

Temperature Dependence of Rate (Arrhenius Equation)

CONCEPT 9 — Master the energetic, statistical, and mathematical models of reaction rates.

Abstract

While thermodynamics tells us if a reaction is feasible, chemical kinetics tells us how fast it occurs. A major factor controlling this speed is temperature.

Historically, it was observed that for many chemical reactions, **a 10° rise in temperature nearly doubles the rate constant.**

In this module, we will deconstruct the Arrhenius Equation to understand the physical and mathematical reasons behind this profound temperature dependence.



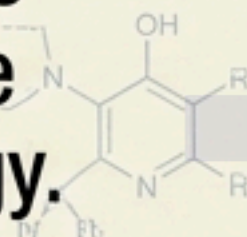
Key Vocabulary

The language of temperature-dependent kinetics



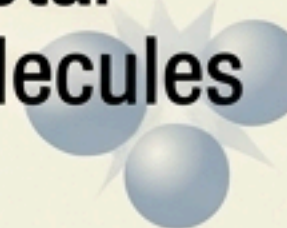
Arrhenius Equation

The mathematical model $k = Ae^{-E_a/RT}$ that quantitatively relates the rate constant of a reaction to absolute temperature and activation energy.



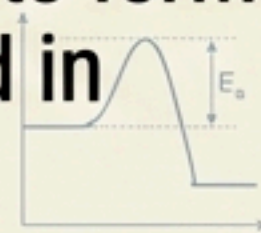
Frequency Factor (A)

Also called the pre-exponential factor. It represents the frequency of total collisions between reactant molecules with the proper orientation.



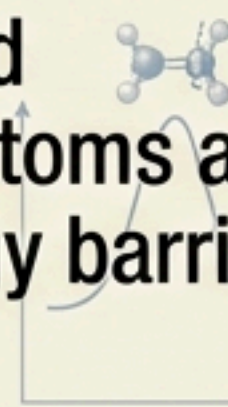
Activation Energy (E_a)

The minimum extra amount of energy absorbed by reactant molecules to form the activated complex. Measured in J/mol.



Activated Complex

The highly unstable, short-lived intermediate arrangement of atoms at the peak of the potential energy barrier during a reaction.



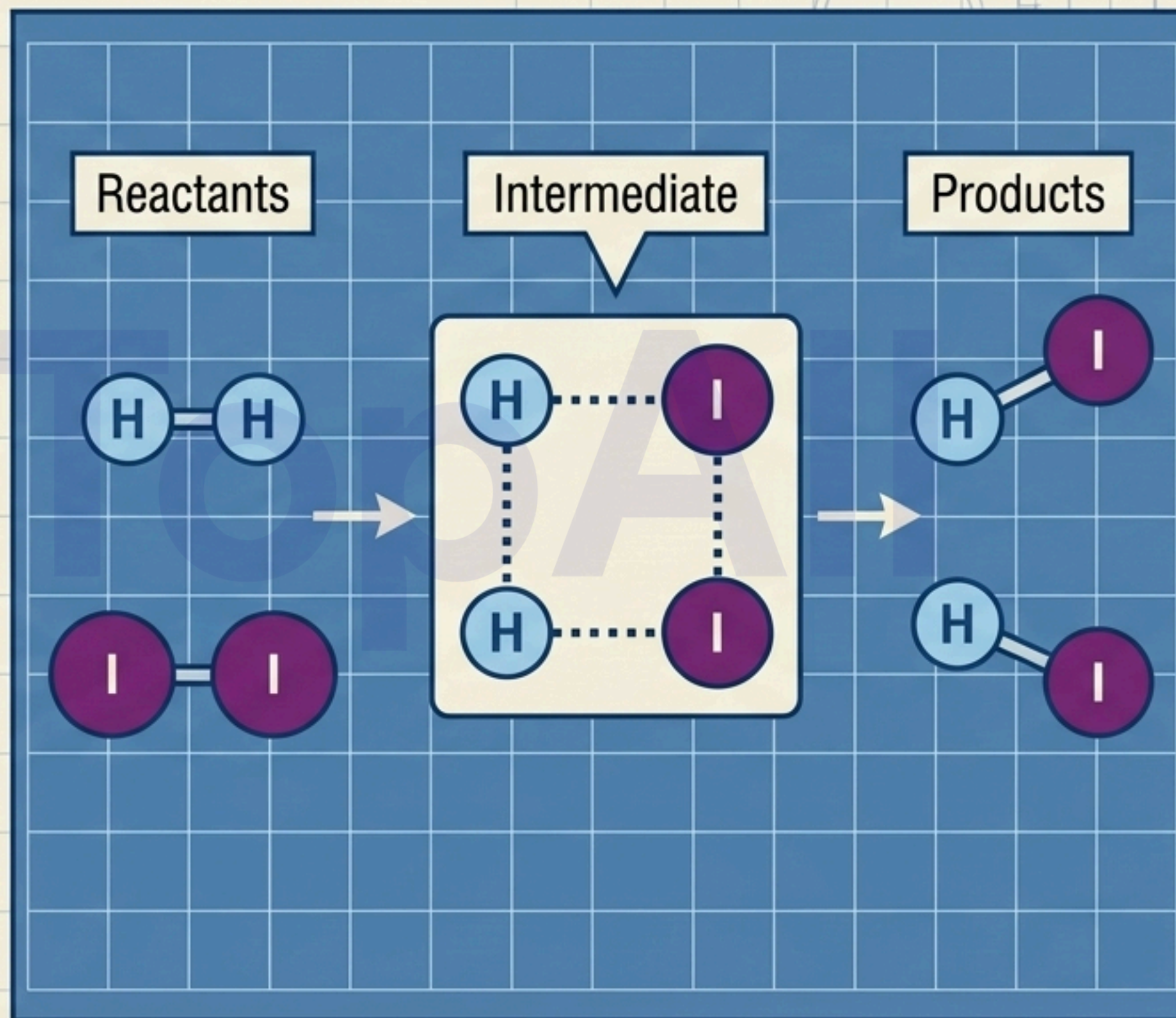
Visual Anchor: The Reaction Mechanism

The physical reality of a collision

According to Arrhenius, a reaction only occurs when molecules collide with sufficient energy to form an unstable intermediate.

The $\text{H}_2 + \text{I}_2$ Example:

1. **Reactants:** Stable molecules approach each other.
2. **Intermediate (Activated Complex):** Old bonds stretch and weaken, while new bonds begin to form. This state exists for a fraction of a second.
3. **Products:** The complex breaks apart to form the final stable molecules (2HI).





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$$E=mc^2$$



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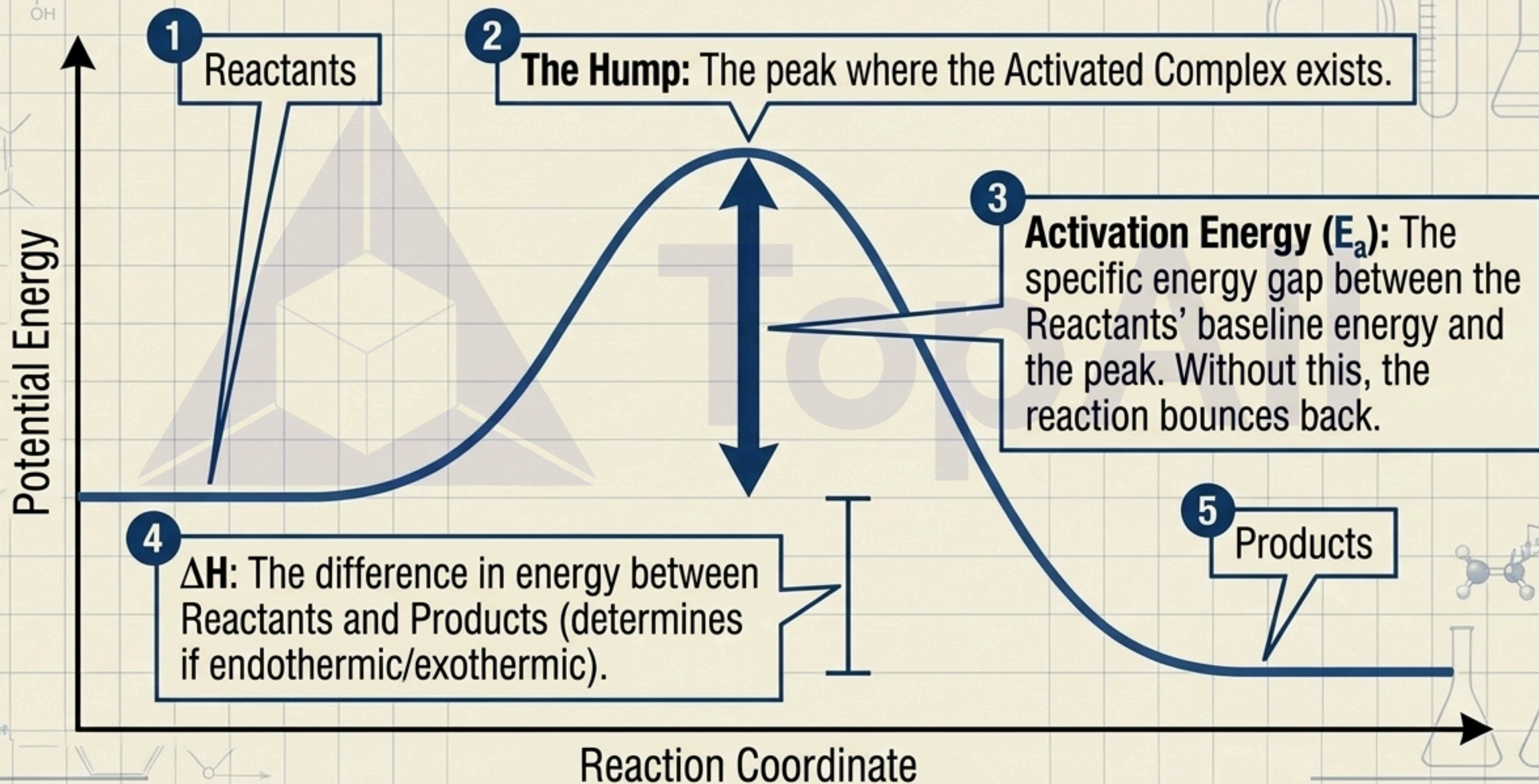
Key Highlights

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- » Get Clear Explanations, Not Just Marks
- » Track Your Performance Smarter



Visual Anchor: Potential Energy Diagram

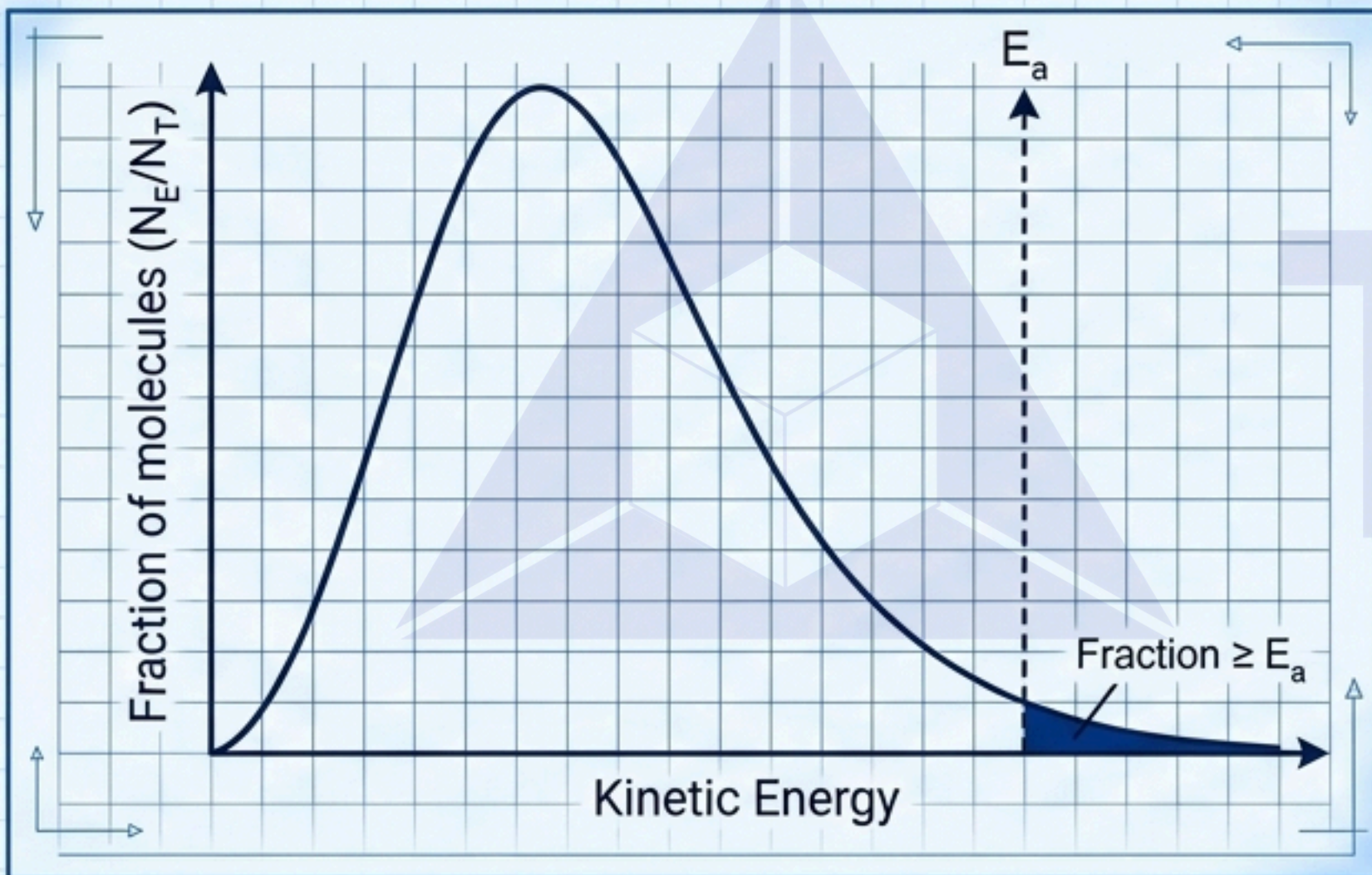
The energy barrier of a reaction



Visual Anchor: Maxwell-Boltzmann Distribution Shift

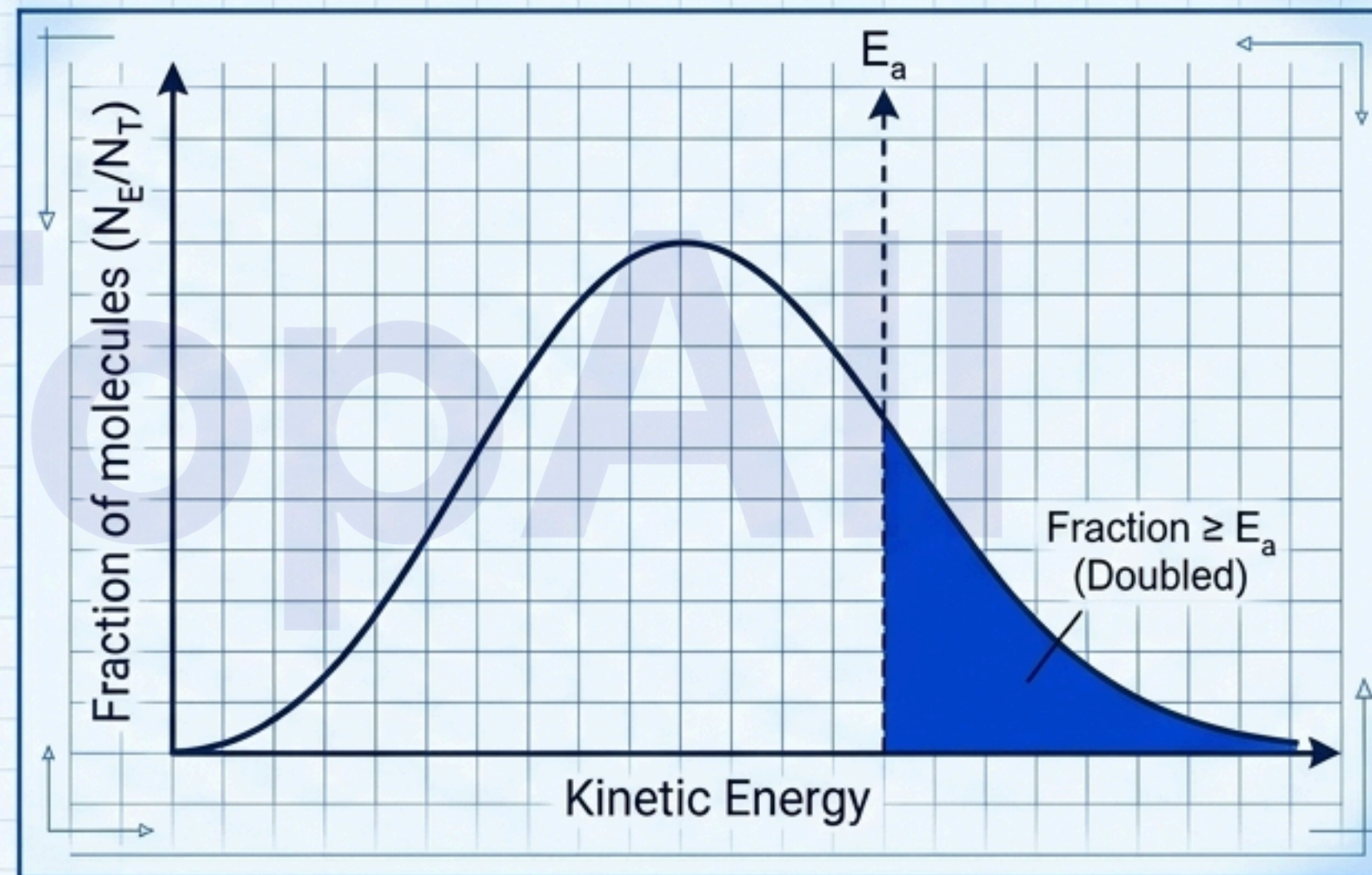
Why a 10° rise doubles the reaction rate

At Temperature t



The peak shows the most probable kinetic energy. Only a small fraction of molecules (shaded area) possess kinetic energy $\geq E_a$.

At Temperature $t + 10$



The curve broadens and shifts right. The peak height drops, but crucially, the shaded area past E_a practically doubles. The fraction of effective collisions doubles, hence the rate doubles.

Formula Bank: The Arrhenius Equation

Three forms for three different applications



1. The Exponential Form

$$k = Ae^{\frac{-E_a}{RT}}$$

Best for:

Conceptual understanding.

Shows that rate constant (k) relies on the collision frequency (A) and the specific fraction of molecules with enough energy ($e^{-E_a/RT}$).

2. The Linear (Graphing) Form

$$\ln k = -\frac{E_a}{RT} + \ln A$$

Best for:

Graphical analysis. By taking the natural log of both sides, it maps perfectly to the equation of a straight line ($y = mx + c$).

3. The Two-Temperature Form

$$\log \frac{k_2}{k_1} = \frac{E_a}{2.303R} \left[\frac{T_2 - T_1}{T_1 T_2} \right]$$

Best for:

Numerical problems. Evaluates how a specific change in temperature ($T_1 \rightarrow T_2$) affects the rate constant ($k_1 \rightarrow k_2$). Note the 2.303 constant from converting $\ln$ to $\log_{10}$.

Visual Anchor: The Linear Plot

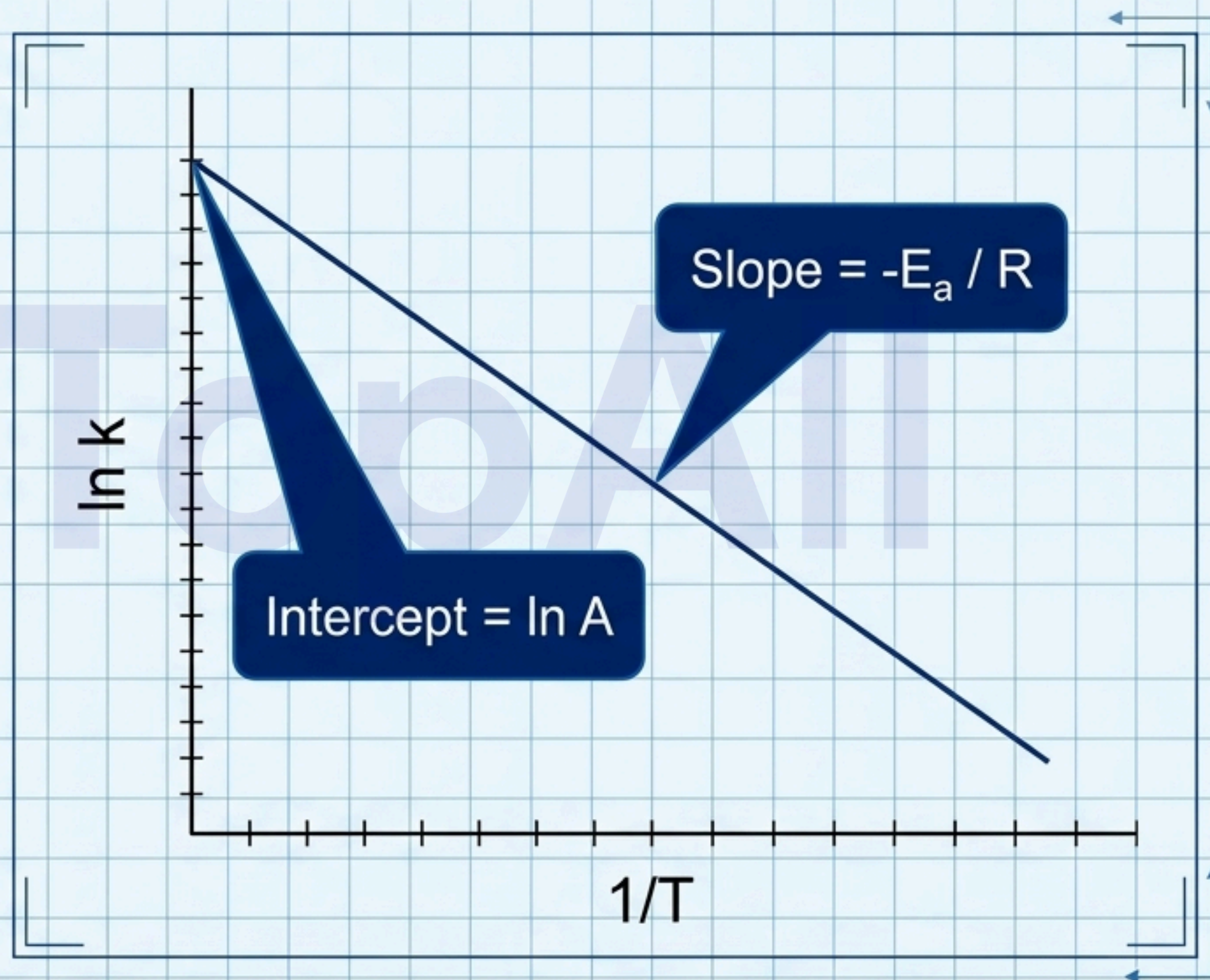
Extracting activation energy from experimental data

Equation Mapping

$$y = mx + c$$

$$\ln k = (-E_a / R) (1 / T) + \ln A$$

- **Y-Axis:** $\ln k$
- **X-Axis:** $1/T$
- **Slope (m):** Evaluates to $-E_a / R$. Because E_a and R are positive, the slope is always negative.
- **Y-Intercept (c):** Evaluates to $\ln A$.



Worked Example: Calculating E_a and A

Based on NCERT Example 3.9 parameters

1. Identify the Data

$$k_1 = 0.02 \text{ s}^{-1} \text{ at } T_1 = 500\text{K}$$

$$k_2 = 0.07 \text{ s}^{-1} \text{ at } T_2 = 700\text{K}$$

$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$$



2. Apply Two-Temperature Formula

$$\log\left(\frac{0.07}{0.02}\right) = \left(\frac{E_a}{2.303 \times 8.314}\right) \times \left[\frac{700 - 500}{500 \times 700}\right]$$



3. Solve for E_a

$$\log(3.5) = \frac{E_a}{19.147} \times \left[\frac{200}{350000}\right]$$

$$0.544 = E_a \times 3.0 \times 10^{-5}$$

$$E_a \approx 18.13 \text{ kJ/mol}$$



4. Solve for A (using T_1 data)

$$\log(0.02) = \log A - \frac{18130}{19.147 \times 500}$$

$$-1.699 = \log A - 1.89 \rightarrow \log A = 0.191$$

$$A = 1.55 \text{ s}^{-1}$$



Trap Alert: Common Exam Pitfalls

Avoid these frequent calculation errors in kinetics

TRAP 1: The Missing Negative Sign

The Error

When extracting E_a from a graph of $\ln k$ vs $1/T$, students often incorrectly set **Slope** = E_a / R .



The Fix

The slope is **ALWAYS** $-E_a / R$. If a graph yields a slope of -5000K , then $-5000 = -E_a / R$, meaning $E_a = 5000 \times R$. Activation energy must end up as a positive value.

TRAP 2: Reversing T_1 and T_2

The Error

In the two-point formula, students frequently write $T_1 - T_2$ in the numerator by mistake, resulting in a negative E_a .

$$\left[\frac{T_2 - T_1}{T_1 - T_2} \right]$$

The Fix

The temperature term is $\left[\frac{T_2 - T_1}{T_1 T_2} \right]$.

Always map k_2 with T_2 on the top, and ensure the larger temperature is assigned to T_2 so the numerator remains positive.

Key Takeaways

Summary of Temperature Dependence

The 10-Degree Rule:

A 10° rise in temperature nearly doubles the reaction rate because it practically doubles the fraction of molecules possessing kinetic energy greater than E_a .

Arrhenius Components:

In $k = Ae^{-E_a/RT}$, A represents collision frequency, and the exponential term $e^{-E_a/RT}$ represents the fraction of successful collisions.

Graphical Analysis:

Plotting $\ln k$ against $1/T$ yields a straight line. The activation energy is derived directly from the slope ($-E_a/R$).

Activation Energy (E_a):

The mandatory energy investment required to form the highly unstable activated complex at the peak of the reaction coordinate.



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