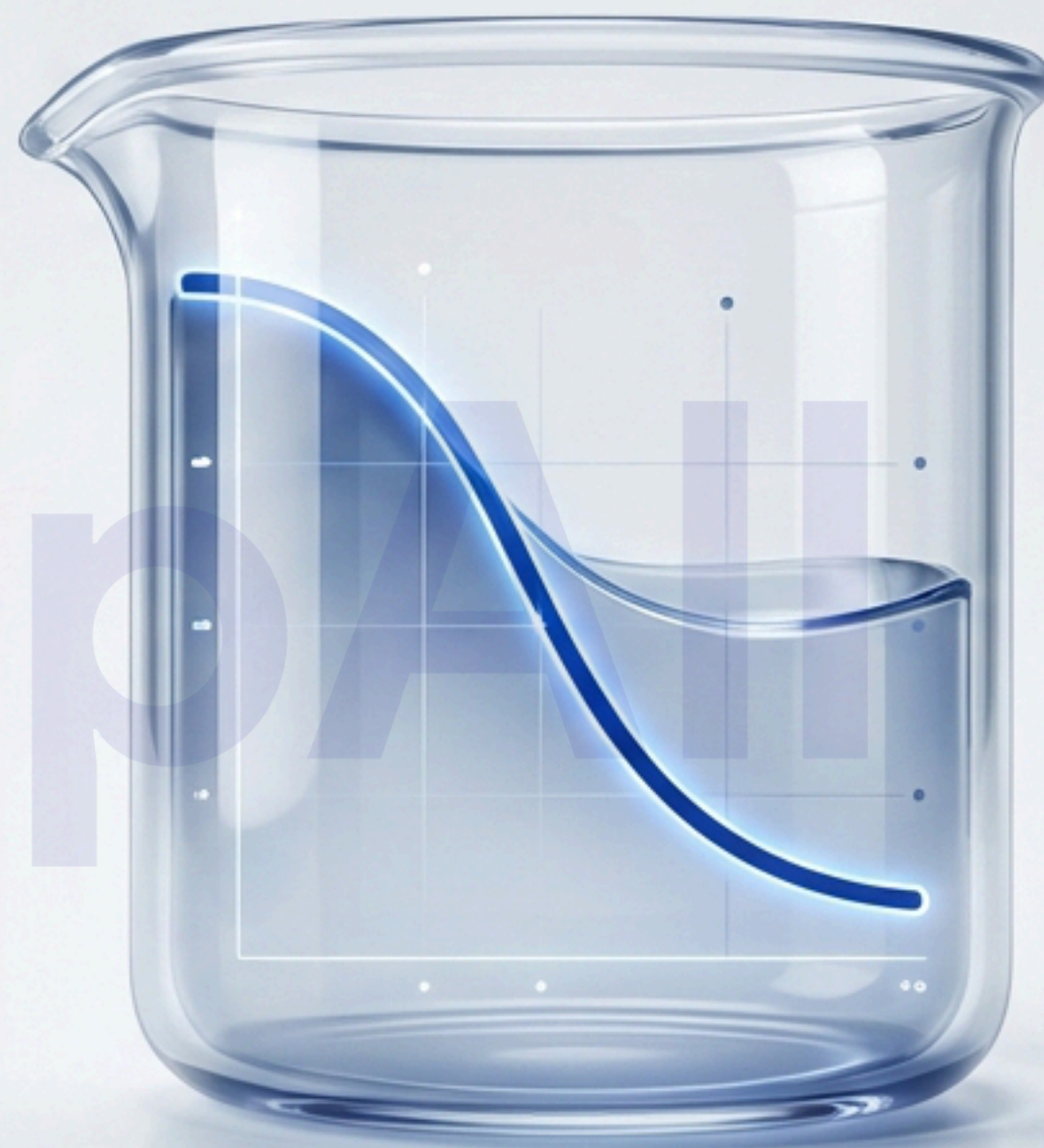


UNIT 3: CHEMICAL KINETICS

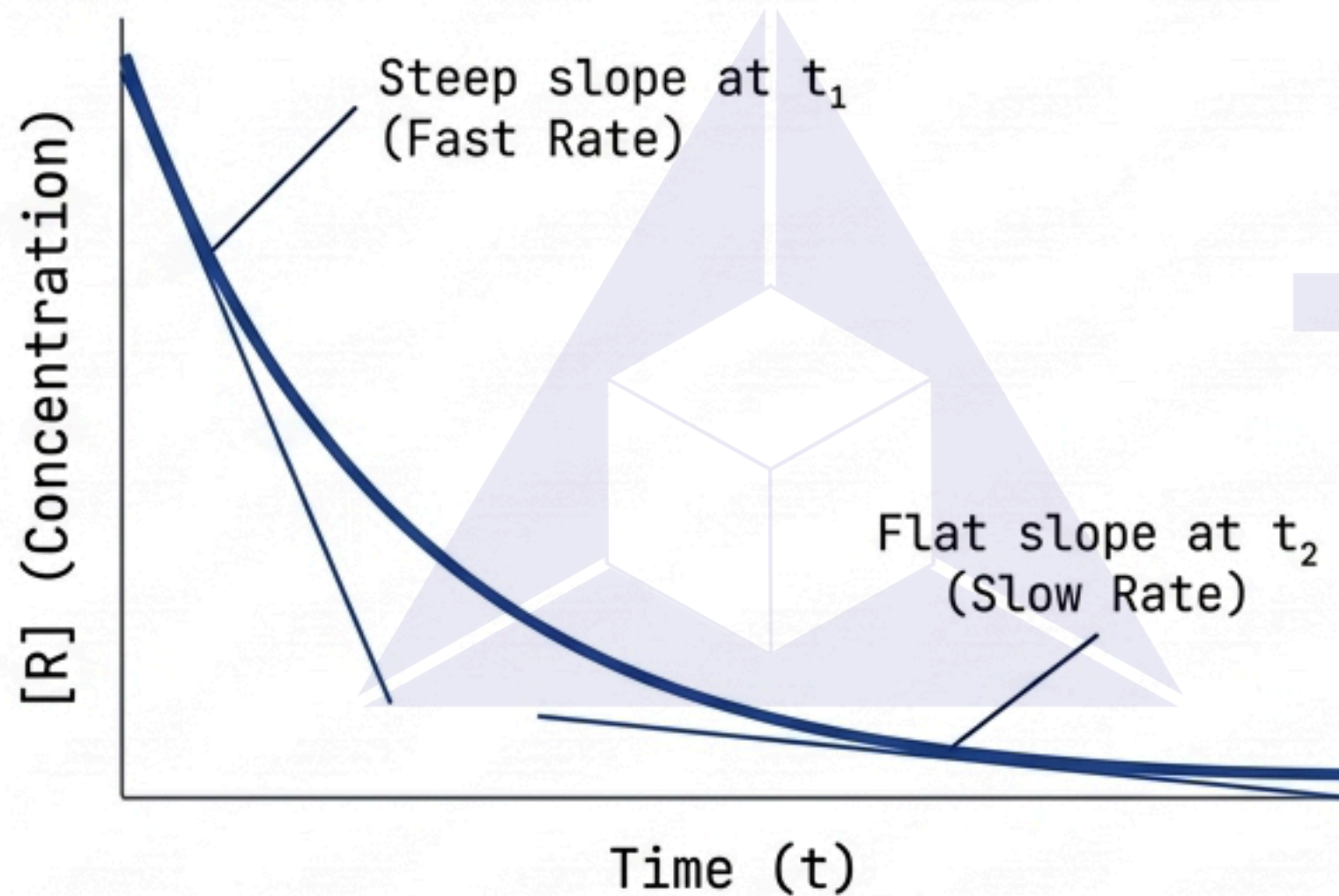
3.2.1 Dependence of Rate on Concentration

Bridging Macroscopic Observations with Microscopic Mechanisms and Mathematical Rate Laws.



Reactions Naturally Decelerate Over Time

Visual Evidence



The Macroscopic Reality

As a reaction proceeds, reactants are consumed. This decreasing availability of reactants causes a continuous, natural drop in the reaction rate.

Observation: Rate is not constant; it falls as reactant concentration falls.



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$$E=mc^2$$



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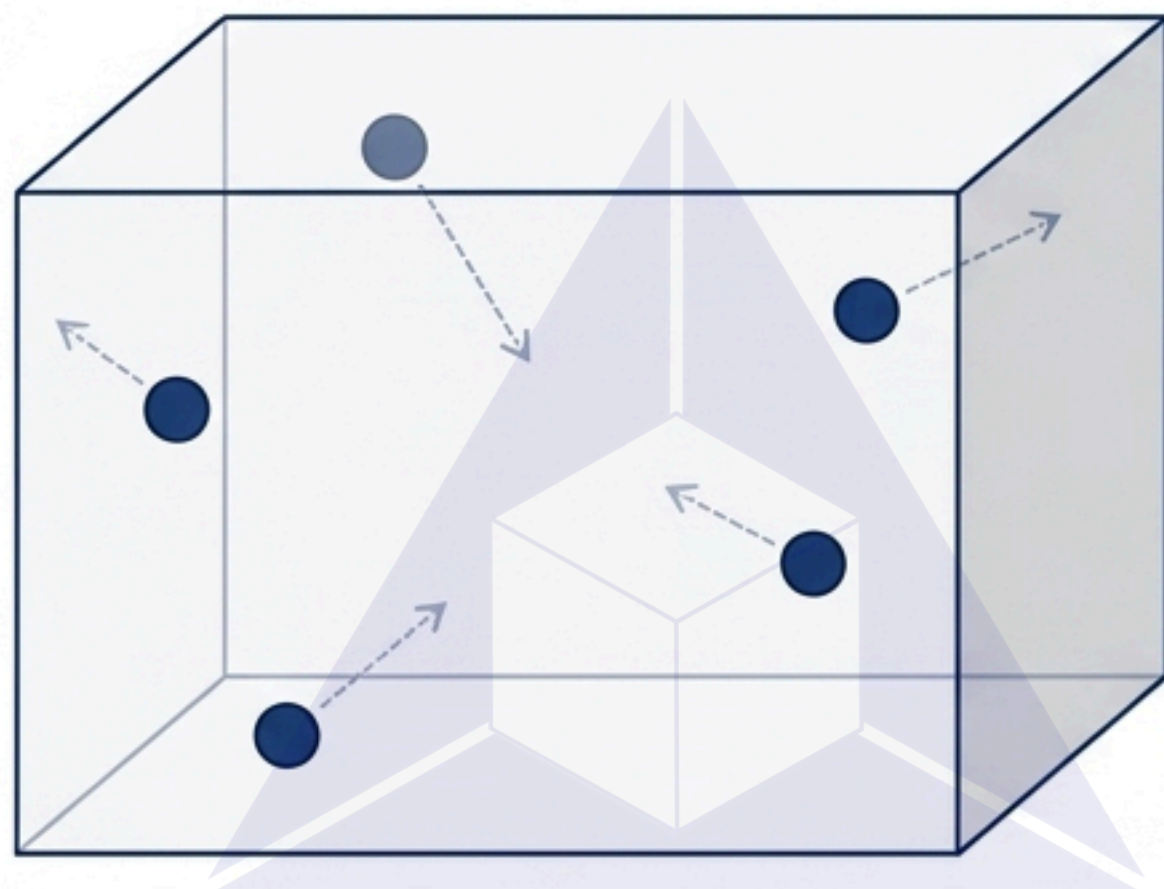
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Key Highlights

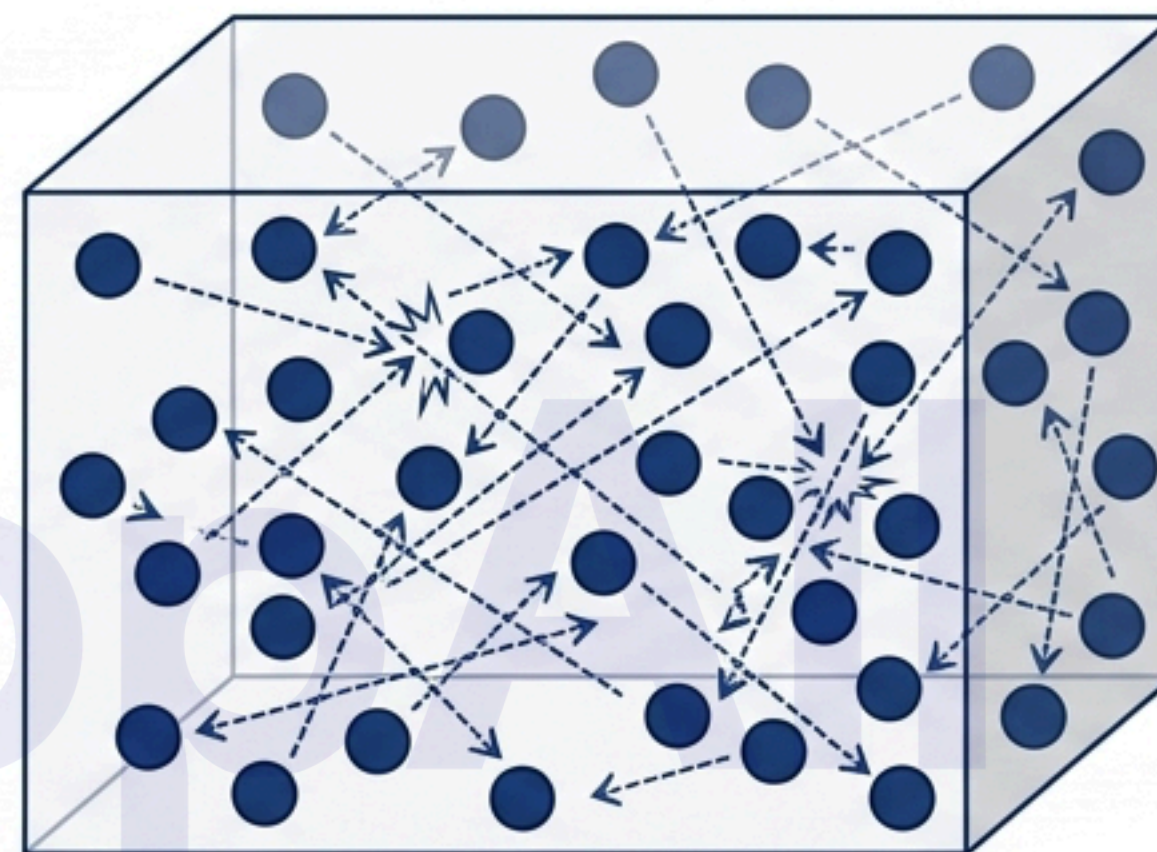
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The Microscopic Mechanism is a Numbers Game



Low Concentration = Fewer Encounters



High Concentration = High Collision Frequency (Z)

Collision Theory dictates that chemical reactions rely on the frequency of effective collisions. More reactant molecules in a given volume inherently increases the statistical probability of these encounters.

Translating Observations into the Rate Law

The Diagnostic Matrix



1	Double $[\text{O}_2]$ (keep NO constant)	→	Rate $\times 2$ (1st Order)
2	Double $[\text{NO}]$ (keep O_2 constant)	→	Rate $\times 4$ (2nd Order)

The Rate Law Formula

$$\text{Rate} = k[\text{NO}]^2[\text{O}_2]^1$$

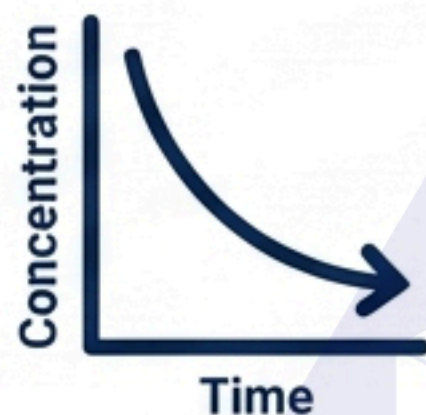
k is the Rate Constant

Exponents represent the Order of the reaction

❗ The Rate Law cannot be deduced from a balanced equation—it must be derived from experimental data.

Synthesis: The Concentration-Rate Connection

Macroscopic (The Lab)



Reaction rates universally **decelerate** as reactant **concentration** is depleted over time.

Microscopic (The Mechanism)



Higher concentrations pack more molecules into a volume, drastically **increasing** the **collision frequency** (Z).

Mathematical (The Model)

k

Experimental changes in concentration allow us to calculate the reaction order and write the definitive Rate Law expression ($\text{Rate} = k[A]^x[B]^y$).

**Concentration dictates particle encounters. Particle encounters dictate reaction speed.
Experimental data gives us the math.**



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