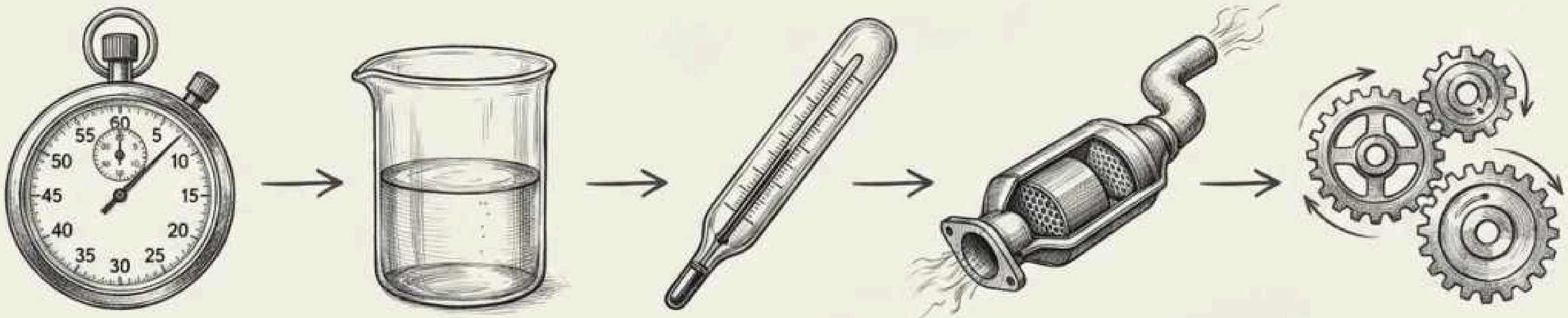
¶ 1.1

Reaction Rates and Kinetics

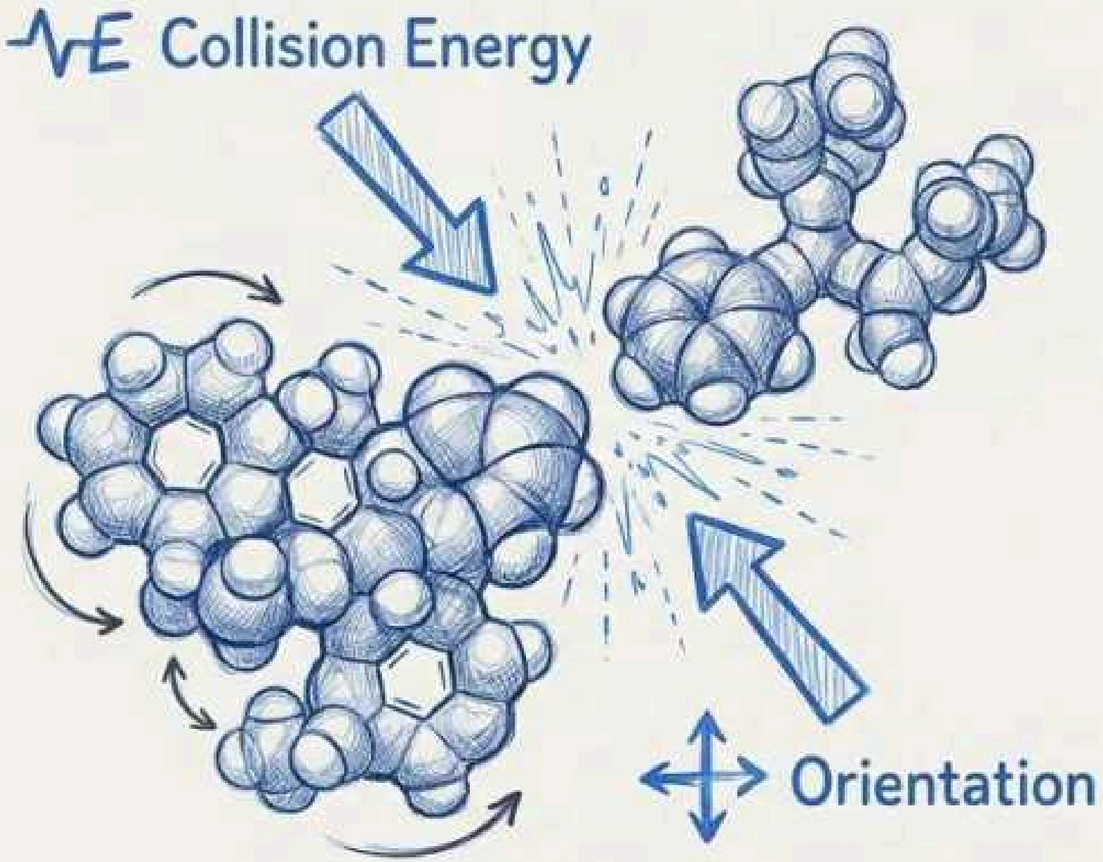


Initial Rates Half-Lives Arrhenius Catalysts Mechanisms

¶ 1.2

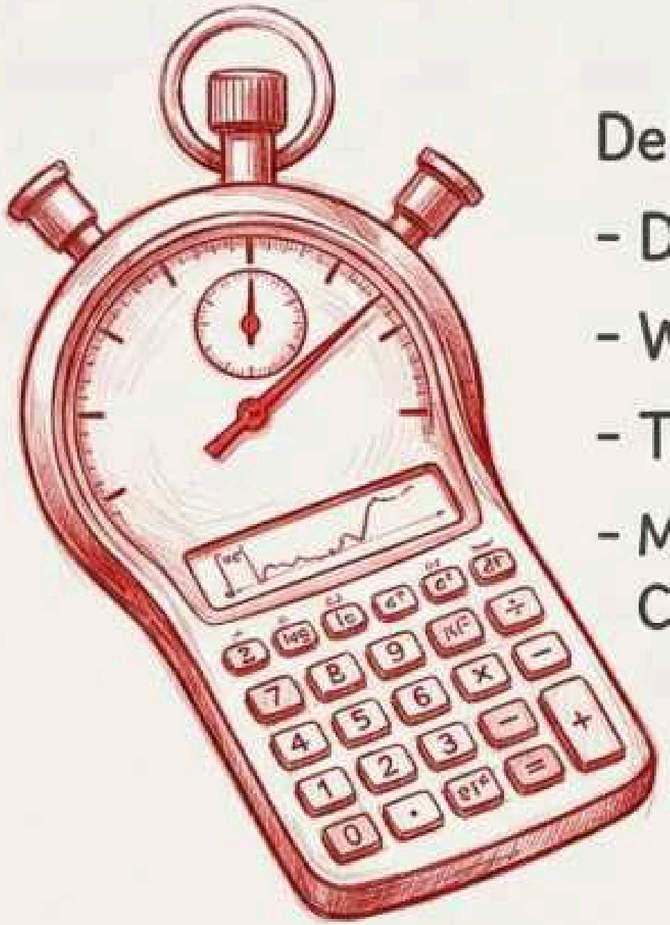
Chemistry 1

HOW reactions happen



Kinetics

HOW FAST they happen



Deliverables:

- Define a Rate
- Write a Rate Law
- Track Half-Lives
- Measure Temperature / Catalyst Effects

¶ 1.3

$$\text{Rate} = \frac{-\Delta[A]}{\Delta t}$$

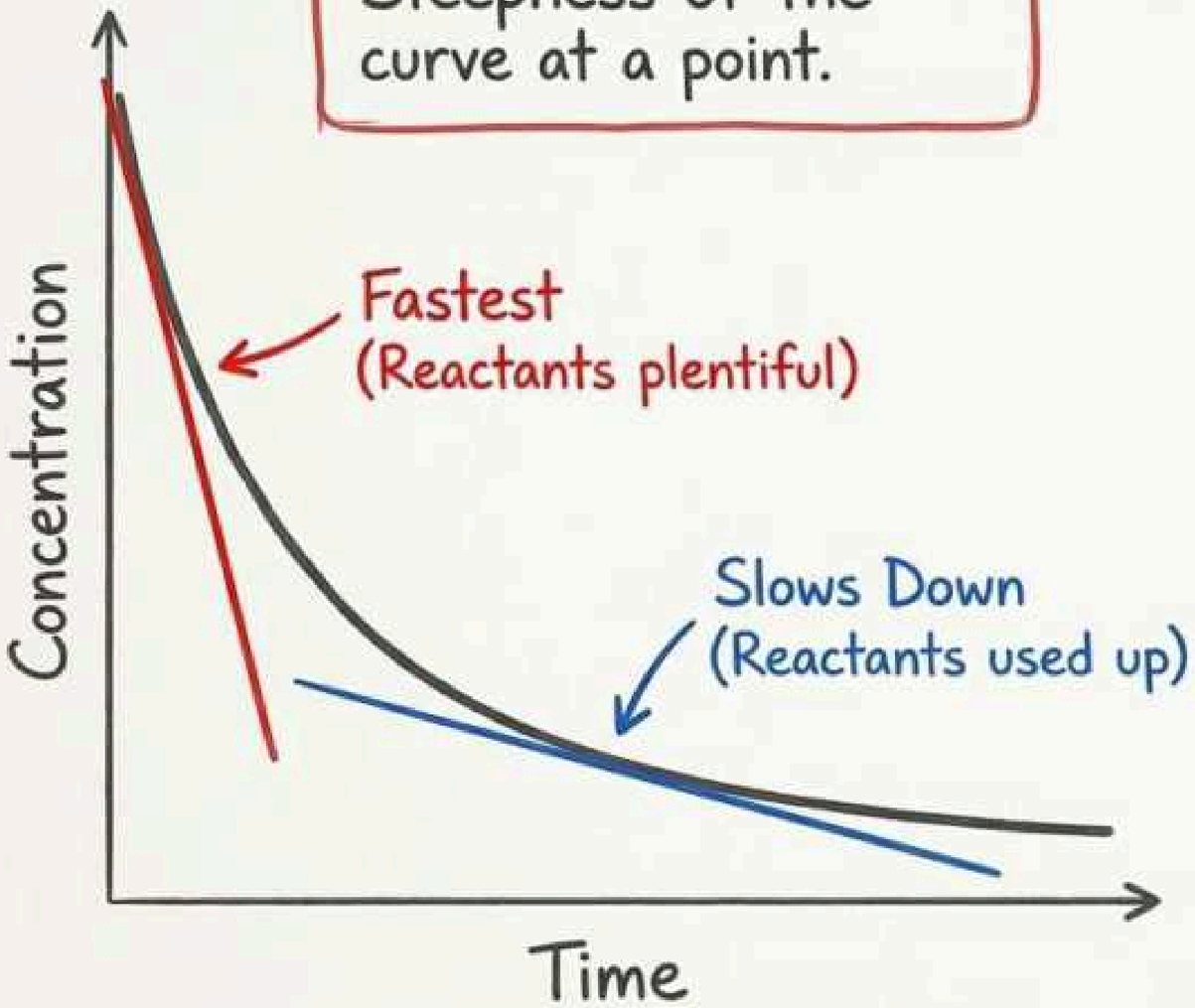
Concentration in moles per litre (Molar)

Time taken

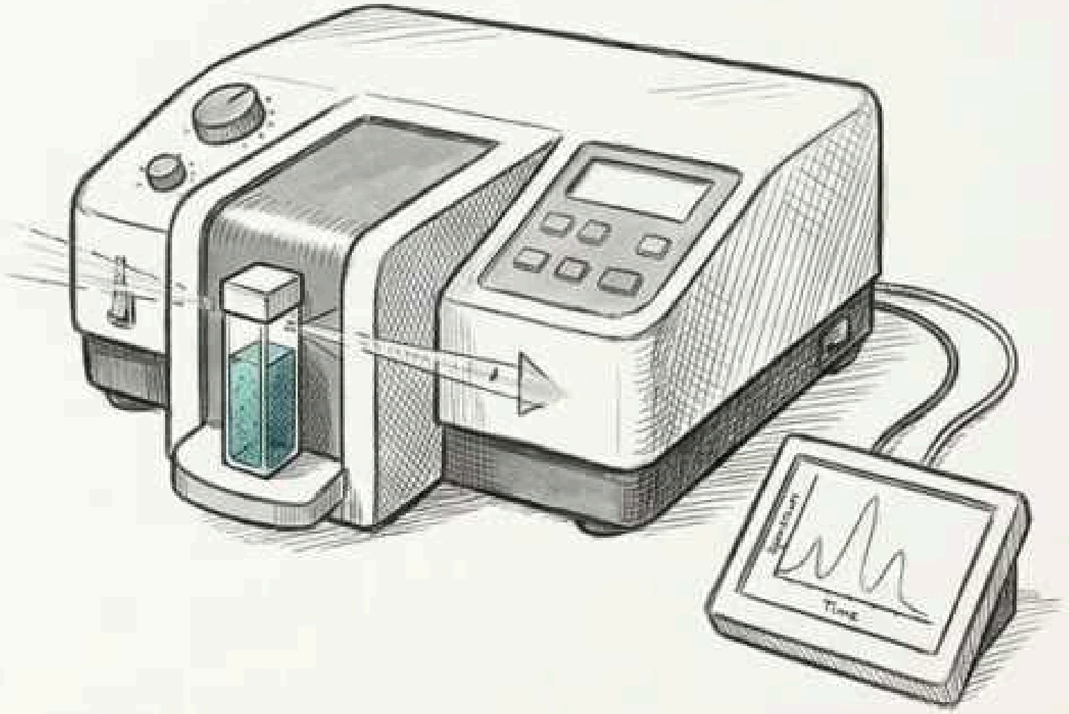
Initial: 0.80 M
Final: 0.60 M
Drop: 0.20 M
Time: 50 seconds

$$\frac{0.20}{50} = 0.004 \text{ M/s}$$

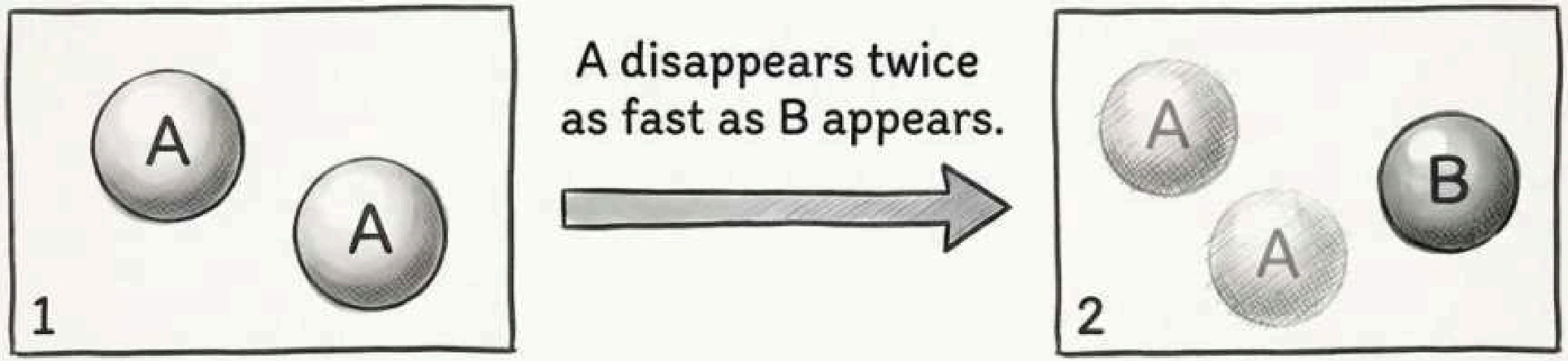
Instantaneous Rate = Steepness of the curve at a point.



In practice, chemists track observable changes: color, pressure, or gas mass.

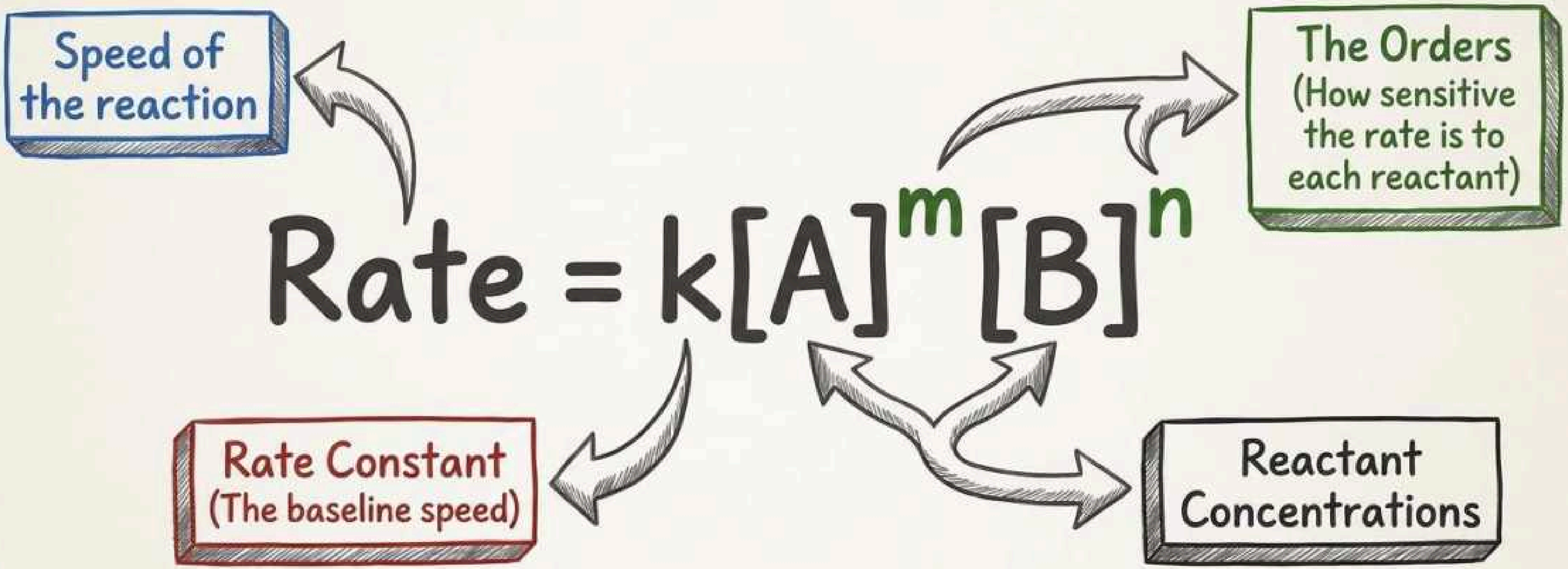


$2A \rightarrow B$



$$\text{Rate} = -\frac{1}{2}\left(\frac{\Delta[A]}{\Delta t}\right) = +\frac{1}{1}\left(\frac{\Delta[B]}{\Delta t}\right)$$

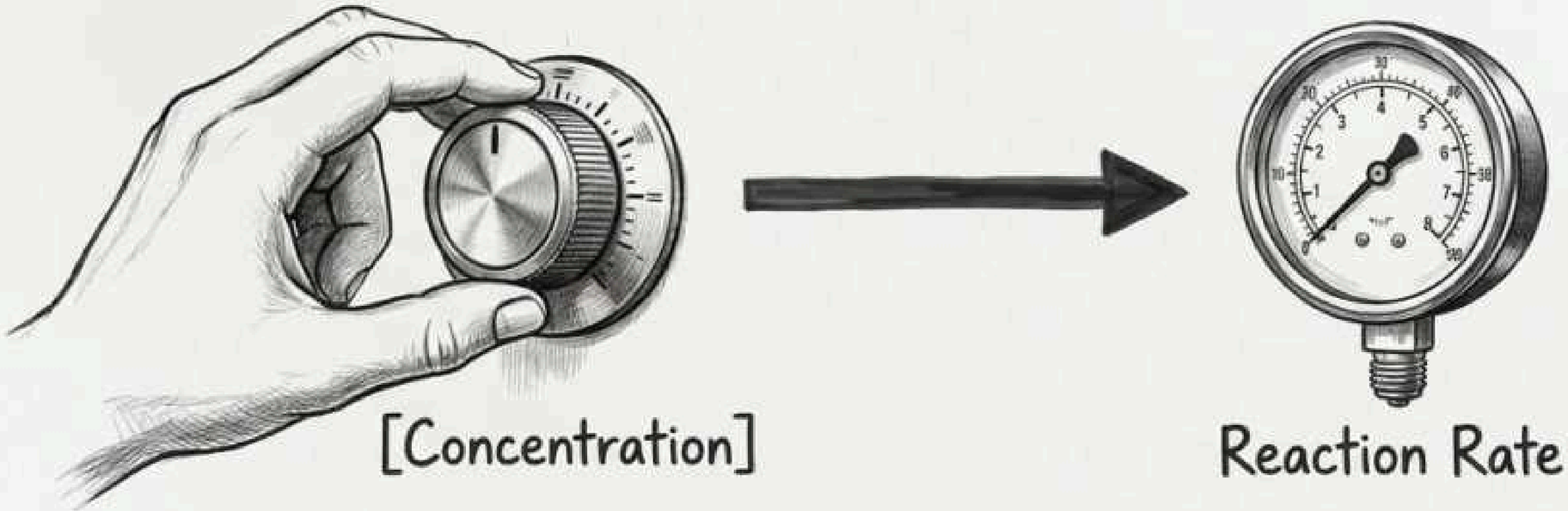
Divide by the coefficients to achieve a single, unified reaction rate.


Speed of the reaction
The Orders (How sensitive the rate is to each reactant)
Rate = k[A]^m[B]^n
Rate Constant (The baseline speed)
Reactant Concentrations

~~Balanced Equation ≠ Rate Law~~

✓ Must be found by experiment.

The Method of Initial Rates



[Concentration] → Reaction Rate

Change ONE concentration at a time ➜ Watch the rate.

¶ 1.8

	[A]	[B]	Rate
Exp 1	0.10 M	0.10 M	2×10^{-3} M/s
Exp 2	0.20 M	0.10 M	8×10^{-3} M/s

×2

×4

Doubling A quadrupled the rate.

Rate $\propto [A]^2$ → Second Order in A.

¶ 1.9

	[A]	[B]	Rate
Exp 1	0.10 M	0.10 M	2×10^{-3} M/s
Exp 3	0.10 M	0.20 M	2×10^{-3} M/s

×2

No Change

Doubling B changed nothing. B is Zero Order.

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

1. Start with the law: $\text{Rate} = k[A]^2$
2. Plug in Exp 1: $2 \times 10^{-3} = k(0.10)^2$
3. Simplify: $2 \times 10^{-3} = k(0.01)$
4. Solve: $k = 0.20$

UNITS MATTER!

$k = 0.20\ M^{-1}s^{-1}$

Units change with the overall order, always carry them.

¶ 1.11

First-Order Reactions

Concentration falls by the same fraction
in every equal interval of time.



Radioactive Decay



Drug Metabolism

The time it takes to fall by exactly half is the Half-Life.

Natural logarithm of 2 (ln 2)

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

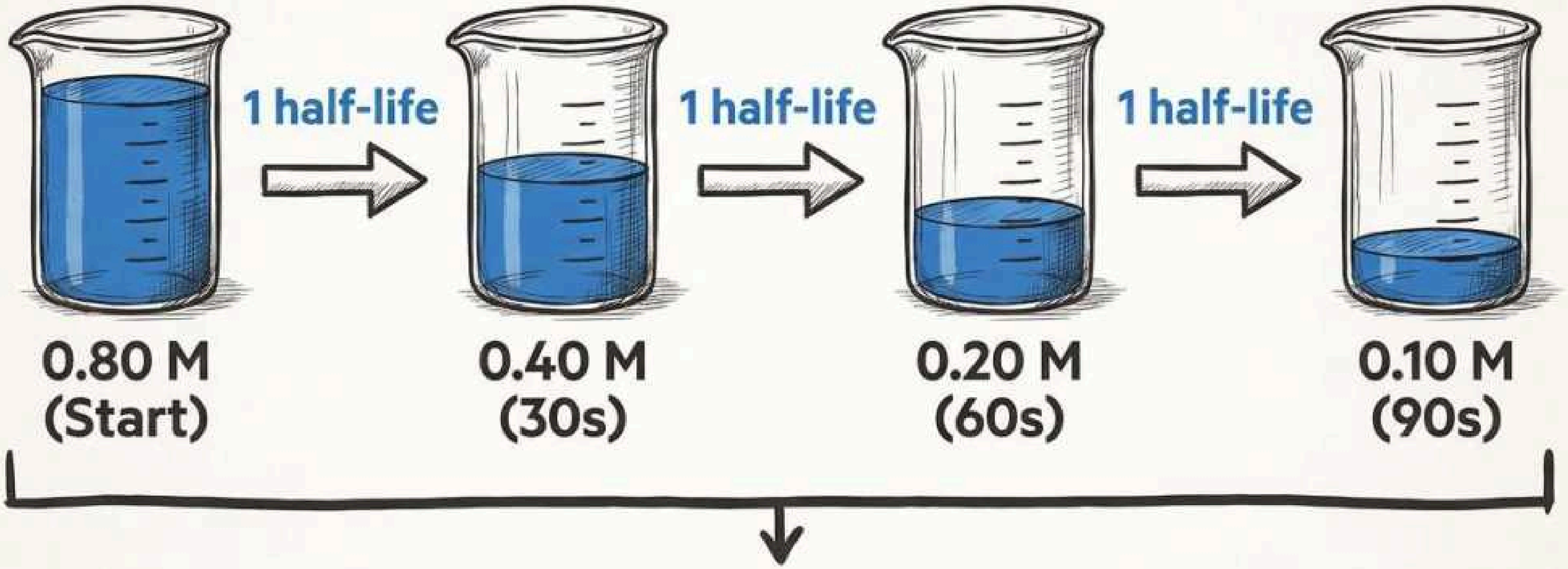
Notice what is MISSING:
The starting concentration $[A]_0$ has no effect on the half-life!

If $k = 0.0231\ s^{-1}$

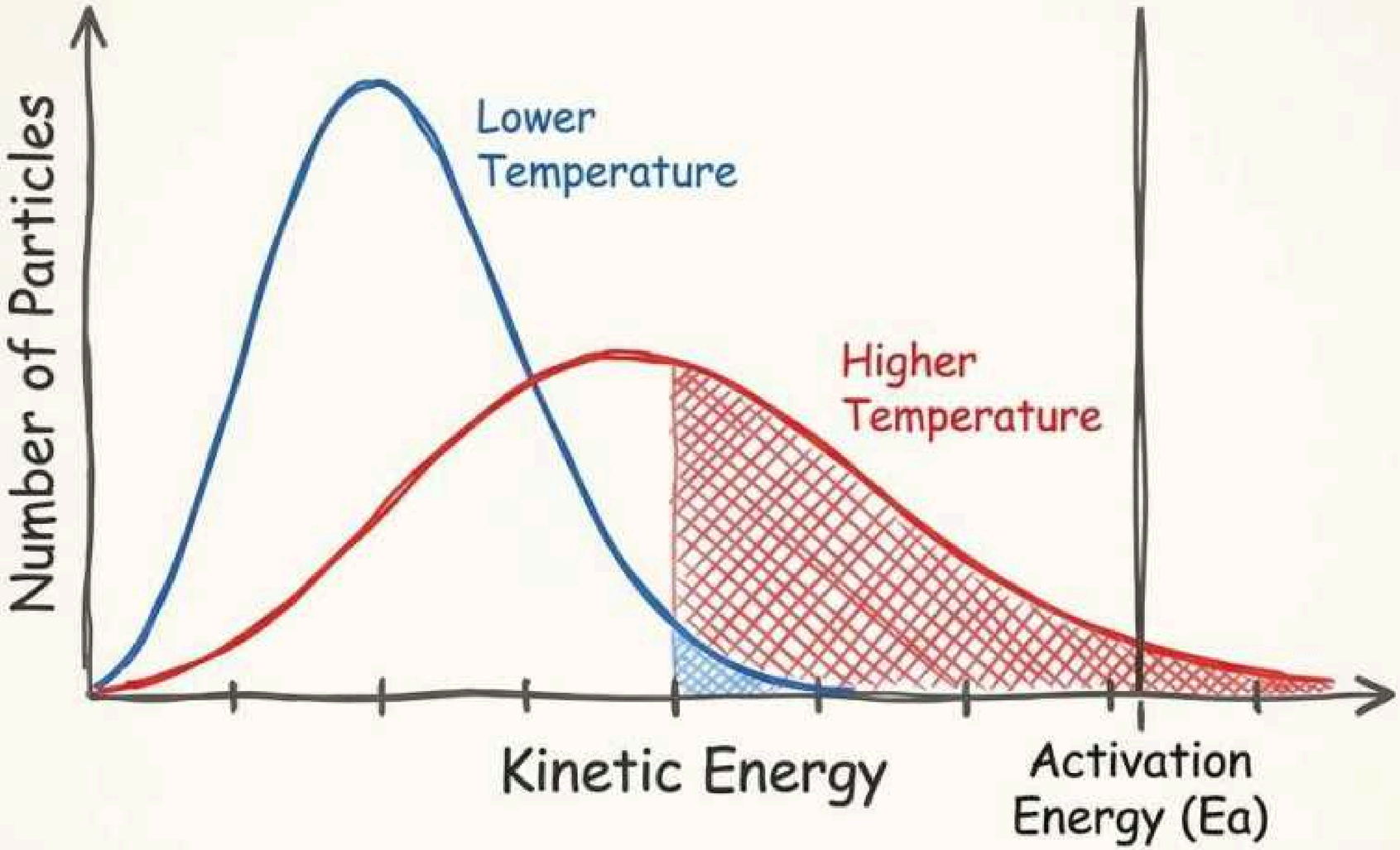
$$t_{1/2} = \frac{0.693}{0.0231}$$

30 seconds

¶ 1.13



Each half-life halves what is left.
After 10 half-lives, less than 1/1,000th remains.



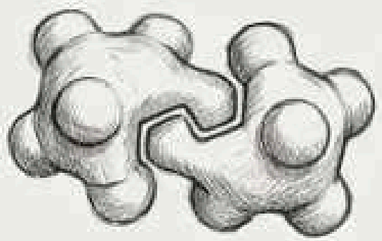
Raising temperature makes particles faster. The higher the Activation Energy (Ea), the more sensitive the rate is to temperature changes.

¶ 1.15

$$k = A e^{-Ea/RT}$$

Rate Constant

Frequency Factor



Counts total collisions & proper orientations.

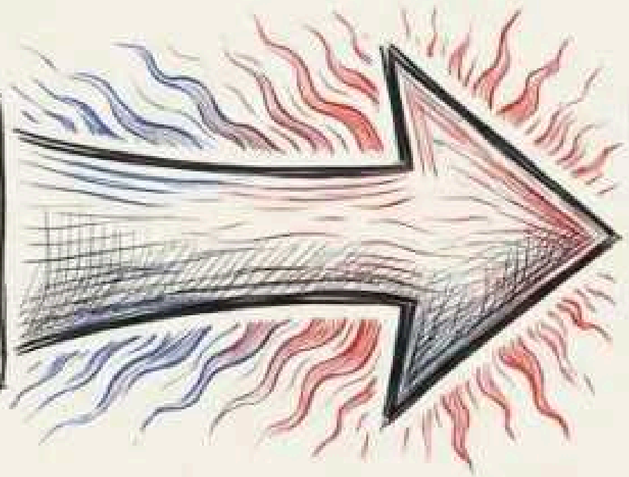
The share of collisions that actually have enough energy to react.

R = Gas constant (8.314 J/mol·K)

T = Temperature in Kelvin

Input Box

Activation Energy = 50 kJ/mol.
Warm from 298K to 308K
(+10 degrees).



Output Box

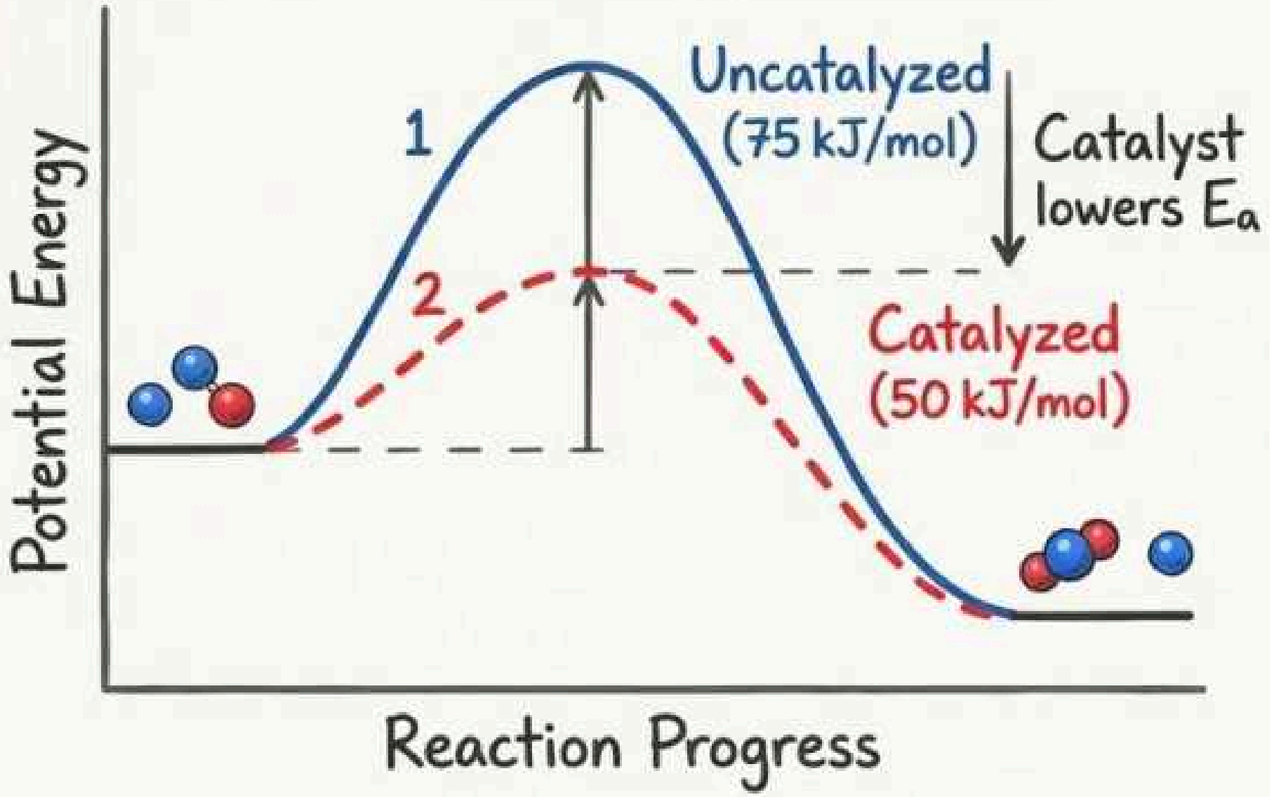
The rate constant (k)
grows by a factor of 1.9.

The 10-Degree Rule:

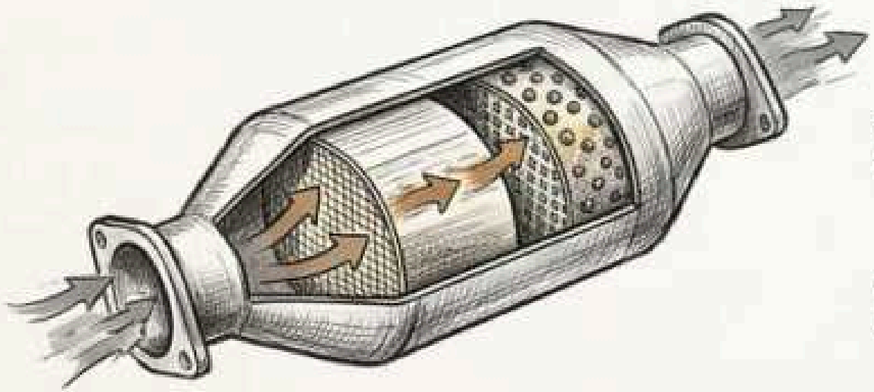
A 10K rise roughly doubles the rate.

*(With a higher Ea, the jump is even bigger!)

Energy Profile

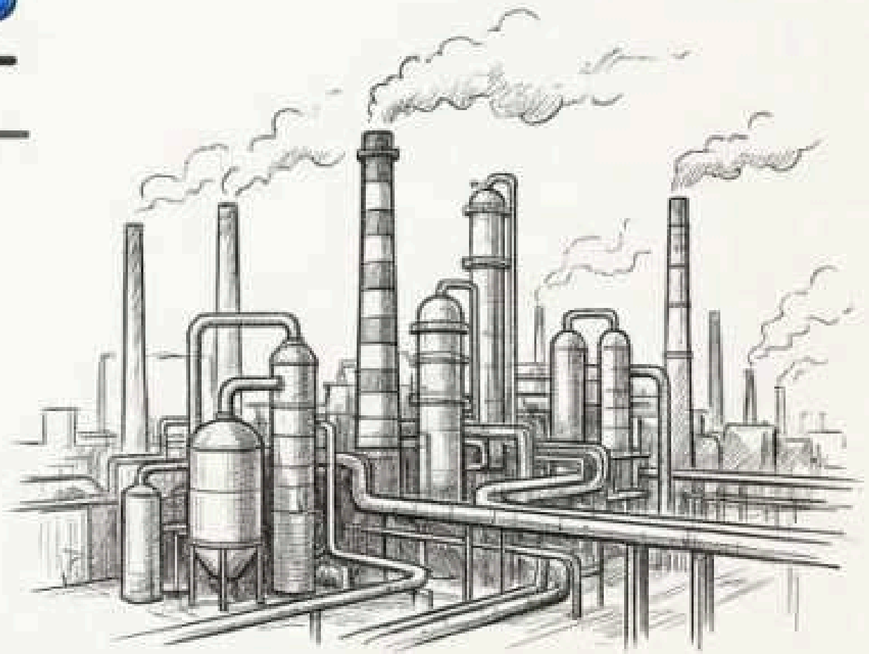


This **25 kJ** drop raises the **rate** constant by a factor of **~24,000** at room temp.



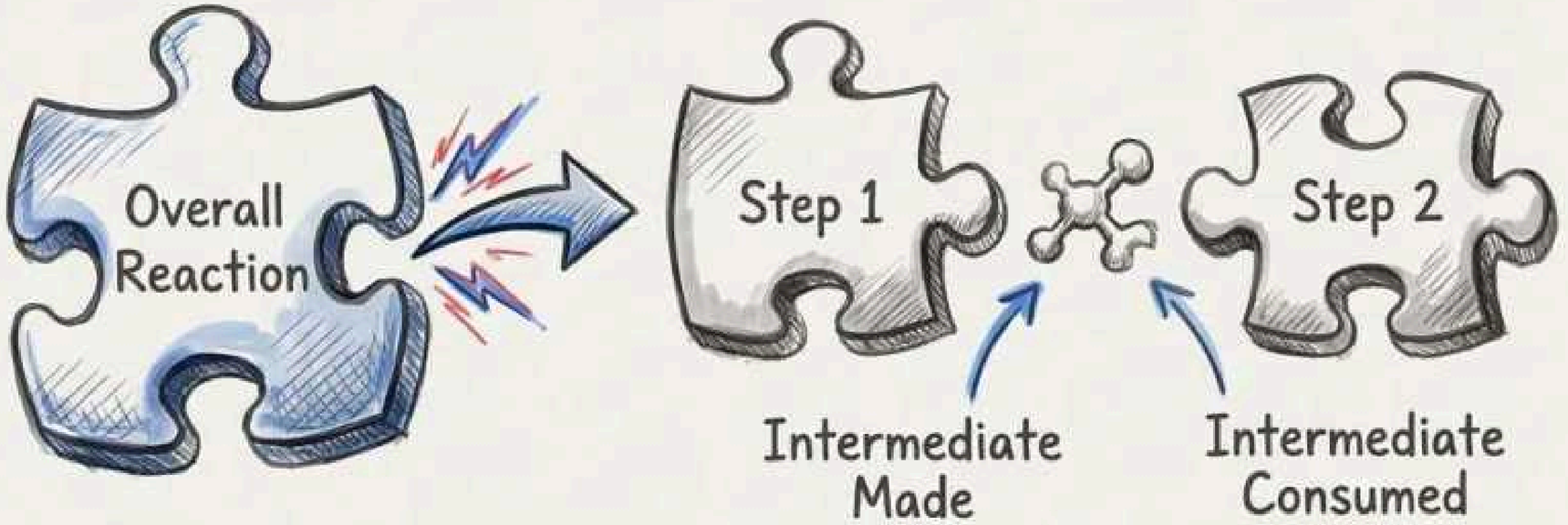
Platinum in Car Exhausts

Iron in Ammonia Production



¶ 1.18

Most reactions are not single collisions. They are sequences called mechanisms made of Elementary Steps.

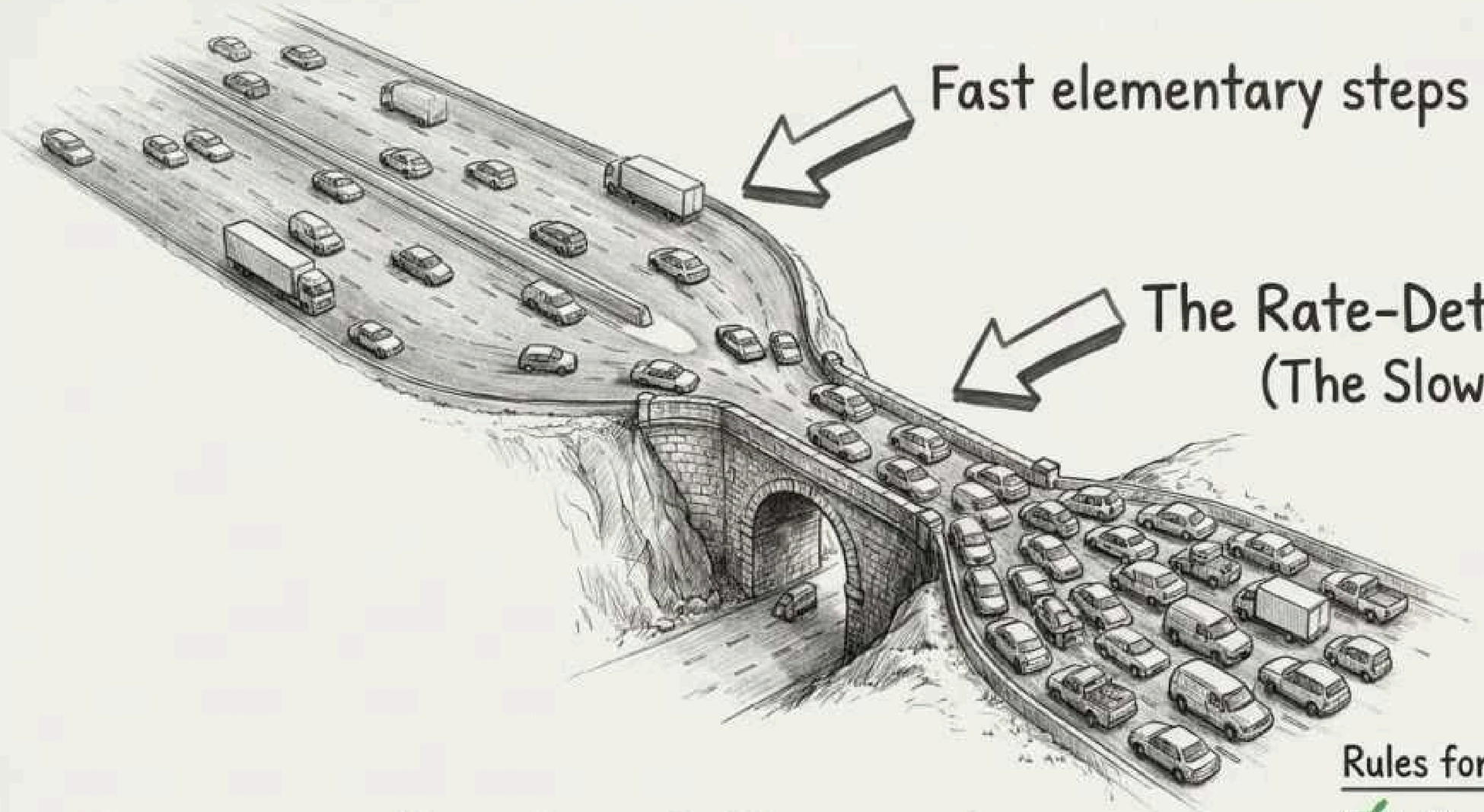


THE CRUCIAL EXCEPTION:

For elementary steps ONLY, stoichiometry dictates the rate law.

- 1 molecule = 1st order step
- 2 molecules = 2nd order step

¶ 1.19

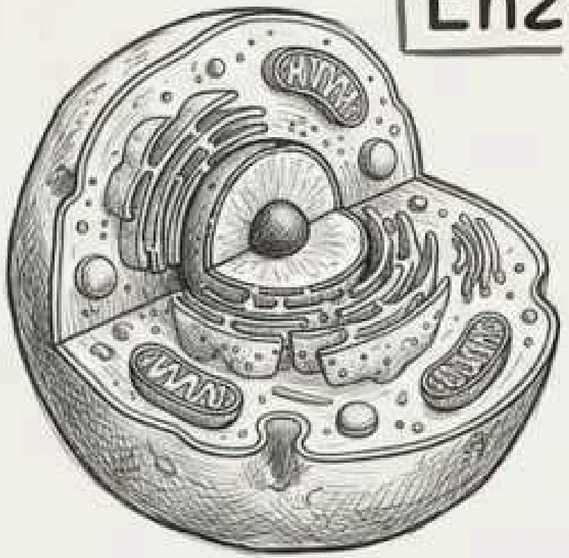


The overall rate of the system is strictly limited by the slowest step.

Rules for Proposed Mechanisms:

- ☑ 1. Must add up to the overall equation.
- ☑ 2. Must correctly predict the experimental rate law.

Enzymes



Biological catalysts running at body temp in milliseconds.

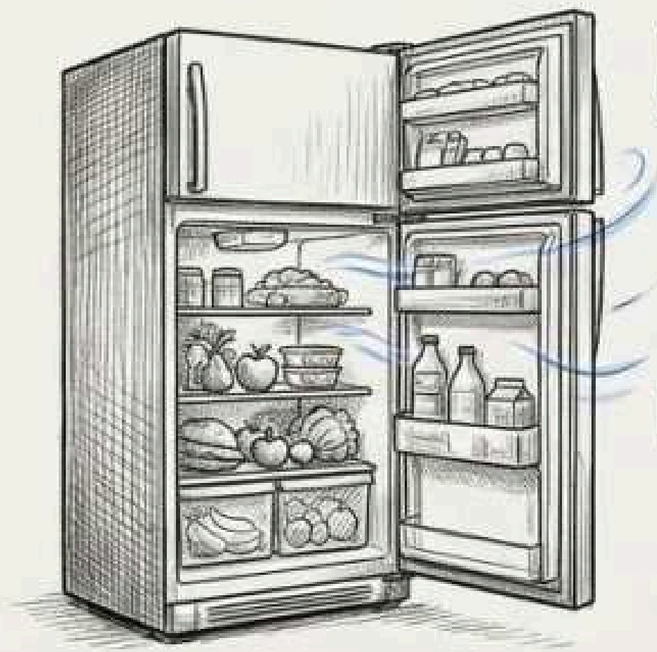
Cooking

¶ 1.20



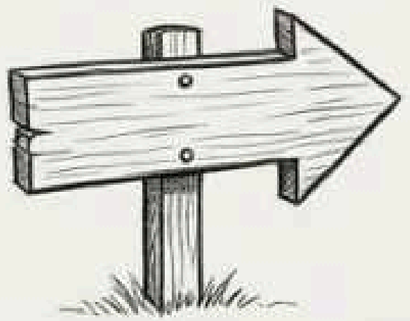
Adding heat to drastically accelerate chemical changes.

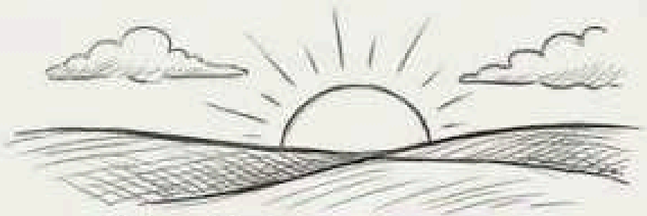
Refrigeration



Cooling lowers k, slowing down the kinetic decision of food spoilage.

The Horizon





Next Lecture: Chemical Equilibrium.

We know how FAST a reaction goes... but how FAR will it go?