathrm{d}[O_2]}{dt}$$

b.

$$3A + \frac{1}{2}B \longrightarrow 2C$$

Solution

Rate of Reactants

$$R = -\frac{1}{3}\frac{\mathrm{d}[A]}{dt}$$

Rate of Reactants

$$R = -2\,\frac{\mathrm{d}[B]}{dt}$$

Rate of Products

$$R = \frac{1}{2}\frac{\mathrm{d}[C]}{dt}$$

c. $\frac{1}{2}N_2 + \frac{2}{3}H_2 \longrightarrow NH_3$

Solution

Rate of Reactants

$$R = -2\frac{\mathrm{d}[N_2]}{dt}$$

Rate of Reactants

$$R = -\frac{3\mathrm{d}[H_2]}{2\,dt}$$

Rate of Products

$$R = \frac{\mathrm{d}[NH_3]}{dt}$$

TRY WORK 1

Consider the following chemical reactions

a. $2N_2O_{5\,(g)} \longrightarrow 4\mathrm{N}O_{2\,(g)} + O_{2\,(g)}$

b. $3\mathrm{A} + \frac{1}{2}\mathrm{B} \longrightarrow 2\mathrm{C}$

c. $\frac{1}{2}N_2 + \frac{2}{3}H_2 \longrightarrow \mathrm{N}H_3$

i. Express the rate of reaction for each reactant and product.

ii. For each reaction, state which substances **increase** and which **decrease** in concentration.

iii. For reaction (a), verify that the rates of all reactants and products are **equal in magnitude** when properly scaled.

iv. For reaction (b), if the rate of A ; $\frac{d[A]}{dt}$ = 0.06 mol/dm^3/s calculate:

a) The rate of B; $\frac{d[B]}{dt}$

b) The rate of C ; $\frac{d[c]}{dt}$

Hint :

$$2\frac{d[B]}{dt}=\frac{1}{3}\frac{d[A]}{dt}$$

$$2\frac{d[B]}{dt}=\frac{1}{3}(0.06)$$

$$\frac{d[B]}{dt} = \frac{0.02}{2}$$

$$\frac{d[B]}{dt} = 0.01 \text{ mol/dm}^3\text{/s}$$

v. For reaction (c), if the rate of formation of NH_3 is 0.12 mol dm^{-1} s^{-1} determine the rates of consumption of NH_3 and H_2

vi. Explain why all the rates of change, when adjusted for their stoichiometric coefficients, **represent the same reaction rate**

QUESTION ONE

Consider the reaction 3A + $\frac{1}{2}B$ ⟶ $2C$.

If the rate of disappearance of B is 1.5×10^{-2} mol/dm^3/s Determine

a. The rate of disappearance of A.

b. The rate of formation of C

Solution

Rate of reactants (A)

$$R = -\frac{1}{3}\frac{d[A]}{dt}$$

Rate of reactants (B)

$$R = -2\frac{d[B]}{dt}$$

Rate of Product (C)

$$R = \frac{1}{2}\frac{d[C]}{dt}$$

a. The rate of disappearance of A

$$2\frac{d[B]}{dt} = \frac{1}{3}\frac{d[A]}{dt}$$

$$2\times(1.5 \times 10^{-2}) = \frac{1d[A]}{3dt}$$

$$0.03 = \frac{1d[A]}{3dt}$$

$R_A = 0.03 \times 3$

$R_A = 0.09$

$R_A = 9.0 \times 10^{-2} moldm^{-3}s^{-1}$

b. The rate of formation of C

$$\frac{2d[B]}{3dt} = \frac{1d[C]}{2dt}$$

$$2\times(1.5 \times 10^{-2}) = \frac{1d[C]}{2dt}$$

$$0.03 = \frac{1d(C)}{2dt}$$

$R_c = 0.03 \times 2$

$R_c = 0.06$

$R_c = 6.0 \times 10^{-2} mol\ dm^{-3}s^{-1}$

QUESTION 2

Consider the reaction $\frac{1}{2}N_2 + \frac{2}{3}H_2 \longrightarrow NH_3$

If the rate of disappearance of hydrogen is 4.5 x $\mathbf{10}^{-3}$ mol/dm^3/s, determine the rate of formation of NH_3

Solution

Rate of Reactants

$$R = -\frac{2\ d[N_2]}{dt}$$

Rate of Reactants

$$R = -\frac{3\mathrm{d}[H_2]}{2\,dt}$$

Rate of Products

$$R = \frac{\mathrm{d}[NH_3]}{dt}$$

$$\frac{3\mathrm{d}[H_2]}{2\,dt} = \frac{\mathrm{d}[NH_3]}{dt}$$

$$\frac{3}{2} \times (4.5 \times 10^{-3}) = \frac{\mathrm{d}[NH_3]}{dt}$$

$R_{NH3} = 6.75 \times 10^{-3}$ mol dm^{-3}s^{-1}

FACTOR THAT AFFECT RATE OF REACTION

1. Temperature
2. Concentration / pressure
3. Catalyst
4. Surface Area/ particle size/ physical state
5. Medium

Collision Theory

For a chemical reaction to occur, reacting molecules must collide with one another.

Not all collisions result in a reaction; only certain collisions are successful.

Effective collisions are those collisions that lead to a chemical reaction. Only a small fraction of all collisions are effective.

To overcome the energy barrier (activation energy), only the most energetic molecules can participate in a reaction.

The orientation of molecules during collision is important; molecules must collide in the correct alignment (e.g., head-on) to react successfully.

Surface Area

If one of the reactants is a solid, increasing its surface area — for example, by crushing it into smaller pieces or grinding it into a powder — exposes more particles to the other reactant. This makes more particles available for collision, increasing the number of collisions per second and, consequently, the number of effective collisions. Making most of the particle attain energy equal to or greater than the activation energy. As a result, the rate of reaction increases.

Catalyst

A catalyst is a substance that increases the rate of a chemical reaction without being chemically changed or used up.

It works by providing an **alternative reaction pathway that has a lower activation energy** than the uncatalysed reaction.

Because the activation energy is reduced, a greater proportion of reacting molecules now have enough energy to collide effectively and form products.

This increases the number of effective collisions per second, thereby increasing the rate of reaction.

Example:

Manganese(IV) oxide (MnO_2) acts as a catalyst in the decomposition of hydrogen peroxide (H_2O_2) into water and oxygen gas.

A catalyst provides an alternative reaction pathway with a lower activation energy.

Because of this, a greater proportion of reacting molecules now have enough energy to overcome the activation barrier during collisions.

This results in more effective collisions per second, thereby increasing the rate of reaction without the catalyst being consumed in the process.

EXAMPLES OF CATALYST IN VARIOUS REACTIONS

1. Inorganic Reactions

(a) *Decomposition of Hydrogen Peroxide:*

$$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$$

Catalyst: Manganese(IV) oxide (**MnO_2**)
Effect: The catalyst speeds up the breakdown of hydrogen peroxide into water and oxygen without being consumed.

(b) *Decomposition of Potassium Chlorate:*

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

Catalyst: Manganese(IV) oxide (**MnO_2**)
Effect: Reduces the activation energy, allowing oxygen to be released faster on gentle heating.

2. Organic Reactions

(a) *Hydrogenation of Ethene (Addition Reaction):*

$$C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_6(g)$$

Catalyst: Nickel (**Ni**) or Platinum (**Pt**)
Effect: The metal surface provides an alternative pathway for hydrogen and ethene molecules to react at a lower activation energy.

(b) *Esterification (Reversible Reaction):*

$$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$$

Catalyst: Concentrated Sulphuric acid (**H_2SO_4**)
Effect: Acts as an acid catalyst, speeding up both forward and reverse reactions.

3. Industrial Processes

(a) *Haber Process (Ammonia Production):*

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

Catalyst: Finely divided Iron (**Fe**)
Effect: Speeds up the rate of ammonia formation without altering equilibrium yield.

(b) *Contact Process (Sulphuric Acid Production):*

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

Catalyst: Vanadium(V) oxide (**V_2O_5**)
Effect: Increases rate of oxidation of sulphur dioxide to sulphur trioxide.

(c) *Ostwald Process (Nitric Acid Production):*

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

Catalyst: Platinum-Rhodium gauze (**Pt/Rh**)
Effect: Enables rapid oxidation of ammonia at a lower temperature.

A positive catalyst increases the rate of a chemical reaction by lowering the activation energy.

It provides an alternative reaction pathway with a lower energy barrier, allowing more reacting molecules to have enough energy to form products.

As a result, the reaction occurs faster.

On an energy profile diagram, the new pathway created by the catalyst appears below the original pathway, showing that less energy is required for the reaction to proceed.

Effect of a Negative Catalyst (Inhibitor)

A negative catalyst, also called an inhibitor, slows down the rate of a chemical reaction by increasing the activation energy required for the reaction to occur.
It interferes with the normal reaction pathway, making it more difficult for reactant molecules to collide with sufficient energy to form products.

As a result, fewer effective collisions occur per second, and the reaction proceeds more slowly.

On an energy profile diagram, the new pathway created by the negative catalyst lies above the original pathway, showing that more energy is required for the reaction to take place.

Example:
The presence of glycerol acts as a negative catalyst in the decomposition of hydrogen peroxide (H_2O_2), slowing down the reaction.

Effect of temperature on rates chemical reactions

Increasing temperature increases the kinetic energy of the molecules, leading to more frequent and energetic collisions. This increases the number of effective collisions per second, and hence, the rate of reaction increases.

Effect of concentration on rates of chemical reactions

An increase in concentration of a reactant increases the rate of reaction because it increases the number of particles or molecules present per unit volume. This increases the number of collisions and hence the effective collisions. So the reaction occurs faster.

Types of Reaction

Exothermic reaction

Endothermic reaction

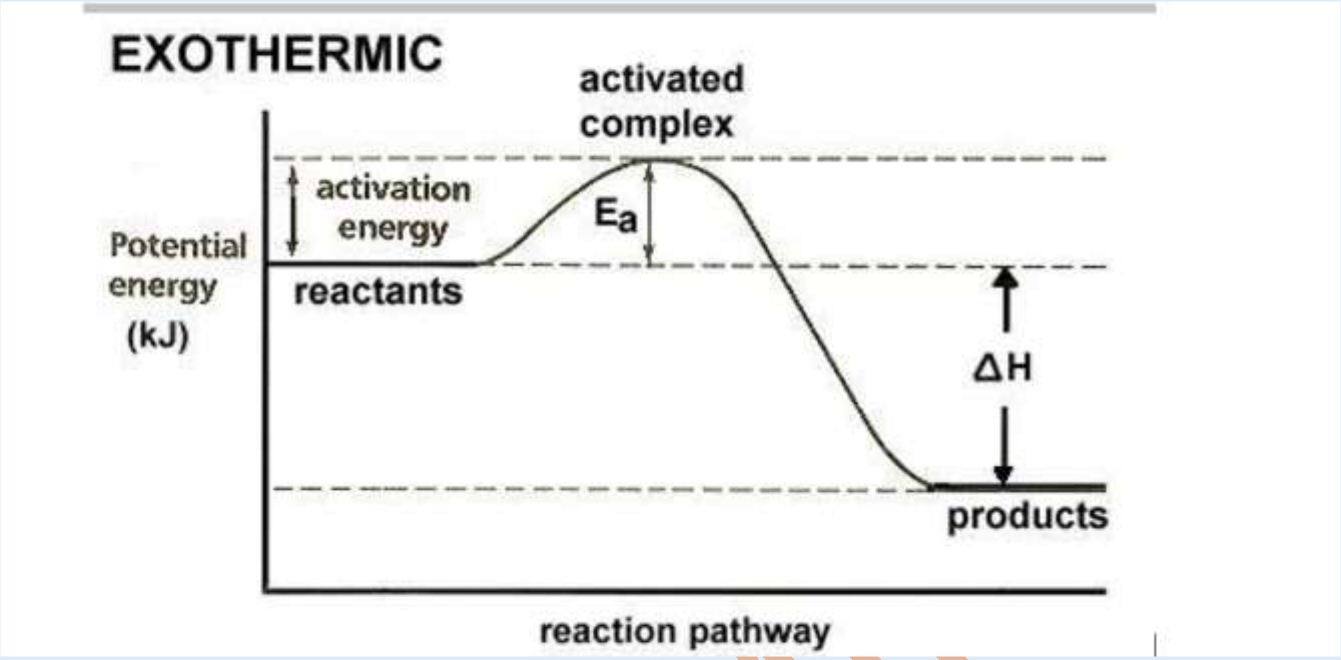
Exothermic reaction

Exothermic reaction is a type reaction in which heat is released into the surrounding.

The enthalpy change is negative

Energy is given out

Feature	Description
Heat flow	Energy is released to surroundings
Temperature change	Surroundings get *hotter*
Energy of products	Lower than reactants
ΔH value	Negative
Examples	Burning, respiration, neutralization



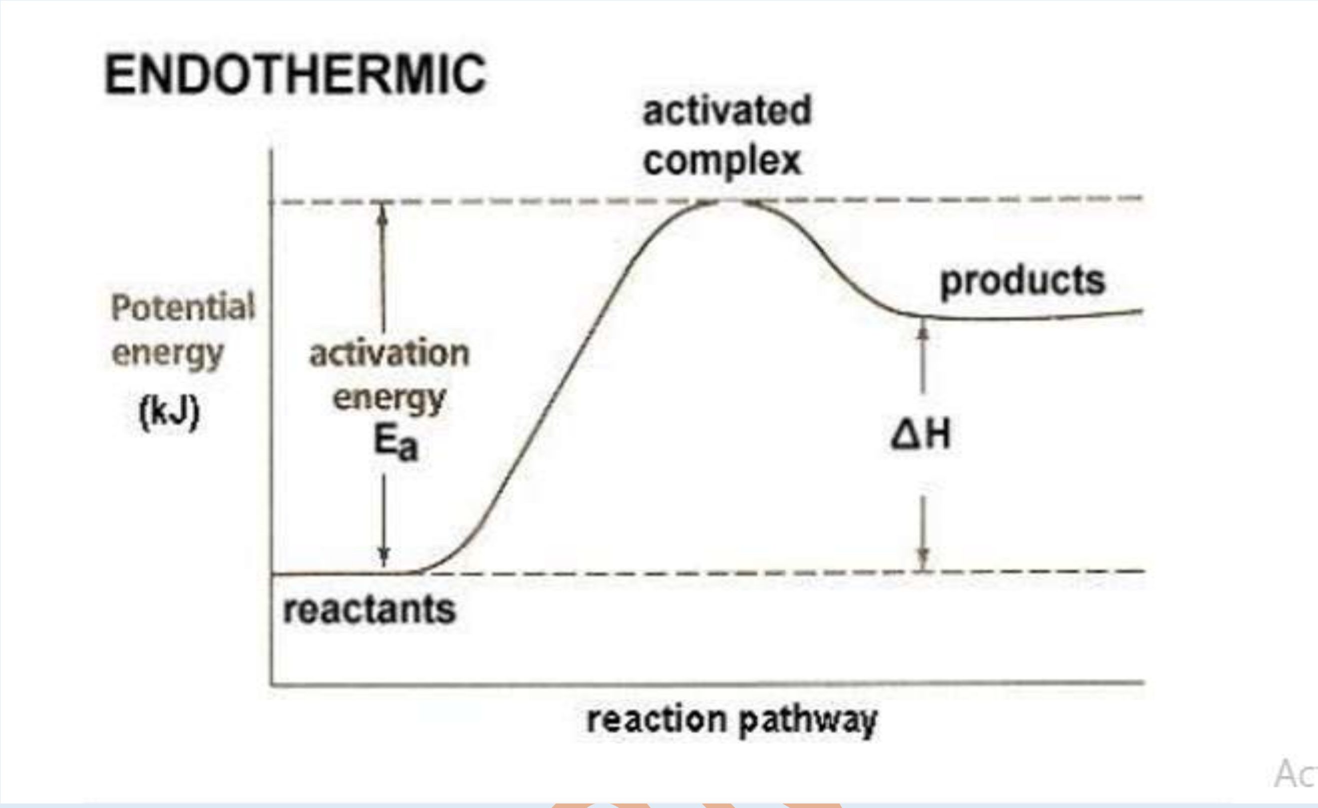
Endothermic reaction

Endothermic reaction is a type of chemical reaction in which heat is absorbed into the surrounding.

The enthalpy change is positive

Energy is absorbed

Feature	Description
Heat flow	Energy is absorbed from surroundings
Temperature change	Surroundings get colder
Energy of products	Higher than reactants
ΔH value	Positive
Examples	Photosynthesis, melting ice, evaporation, thermal decomposition



The Rate law

The rate law is an equation or expression that relates the **rates of a reaction** to the **rate constant** and the **concentrations of the reactants** raised to their **corresponding powers.**

Consider this equation

$$xA + yB \rightarrow C + D$$

Rate law

$$R = k[A]^x[B]^y$$

k = rate constant and is the proportionality constant in every rate law

A and B = concentrations of reactant A and reactant B in mol/dm^3

x and y = order of reaction with respect to A and B

Order of a Reaction

Order of a reaction is the sum of the powers to which the individual concentrations of the reactants are raised in the rate law.

Let consider this equation

$$xA + yB \qquad C + D$$

The order of the reaction with respect to A = x

The order of the reaction with respect to B = y

The overall order of the reaction is = $x + y$

Types of Reaction Orders and Their Implications

1. First-Order Reaction (Order = 1)
The rate depends on the concentration of **one reactant raised to the power 1.**
Rate law: Rate $\propto$ [A]
Implication: If concentration doubles, the reaction rate also doubles.
Example: Radioactive decay, many simple decomposition reactions.

2. Second-Order Reaction (Order = 2)
The rate depends on either:

the **square** of one reactant's concentration, or
the **product** of two different reactant concentrations.

Rate law: Rate $\propto [A]^2$ **or** Rate $\propto$ [A][B]
Implication: If concentration doubles, rate becomes four times faster (because $2^2 = 4$).

3. Third-Order Reaction (Order = 3)

The rate depends on the concentrations of reactants whose **powers add up to 3**.
Rate law: Rate ∝ $[A]^3$ or $[A]^2[B]$ or $[A][B][C]$
Implication: A third-order reaction is extremely sensitive to changes in concentration.
If the concentration of a reactant is doubled, the reaction rate increases by a factor of 8 (because $2^3 = 8$).

The order of a reaction helps to indicate how the rate of a reaction depends on the changes of the concentration of that reactant.

Rate Determining Step

Rate determining step is the slowest of the reaction that takes place in more than one step and help to find the order of a reaction.

Questions 1

At a certain temperature, an alkylbromide reacts with hydroxide ions in a one step process as follows :

$$RBr + OH^- \rightarrow ROH + Br^-$$

i. write down the rate equation for the reaction and give the units of the rate constant.

ii. What is the order of the reaction?

iii. What will happen to the rate of the reaction when the concentration of the hydroxide ion is tripled?

Solution

i. R = k$[RBr][\mathrm{OH}]$

unit of R, RBr, OH

R = mol/dm^3/s

RBr = mol/dm^3

OH= mol/dm^3

$$\text{mol}/dm^3/s = \text{k}\,[\mathrm{mol}/dm^3][\mathrm{mol}/dm^3]$$

$$\text{k} = \frac{\mathrm{mol}dm^{-3}s^{-1}}{[\mathrm{mol}dm^{-3}][\mathrm{mol}dm^{-3}]}$$

$$\text{k} = \frac{\mathrm{mol}dm^{-3}s^{-1}}{[mol^2dm^{-6}]}$$

$$\text{k} = \frac{dm^3s^{-1}}{[mol]}$$

$$\text{k} = mol^{-1}dm^3s^{-1}$$

Question 2

In a certain reaction the rate determining step is

A + 2B → C + D

i write an expression for the reaction rate, giving reasons

ii. Deduce what will be the effect on the reaction rate when

a. Concentration of A is doubled

b. When concentration of B is doubled

iii. how would the concentration of B be changed in order to halve the reaction rate ?

Solution

$$Rate = R = k[A][B]^2$$

The reason for this rate equation is because the reaction is the rate determining step so the orders of the reactants are equal to the stoichiometric coefficients in the reaction.

ii. Effect on the rate when Concentration of A is doubled

$$R_1 = k[2A][B]^2$$

$$R_1 = 2k[A][B]^2$$

But $R = k[A][B]^2$

$$R_1 = 2R$$

When concentration is doubled, the rate of reaction is also doubled.

ii. Effect on the rate when Concentration of B is doubled

$$R_1 = k[A][2B]^2$$

$$R_1 = 4k[A][B]^2$$

But $R = k[A][B]^2$

$$R_1 = 4R$$

When concentration of B is doubled, the rate of reaction is quadrupled or increases by 4 folds.

iii. how would the concentration of B be changed in order to halve the reaction rate ?

$$R_1 = \frac{1}{2} R$$

$$R_1 = \frac{1}{2} k[A][B]^2$$

$$[xB^*]^2 = \frac{1}{2}[B]^2$$

$$x^2[B^*]^2 = \frac{1}{2}[B]^2$$

$$x^2 = \frac{1}{2}$$

$$x = \sqrt{\frac{1}{2}}$$

$$x = \frac{\sqrt{2}}{2} = 0.707$$

B should be reduced by 0.707 or 70.7% is halve the rate equation.

QUESTION 3

A reaction is found to have a rate law

$$\text{Rate} = k\,[A]^2[B]$$

Define each of the terms in this rate law and show how they are affected by each of the following changes.

ii. The concentration of A is doubled

iii. A catalyst is added

iv. The concentration of A is increased by a factor of 2 where as that of B is decreased by a factor of 4

v. Temperature is increased

Solution

$$R = k\,[A]^2[B]$$

R = Rate

The rate of the chemical equation is the amount of reactant consumed per unit time.

k - k is the rate constant and is defined as the proportionality constant in any rate law at a given temperature.

[A] - The concentration of A in mol/dm^3

[B] - The concentration of B in mol/dm^3

ii. The concentration of A is doubled

$$R_1 = k\,[2A]^2[B]$$

$$R_1 = 4k[A]^2[B]$$

But R = k$[A]^2[B]$

$$R_1 = 4R$$

When concentration of A is doubled then the rate also Quadruples or is increased by 4 folds.

Catalyst

The addition of a catalyst increases the rate of reaction by creating a new pathway and lowering the activation energy of the reaction.

The concentration of A is increased by a factor of 2 where as that of B is decreased by a factor of 4

$$R_1 = k\ [2A]^2[\tfrac{1}{4}B]$$

$$R_1 = 4\times\tfrac{1}{4}\,k\ [A]^2[B]$$

$$R_1 = k\ [A]^2[B]$$

But R = k $[A]^2[B]$

$$R_1 = R$$

So when the concentration of A is increased by a factor of 2 where as that of B is decreased by a factor of 4 then the rate of reaction remains the same.

v. Temperature is increased

When temperature is increased it increases the kinetic energy of the molecules, hence increasing collision. And when collision is increased it increases effective collision therefore more molecules attain energy that is equal or greater than the activating energy and hence increasing the rate of reaction.

DETERMINATION OF REACTION ORDER AND RATE CONSTANT FROM EXPERIMENTAL DATA

Question 1

The results of the rate measurement for the reaction between P$Cl_{3(g)}$ and $Cl_{2(g)}$ to form P$Cl_{5(g)}$are summarized in the following tables:

Experiment	Conc of P$Cl_{3(g)}$ in mol/dm^3	Conc of $Cl_{2(g)}$ in mol/dm^3	Rate mol/dm/s
1	1.0	1.0	0.004
2	1.0	0.5	0.002
3	1.0	2.0	0.008
4	0.5	1.0	0.001
5	2.0	1.0	0.016

From the results of the experiment

i. Determine the overall order of the reaction and hence write the rate equation of the reaction

ii. Calculate the specific rate constant.

Solution

From the experiment 1 and 2 when P$Cl_{3(g)}$ is kept constant and the concentration of $Cl_{2(g)}$ is halved the rate of reaction is also halved.

From experiment 4 and 5 when the concentration of $Cl_{2(g)}$ is kept constant and concentration of P$Cl_{3(g)}$ is increased four folds the rate increases by a factor of 4

The rate is directly proportional to the square of the concentration of $PCl_{3(g)}$

$$\text{Rate} = R = k[PCl_{3(g)}]^2[Cl_{2(g)}]$$

Or

$$R = k[PCl_{3(g)}]^x[Cl_{2(g)}]^y$$

Where x and y represent the order of the rate of each reactants

Considering the 1 and 2 experiment,

FINDING FOR THE ORDER OF $Cl_{2(g)}$ WHICH IS Y

$$\frac{Exp\ 1}{Exp\ 2} = \frac{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}$$

$$\frac{0.004}{0.002} = \frac{k[1.0]^x[1.0]^y}{k[1.0]^x[0.5]^y}$$

$$2 = \frac{[1.0]^y}{[0.5]^y}$$

$$2 = 2^y$$

$$y = 1$$

$$\frac{Exp\ 2}{Exp\ 1} = \frac{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}$$

$$\frac{1}{2} == \frac{k[1.0]^x[0.5]^y}{k[1.0]^x[1.0]^y}$$

$$\frac{1}{2} = \frac{[0.5]^y}{[1.0]^y}$$

$$\frac{1}{2} = \frac{1}{2}^y$$

-1 = -y

y = 1

FINDING FOR THE ORDER OF $PCl_{3(g)}$ WHICH IS X

$$\frac{Exp\ 4}{Exp\ 5} = \frac{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}{k[PCl_{3(g)}]^x[Cl_{2(g)}]^y}$$

$$\frac{0.001}{0.016} = \frac{k[0.5]^x[1.0]^y}{k[2.0]^x[1.0]^y}$$

$$0.0625 = \frac{[0.5]^x}{[2.0]^x}$$

$$0.0625 = 0.25^x$$

$$5^4 = 5^{2x}$$

$$4 = 2x$$

$$x = 2$$

The rate of the reaction is therefore
R = $k[PCl_{3(g)}]^2][Cl_{2(g)}]$

The order of the reaction is = x + y

1+ 2 = 3

This is therefore a third order reaction

Specific rate constant

$$R = k[PCl_{3(g)}]^2][Cl_{2(g)}]$$

$$k = \frac{R}{[PCl_{3(g)}]^2][Cl_{2(g)}]}$$

Fix data for any of the experiment to obtain the value for k

$$k = \frac{0.016}{[2.0]^2][1.0]}$$

k = 0.004 $mol^{-2}dm^{-6}s^{-1}$

Question 2

For the reaction A + 2B → C the experimental rate was formed to vary with the initial concentrations of the reactants as follows :

Conc of A [mol/dm^3]	Conc of B [mol/dm^3]	Initial rate of formation of C {mol dm^{-1} s^{-1}]
0.1	0.1	3.0×10^{-2}
0.2	0.2	24.0×10^{-2}
0.2	0.1	12.0×10^{-2}

i. Deduce the rate law of the reaction

ii. Calculate the specific rate constant and give units

iii. What would be the initial rate of formation of c starting with 0.12 moldm^{-3} of A and 0.12 moldm^{-3} of B

Solution

Rate = R = k$[A]^x[B]^y$

By comparing equation 2 and 3

$$\frac{Exp\ 2}{Exp\ 3} = \frac{\text{k}\ [A]^x[B]^y}{\text{k}\ [A]^x[B]^y}$$

$$\frac{24.0 \text{ x } 10^{-2}}{12.0 \text{ x } 10^{-2}} = \frac{\text{k}\ [0.2]^x[0.2]^y}{\text{k}\ [0.2]^x[0.1]^y}$$

$$\frac{24.0 \text{ x } 10^{-2}}{12.0 \text{ x } 10^{-2}} = \frac{[0.2]^y}{[0.1]^y}$$

$2 = 2^y$

y = 1

Rate = R = k$[A]^x[B]^y$

By comparing equation 1 and 3

$$\frac{Exp\ 3}{Exp\ 1} = \frac{\text{k}\ [A]^x[B]^y}{\text{k}\ [A]^x[B]^y}$$

$$\frac{12.0 \text{ x } 10^{-2}}{3.0 \text{ x } 10^{-2}} = \frac{\text{k}\ [0.1]^x[0.1]^y}{\text{k}\ [0.2]^x[0.1]^y}$$

$$\frac{12.0 \text{ x } 10^{-2}}{3.0 \text{ x } 10^{-2}} = \frac{[0.1]^x}{[0.2]^x}$$

$4 = 2^x$

$$2^2 = 2^x$$

$$2 = x$$

Specific rate constant

$$R = k[A]^2[B]$$

$$k = \frac{R}{[A]^2[B]}$$

Fix data for any of the experiment to obtain the value for k

$$k = \frac{3.0 \text{ x } 10^{-2}}{[0.1]^2][0.1]}$$

$$k = 30 \text{ mol}^2/\text{dm}^6/\text{s}$$

What would be the initial rate of formation of c starting with 0.12 mol/dm-3 of A and 0.12mol/dm-3 of B

$$R = k[A]^2[B]$$

K = 30
A = 0.12
B = 0.12

$$R = 30[0.12]^2[0.12]$$