3. Chemical Kinetics

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🎯 Learning outcomes

Candidates should be able to:

3.1 Define chemical kinetics.
3.2 Explain and use the terms: rate of reaction, rate equation, and rate constant.
3.3 Qualitatively explain the factors affecting the rate of reaction.
3.4 Use collision theory to explain the influence of temperature, concentration, and particle size on chemical reaction rate
3.5 Define activation energy and activated complex.
3.6 Derive and use the integrated rate equations and half-life for zero- and first-order reactions.
3.7 Construct and use rate equations, calculating an initial rate using concentration data.
3.8 Explain the significance of the Arrhenius equation and solve related problems.
3.9 Explain and use the terms catalyst and catalysis (homogeneous, heterogeneous).
3.10 Describe enzymes as biological catalysts.
3.11 Explain the role of a catalyst in the reaction mechanism.
3.12 Solve numerical problems based on rate, rate constant, and order of zero and first-order reactions.

Introduction

Have you ever wondered why food spoils faster in summer than in winter? Why does a car engine rust over the years, but a sparkler burns in seconds? Or why do doctors use particular drug dosages? All these phenomena involve the speed of chemical reactions — and that is exactly what Chemical Kinetics is about.

Chemical kinetics: The branch of Chemistry that deals with the **rates of chemical reactions, the factors that affect them, and the mechanisms (step-by-step pathways) of those** reactions.

Some reactions occur very rapidly, such as the explosion of fireworks, while others occur very slowly, such as the rusting of iron. Chemical kinetics helps us understand why these rates differ.

Kinetics tells us *how fast* a reaction happens — not just whether it can happen (thermodynamics).

Why Does Kinetics Matter?

- Industrial chemistry: Controlling reaction speed maximises product yield and reduces cost.
- Medicine: Drug metabolism rates determine safe dosages and dosing schedules.
- Environment: Understanding rates of atmospheric reactions helps control pollution.
- Food science: Preservative chemistry relies on slowing down decomposition reactions.
- Biology: Every enzyme in your body is a natural kinetics optimiser.

Concept of reaction rate

The rate of a chemical reaction tells us how quickly reactants are consumed or products are formed over time.

Rate of Reaction: The decrease in concentration of a reactant (or increase in concentration of a product) per unit time.

Average rate of reaction:

It is the rate of reaction measured over a long interval of time. Let us consider a general reaction:

$$A \rightarrow C$$

If the initial concentration of reactant at time t_1 is $[A_1]$ and the final concentration at time t_2 is $[A_2]$. Then, the average rate of reaction in terms of the reactants is:

$$\text{Average rate} = \frac{\text{change in concentration}}{\text{change in time}}$$

$$= -\frac{[A2]-[A1]}{t2-t1} \qquad = -\frac{\Delta A}{\Delta t}$$

Why the Negative Sign?

Reactant concentration decreases over time, so Δ[A] is negative. The negative sign is included to ensure the rate is always a positive quantity. When expressing rate in terms of products, no negative sign is needed because product concentration increases.

Hence, the rate of the reaction in terms of product can be expressed as follows.

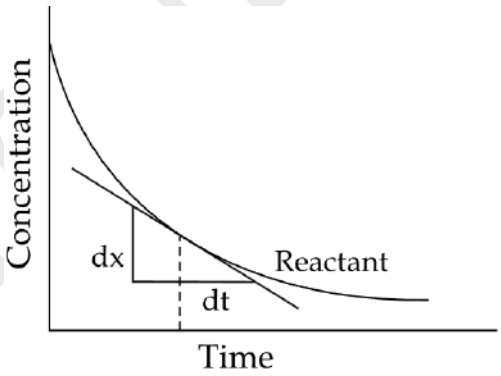
Rate (R) = $\frac{\Delta C}{\Delta t}$ where, $\Delta C = C_2 - C_1$(increase in concentration) and $\Delta t = t_2 - t_1$ (time taken)

Instantaneous rate of reaction: the rate of reaction measured at any instant of time.

It is expressed as follows:

$\frac{dx}{dt} = \lim(\Delta t \rightarrow 0) \frac{\Delta C}{\Delta t}$ where dx = very small change in concentration and dt = very small change in time

On a concentration vs. time graph, this is the slope of the tangent line at that point.



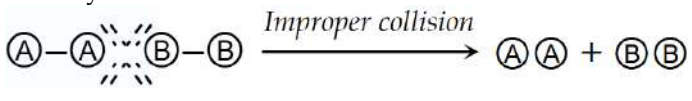
Collision theory of reaction rate

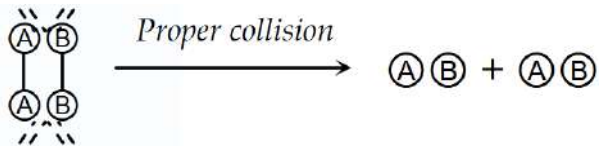
Ineffective collision: particles just bounce apart unchanged.

Effective collision: leads to product formation.

Proper orientation

The molecules must align correctly for bonds to break and form.





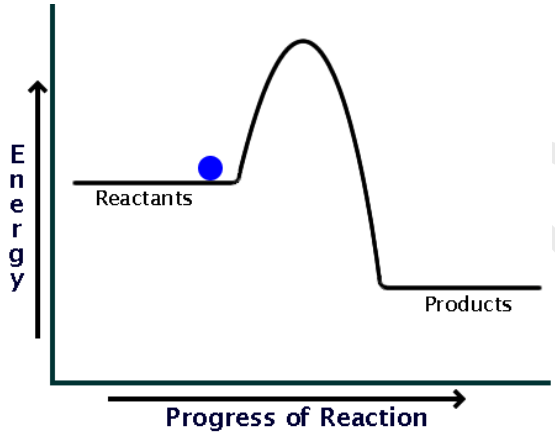
Sufficient energy, enough kinetic energy to overcome the activation energy barrier.

Ineffective Collision	Effective Collision
Particles bounce apart unchanged	Bonds break, and new bonds form → Products are formed
Energy < activation energy, OR wrong orientation	Energy ≥ activation energy AND correct orientation
No reaction occurs	Reaction occurs successfully

[How to speed up chemical reactions (and get a date) - Aaron Sams]

Concept of activation energy

The **Activation Energy (E_a)** of a reaction is the minimum energy required by colliding particles for a collision to be effective. Reaction pathway diagrams show how the activation energy provides a barrier to reaction.



Activated Complex (Transition State) is the high-energy, unstable arrangement of atoms at the peak of the energy barrier. It exists for an extremely short time before breaking apart into products.

Factors affecting reaction rate

Nature of Reactants

- Ionic reactions are usually fast.
- Covalent reactions are generally slow because bonds must be broken first.

Surface Area of the reactants

Greater surface area increases the rate of reaction because more particles are exposed for collision.
For example, powdered calcium carbonate reacts faster than marble chips.

Concentration

As concentration increases, the **frequency of collisions** increases, resulting in an increased reaction rate.

Pressure: Similarly, when **pressure is increased** in a gaseous reaction, the **frequency of collisions increases**, **increasing the reaction rate**.

Temperature

At higher temperatures, molecules have more kinetic energy, so a higher percentage of successful collisions occurs between reactant molecules.

Temperature increases the rate of reaction because of two factors:
1. The frequency of effective collisions is greater because of the greater kinetic energy of the molecules
2. A greater proportion of the molecules have kinetic energy greater than the activation energy.
The second reason has a far greater effect.

Factor	Effect	Explanation (Collision Theory)
Nature of reactants	Ionic reactions are fast; covalent bonds require breaking → slower	Ionic species are already separated; covalent bonds need energy to break
Surface area (solid)	More surface area → faster	More particles are exposed to collisions
Concentration	Higher concentration → faster rate	More particles per volume → more frequent collisions
Pressure (gases)	Higher pressure → faster rate	Particles pushed closer together → collision frequency increases
Temperature	Higher temperature → much faster rate	① Molecules move faster (more collisions) ② **Far more important:** a larger fraction of molecules have energy ≥ E_a

The Arrhenius Equation

The Arrhenius equation mathematically shows how temperature affects the rate constant.

Arrhenius equation: $k = A\, e^{-Ea/RT}$

Here, k= is the rate constant,

A = pre-exponential factor or Arrhenius factor

E_a = activation energy,

R = universal gas constant,

T = absolute temperature.

Even a small increase in temperature causes a *large* increase in k because of the exponential term.

Worked Examples

1. The rate constant for a reaction is $2.0 \times 10^{-3}\ s^{-1}$ at 300 K and $8.0 \times 10^{-3}\ s^{-1}$ at 320 K. Calculate the activation energy.

Solution:

Using: $\log(k_2/k_1) = E_a / 2.303R \times (1/T_1 - 1/T_2)$

$\log(8.0\times10^{-3} / 2.0\times10^{-3}) = E_a / (2.303 \times 8.314) \times (1/300 - 1/320)$

$\log(4) = E_a / 19.14 \times (2.083 \times 10^{-4})$

$0.602 = E_a \times 1.088 \times 10^{-5}$

$E_a = 0.602 / 1.088 \times 10^{-5} \approx 55{,}330\ J\ mol^{-1} \approx 55.3\ kJ\ mol^{-1}$

2. The activation energy for a reaction is 80 kJ mol^{-1} and A = 1.0 × 10^{1} s^{-1}. Calculate the rate constant at 27°C.

Solution:

T = 27 + 273 = 300 K, E_a = 80,000 J mol^{-1}, R = 8.314 J mol^{-1} K^{-1}

Using $k = A \times e^{-(E_a/RT)}$:

E_a/RT = 80,000 / (8.314 × 300) = 80,000 / 2,494.2 = 32.08

$k = 1.0 \times 10^{9} \times e^{-32.08}$

$k = 1.0 \times 10^{9} \times 9.16 \times 10^{-14}$

$k \approx 9.16 \times 10^{-5}$ s^{-1}

Catalysis

Catalysis: the increase in the rate of a chemical reaction brought about by the addition of particular substances that are not used up in the reaction.

Catalysts provide an *alternative reaction pathway* with lower activation energy.

Catalyst

Catalysts increase the rate because they make the reaction go by a different reaction pathway (mechanism) that has a lower activation energy than the uncatalysed reaction.

A catalyst **is a substance that increases the rate of a reaction but remains chemically unchanged itself at the end of the reaction.**

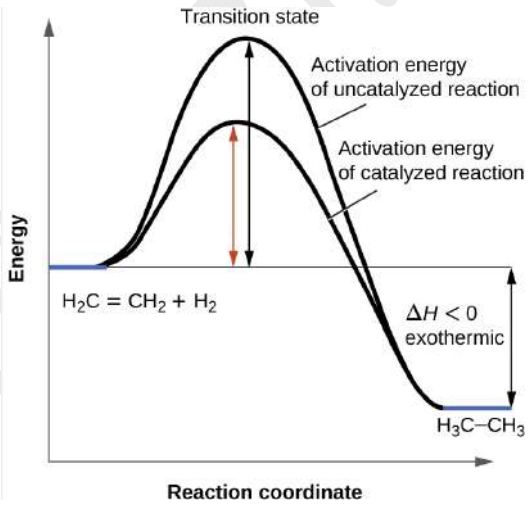


Figure: Effect of a catalyst on the activation energy of an exothermic reaction

Homogeneous Catalyst: When a catalyst and the reactants are in the same phase.

- Example 1: Concentrated H_2SO_4 (liquid) catalysing the esterification of an alcohol with a carboxylic acid (both liquids).
- Example 2: Fe^{2+} ions (aqueous) catalysing the reaction between $S_2O_8^{2-}$ and I^- in aqueous solution.
- Example 3: Acid-catalysed hydrolysis of sucrose in aqueous solution.

Heterogeneous Catalyst: A catalyst that is in a different phase from the reactants.

- Example 1: Iron (Fe) catalyst in the Haber process — $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
- Example 2: Nickel in the hydrogenation of vegetable oils (margarine production)
- Example 3: Vanadium(V) oxide (V_2O_5) in the Contact process for manufacturing H_2SO_4

Enzyme catalysis

Inside every living cell, thousands of chemical reactions are occurring every second — made possible by remarkable protein catalysts called enzymes.

Enzyme: A protein molecule that acts as a biological catalyst, increasing the rate of a specific biochemical reaction in living organisms.

1. **Key Features**
   - **Highly specific:** One enzyme catalyzes one type of reaction.
     Example: Diastase hydrolyzes starch → maltose, but does not hydrolyze cellulose.
   - **Optimum temperature:** Maximum activity at a specific temperature (~37°C for human enzymes). Too high → denatures (loses shape). Too low → too slow.
   - **Optimum pH:** Each enzyme works best at a specific pH.
     Example: Pepsin (stomach) works at pH ~2; trypsin (small intestine) works at pH ~8.
   - **Analogy:** An enzyme is like a specific key that fits only one lock. Change the lock (different substrate), and the key doesn't work.

Pre-lab: Liver and Enzyme activity - YouTube

Role of Catalyst in Reaction Mechanism:

A catalyst works by providing an alternative mechanism with a lower activation energy for the rate-determining step. This can involve:

1. The catalyst forms an intermediate with one of the reactants (homogeneous catalysis).
2. Reactants adsorbing onto the catalyst surface, where bonds are weakened (heterogeneous catalysis).

Either way, the energy barrier is reduced, and the reaction proceeds faster.

Rate law equation, rate constant, and its units

Experiments show that the reaction rate depends on the concentration of reactants in a specific mathematical way. This relationship is captured in the rate equation (also called the rate law).

The **rate equation** is an expression showing how the reaction rate depends on the concentration of the reactants. The rate equals the rate constant (k) multiplied by the concentration of each reactant raised to certain whole-number powers (called orders).

The rate equation for the general reaction P + Q → R + S is:

$$\text{rate} = k[\text{P}]^x [\text{Q}]^y$$

where k = **rate constant**

x = **order of reaction** with respect to reactant P,

y = order of reaction with respect to reactant Q, overall order = x + y

Rate equation: an equation showing the relationship between the rate constant and the concentrations of the species that affect the rate of reaction.

Rate constant: the proportionality constant, k, in a rate equation

The **overall order of the reaction** is the sum of all the orders in the rate equation. For this reaction, the overall order of reaction would be $x + y$.

Rate has units of mol L^{-1} s^{-1}. Concentration has units of mol L^{-1}.

The units of the rate constant, k, depend on the overall order of the reaction, and the units can be calculated.

Order of reaction: the power to which the concentration of the reactant is raised in the experimentally determined rate equation. If a rate is directly proportional to concentration, it is a first-order reaction; if the rate is directly proportional to the square of the concentration, it is a second-order reaction. The overall order of reaction is the sum of these powers.

⚠ Critical Point — Order is Experimental, NOT Stoichiometric

The order of a reaction CANNOT be determined simply by looking at the balanced equation. It must be determined experimentally. For example, the reaction $2NO_2 \rightarrow 2NO + O_2$ has a stoichiometric coefficient of 2 for NO_2, but the actual rate law is Rate = $k[NO_2]^2$, which happens to match — but this is coincidental, not a rule!

Zero-order reaction

A zero-order reaction proceeds at a constant rate, unaffected by the reactant's concentration.

e.g., Formation of HCl gas from H_2 and Cl_2 gas in the presence of sunlight on the surface of water.

$H_2(g) + Cl_2(g)\ h\upsilon \rightarrow 2\ HCl\ (g)$

For a zero-order reaction, $\mathbf{Rate = k\,[A]^0 = k}$

i.e., rate = k

So, units of k = mol L^{-1} s^{-1}

What Does Zero-Order Mean Physically?

Zero-order behaviour often occurs when a catalyst surface is completely saturated with reactant. Adding more reactant cannot increase the rate because there are no free active sites remaining. The catalyst (not the reactant concentration) is the limiting factor.

First-order reaction

A first-order reaction is a reaction whose rate depends on the concentration of one reactant raised to the power one, meaning the rate is directly proportional to the first power of its concentration (Rate $\propto [A]^1$).

e.g., Decomposition of hydrogen peroxide:
$2H_2O_2 \rightarrow 2H_2O + O_2$
Radioactive decay of nuclei:
(e.g., decay of carbon-14 or uranium-238)
Decomposition of nitrogen dioxide:
$2N_2O_5 \rightarrow 4\ NO_2 + O_2$

For a first-order reaction, $\mathbf{Rate = k\,[A]}$

So, units of k = s^{-1}

Pseudo-order reaction

A pseudo-order reaction is a reaction in which one reactant is present in large excess, so its concentration remains nearly constant during the reaction. As a result, the rate appears to depend only on the concentration of the other reactant.

A pseudo-first-order reaction is actually a second-order reaction that behaves like a first-order reaction.

Example: Hydrolysis of ethyl ethanoate (ethyl acetate) with excess water:

$CH_3COOC_2H_5 + H_2O \rightarrow CH_3COOH + C_2H_5OH$

Rate = $k[CH_3COOC_2H_5][H_2O]$, but since water is in vast excess, $[H_2O] \approx$ constant.

So Rate $\approx k'[CH_3COOC_2H_5]$, where $k' = k[H_2O]$ = pseudo-first-order rate constant.

Second-order reaction

Rate $\propto$ [reactant]^2 — if concentration doubles, rate quadruples

e.g., $2NO_2 \rightarrow 2NO + O_2$

For a second order reaction, rate = $k\,[A]^2$

So, units of rate constant, k = $mol^{-1}\ L\ s^{-1}$

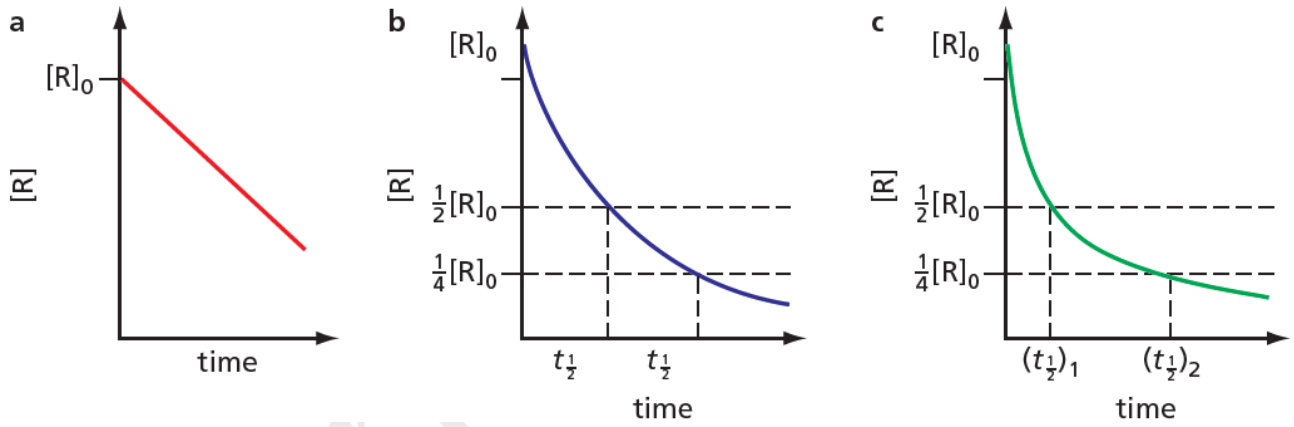


Figure: The three most common types of concentration-time graphs: **a** zero order, **b** first order, **c** second order. $[R]_0$ is the concentration at time t = 0.

Overall Order	Units of k
Zero order	$mol\ L^{-1}\ s^{-1}$
First order	s^{-1}
Second order	$mol^{-1}\ L\ s^{-1}$

Order and molecularity of reaction

Feature	Molecularity	Order
Definition	No. of particles colliding simultaneously in the rate-determining step	Sum of powers of concentrations in the experimental rate equation
Values	Whole numbers only: 1, 2, 3	Can be 0, 1, 2, or even fractional
How determined	From the reaction mechanism (theoretical)	From experimental data (empirical)
For complex rx^{n}s	Refers to the slowest (rate-determining) step	Overall observed value from the experiment

Integrated rate law for a zero-order reaction

Consider a general zero-order reaction in which reactant A is converted to a product. Let:

- a = initial concentration of reactant (mol L^{-1})
- x = amount of reactant converted to product after time t (mol L^{-1})
- $(a - x)$ = concentration of reactant at time t (mol L^{-1})

$$\mathbf{A} \longrightarrow \text{Product}$$

At $t = 0$: $[A] = a$ | At $t = t$: $[A] = (a - x)$

For a zero-order reaction, the rate is **independent** of the concentration of the reactant. The rate depends on the zeroth power of concentration:

$$\text{Rate} \propto [A]^0 = \text{constant}$$

Expressing the rate in terms of the change in x with respect to time:

$$dx / dt = k_0(a - x)^0$$

$$dx / dt = k_0 \quad \ldots (i)$$

where k_0 is the **zero-order rate constant**. Since $[A]^0 = 1$, the rate equals the rate constant at all times. Equation (i) is the **differential rate equation** for the zero-order reaction.

Rearranging equation (i) to separate variables:

$$dx = k_0\, dt$$

Integrating both sides:

$$\int dx = k_0 \int dt$$

$$x = k_0 t + C \quad \ldots (ii)$$

where C is the constant of integration.

At $t = 0$, $x = 0$. Substituting into equation (ii):

$$0 = k_0 \times 0 + C$$

$$C = 0$$

Substituting $C = 0$ back into equation (ii):

$$x = k_0 t + 0$$

$$x = k_0 t$$

Since $x = a - (a - x)$, we can also write this in terms of the remaining concentration $[A] = (a - x)$:

$$(a - x) = a - k_0 t$$

$$x = k_0 t \quad \ldots (iii)$$

Rearranging for the rate constant:

$$k_0 = x / t \quad \ldots (iv)$$

Equations (iii) and (iv) are the **integrated rate equations for a zero-order reaction**. Equation (iii) shows that x varies linearly with time, and equation (iv) gives the rate constant directly from the amount of reactant consumed per unit time.

- The rate is **constant** and independent of the concentration of the reactant.
- A plot of concentration [A] vs. time t gives a **straight line** with slope $-k_0$.
- The rate constant k_0 has units of mol L^{-1} s^{-1} (concentration per unit time).

- Common examples include enzyme-catalysed reactions at saturating substrate concentrations, and reactions on solid catalytic surfaces.

Integrated rate law for a first-order reaction

A first-order reaction is a chemical reaction in which the rate depends on the concentration of only one reactant raised to the first power.

Some examples of first-order reactions are given below.

Decomposition of H_2O_2: $2\,H_2O_2 \longrightarrow H_2O + O_2$, $\frac{dx}{dt} = k_1\,[H_2O_2]^1$

Decomposition of N_2O_5: $2\,N_2O_5 \longrightarrow 4\,NO_2 + O_2$, $\frac{dx}{dt} = k_1\,[N_2O_5]^1$

Let us consider a reaction in which reactant A changes into a product. Let a be the initial concentration of reactant A.

After a certain time t, x mol of reactant changes into product.

$$k = \frac{2.303}{t} \log \frac{a}{a-x}$$

Consider a general first-order reaction in which reactant *A* is converted to a product. Let:

- a = initial concentration of reactant (mol L^{-1})
- x = amount of reactant converted to product after time t (mol L^{-1})
- $(a - x)$ = concentration of reactant at time t (mol L^{-1})

A $\longrightarrow$ Product

At $t = 0$: [A] = a | At $t = t$: [A] = $(a - x)$

For a first order reaction, the rate depends on only **one** concentration term:

Rate $\propto$ $[A]^1$

Expressing the rate in terms of the change in x with respect to time:

$dx / dt \propto (a - x)$

$dx / dt = k_1(a - x)$ … (i)

where k_1 is the **rate constant** (also called the velocity constant or specific rate of reaction). Equation (i) is the **differential rate equation** for the first order reaction.

Rearranging equation (i) to separate variables:

$dx = k_1(a - x)\,dt$

$dx / (a - x) = k_1\,dt$

Integrating both sides, and using the standard result $\int dx/(ax + b) = \ln(ax + b)/a + C$:

$\int dx / (a - x) = k_1 \int dt$

$-\ln(a - x) = k_1 t + C$ … (ii)

where C is the constant of integration.

At $t = 0$, $x = 0$. Substituting into equation (ii):

$-\ln(a - 0) = k_1 \times 0 + C$

$C = -\ln a$

Substituting the value of C back into equation (ii):

$-\ln(a - x) = k_1 t - \ln a$

$\ln a - \ln(a - x) = k_1 t$

Applying the logarithm rule $\ln a - \ln b = \ln(a / b)$:

$\ln \times [a / (a - x)] = k_1 t$

$k_1 = (1 / t) \times \ln [a / (a - x)]$ … (iii)

Converting to base-10 logarithm using $\ln a = 2.303 \log a$:

$k_1 = (2.303 / t) \times \log [a / (a - x)]$ … (iv)

- The rate depends linearly on the concentration of a single reactant.
- The integrated rate equation is logarithmic, giving a straight line when ln[A] is plotted against *t*.
- The rate constant k_1 has units of s^{-1} (or min^{-1}, h^{-1}, etc.).

Half-life reaction

The half-life period is the time for the concentration of a reactant to decrease to half its initial value. It is denoted by $t_{1/2}$.

Half-life t½: time taken for the amount (or concentration) of the limiting reactant in a reaction to decrease to half its initial value.

By applying the strict definition of half-life, we evaluate the system at the specific moment when the elapsed time *t* is equal to the half-life $t_{½}$. At this exact threshold, the remaining reactant concentration reduces to half of its initial amount:

$$At\ t = t_{½},\ [A]_t = [A]0_2$$

Substituting this boundary condition directly into the integrated rate law formula:

$$[A]0_2 = [A]_0 - k \cdot t_{½}$$

Isolating the term containing the half-life by rearranging the terms:

$$k \cdot t_{½} = [A]_0 - [A]0_2$$

$$k \cdot t_{½} = [A]0_2$$

Finally, dividing by the rate constant yields the definitive expression for the zero-order half-life:

$$t_{½} = [A]0_{2k}$$

First-order reactions have a fixed half-life time, which is given by the following relation:

$$\boldsymbol{k = 0.693/t_{1/2}}$$

To find the half-life expression, we insert the foundational half-life boundary conditions:

$$At\ t = t_{½},\ [A]_t = [A]0_2$$

Substituting this into the quotient natural logarithmic integrated rate law:

$$ln([A]0 / 2_{[A]0}) = -k \cdot t_{½}$$

The initial concentration terms completely cancel out within the fraction, revealing a critical property:

$$ln(1_2) = -k \cdot t_{½}$$

Using log properties, $ln(1/2) = -ln(2)$. Substituting this removes the negative signs from both sides of the kinetic equality:

$$-\ln(2) = -k \cdot t_{½}$$

$$\ln(2) = k \cdot t_{½}$$

Isolating $t_{½}$ gives the exact value expression:

$$t_{½} = \ln(2)_k$$

Since the transcendental constant $\ln(2)$ evaluates approximately to **0.693**, we arrive at the standard practical engineering format:

$$t_{½} = 0.693_k$$

Kinetic Parameter	Zero-Order Reaction	First-Order Reaction
Differential Rate Law	Rate = k	Rate = k[A]
Integrated Rate Law	[A]t = [A]0 − kt	[A]t = [A]0 · e^−kt
Half-Life Expression	t½ = [A]0 / (2k)	t½ = 0.693 / k
Rate Constant (k) Units	M · s^−1	s^−1
Concentration Dependence	Directly proportional to [A]0	Completely independent of [A]0

Chapter Summary

- Chemical kinetics is the study of reaction rates, the factors that affect them, and reaction mechanisms.
- Rate of reaction = Δ[concentration] / Δtime. Units: mol L^{-1} s^{-1}.
- Collision theory: Effective collisions require (i) energy ≥ E_a and (ii) correct orientation.
- Activation energy (E_a): minimum energy needed for a reaction; the activated complex forms at the energy peak.
- Factors increasing rate: increased surface area, concentration, pressure (gases), temperature; adding a catalyst.
- Arrhenius equation: $k = Ae^{-}(E_a/RT)$. Temperature has an exponential effect on k.
- Rate equation: Rate = $k[A]^x[B]^y$. Order is determined experimentally, NOT from stoichiometric coefficients.
- Zero order: Rate = k; $t½ = [A]_o/2k$ (half-life decreases over time).
- First order: Rate = k[A]; t½ = 0.693/k (constant half-life — key feature).
- Second order: Rate = $k[A]^2$; k units: mol^{-1} L s^{-1}.
- A catalyst lowers E_a by providing an alternative reaction pathway; it is regenerated.
- Homogeneous catalysis: catalyst in the same phase as reactants. Heterogeneous: different phases.
- Enzymes: biological protein catalysts; highly specific; have optimum temperature and pH; follow the lock-and-key model.

Key Terms Glossary

Term	Definition
Activation Energy (E_a)	Minimum energy needed for colliding particles to react.
Activated Complex	High-energy, unstable arrangement of atoms at the energy peak.

Arrhenius Equation	$k = Ae^{-}(E_a/RT)$; mathematically links rate constant to temperature.
Catalyst	Substance increasing reaction rate without being consumed; lowers E_a.
Collision Theory	A theory that explains reactions occur through effective collisions.
Enzyme	Biological protein catalyst; highly specific; has optimum T and pH.
Half-Life (t½)	Time for the reactant concentration to fall to half its initial value.
Molecularity	Number of particles colliding in the rate-determining elementary step.
Order of Reaction	Power to which the concentration is raised in the experimental rate equation.
Rate Constant (k)	Proportionality constant in the rate equation; depends on T and E_a.
Rate Equation	Rate = $k[A]^x[B]^y$; experimental relationship between rate and concentrations.
Rate-Determining Step	Slowest step in a mechanism; controls the overall reaction rate.
Reaction Mechanism	A sequence of elementary steps by which reactants form products.

*****This is not a complete note. It is to guide you. It is recommended to study the prescribed textbooks along with this material. *****