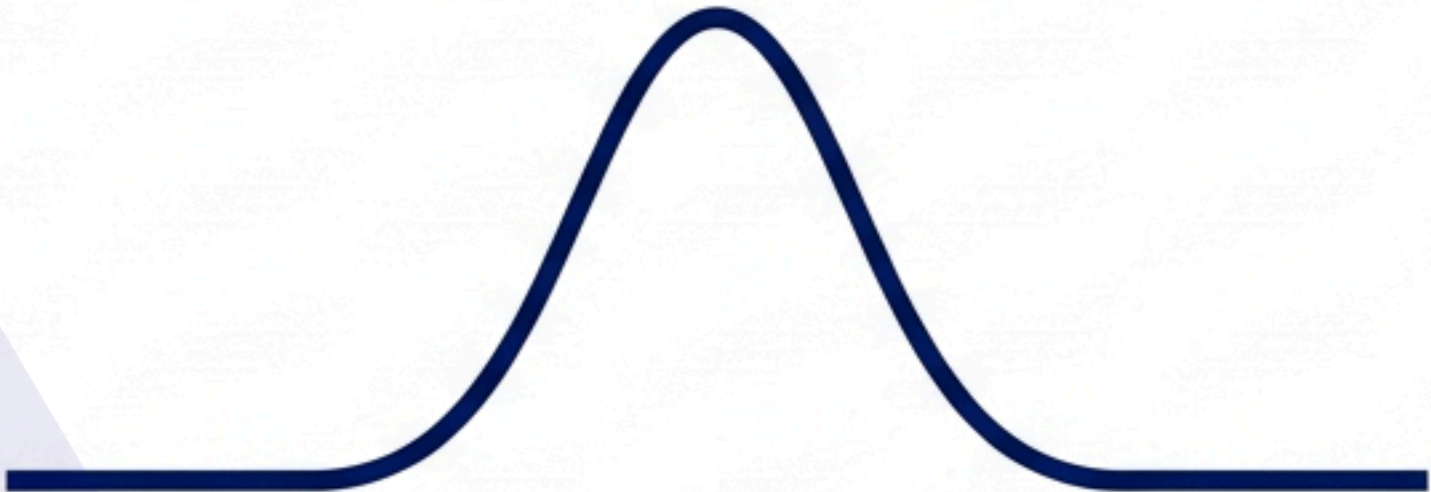




Chemical Kinetics • Unit 3

3.4 Temperature Dependence of the Rate of a Reaction

Understanding how thermal energy drives molecular collisions, activation barriers, and the Arrhenius Equation.

Heat Accelerates Chemical Change

**+10°C =
2x Speed** 

For most chemical reactions, a rise in temperature by 10° nearly doubles the rate constant (k).

Thermodynamics vs. Kinetics:

Thermodynamics dictates if a reaction is feasible (e.g., diamond converting to graphite). **Kinetics** tells us how fast it happens.

The Energy Factor: Temperature is a primary driver because it directly alters the kinetic energy of the system.

The Exponential Reality: The relationship between temperature and reaction rate is not linear—it is **exponential**. To understand exactly how, we rely on the Arrhenius Equation.

The Arrhenius Equation

Proposed by J.H. van't Hoff and physically justified by Svante Arrhenius, this equation accurately predicts how the rate constant (k) scales with temperature (T).

$$k = A \cdot e^{(-E_a / RT)}$$

k (Rate Constant):
The speed factor of the reaction.

A (Frequency Factor): A constant specific to the reaction, related to collision frequency.

E_a (Activation Energy): The minimum energy required to form products (J mol^{-1}).

R (Gas Constant):
 $8.314 \text{ J K}^{-1} \text{ mol}^{-1}$.

T (Temperature):
Absolute temperature in Kelvin (K).

$e^{(-E_a / RT)}$: The crucial fraction of molecules that have kinetic energy greater than E_a .



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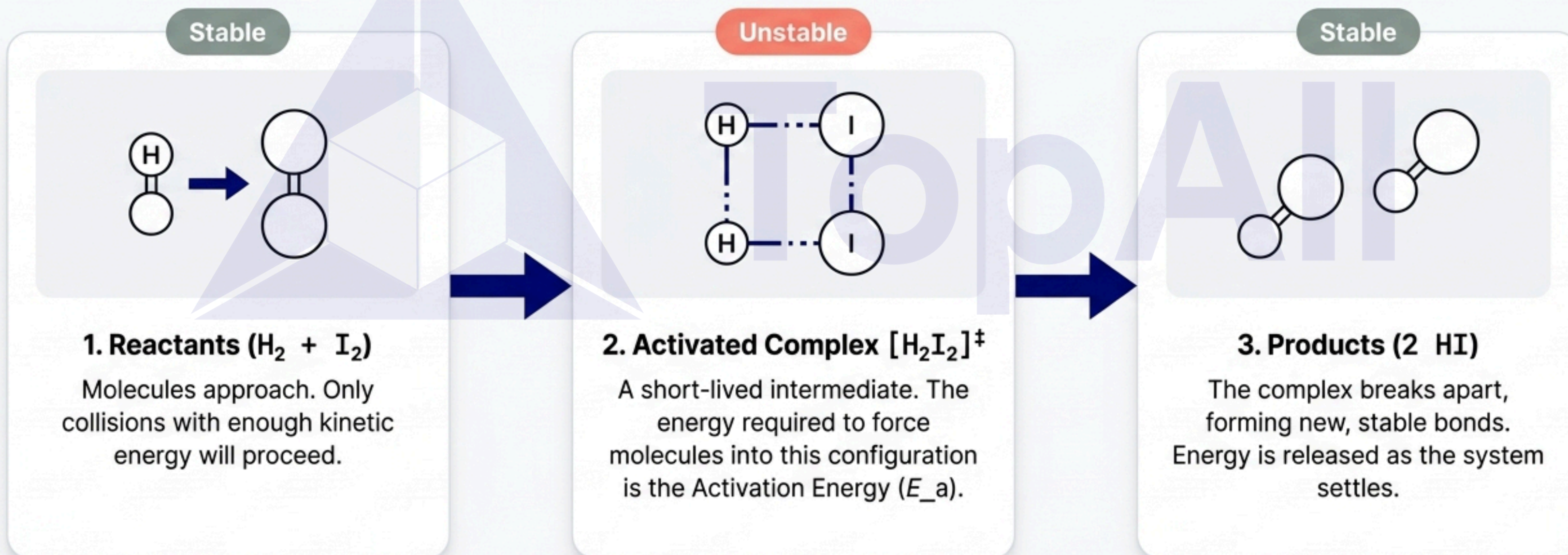
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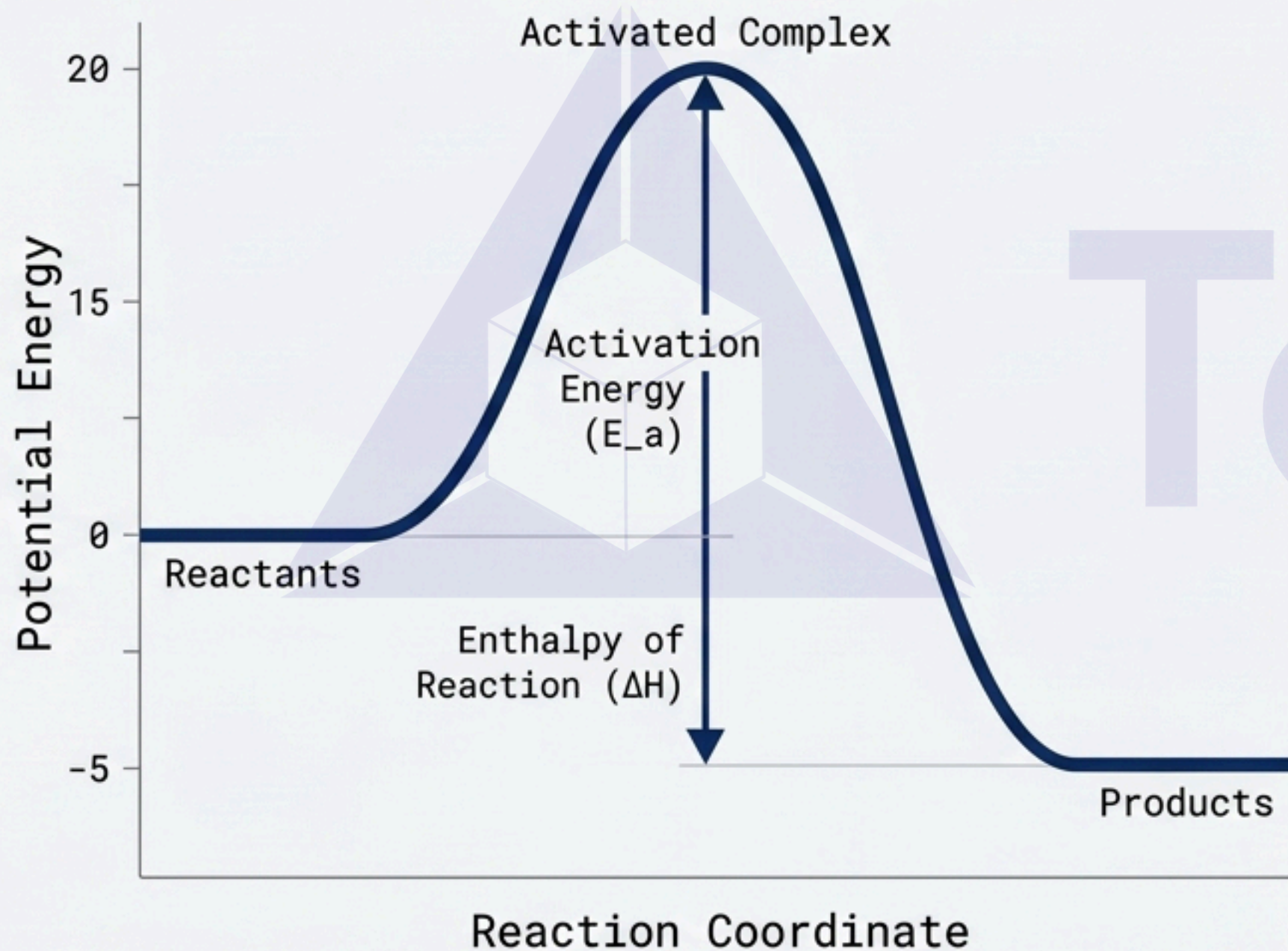


Crossing the Energy Barrier

According to collision theory, molecules must collide with sufficient energy and proper orientation to react. They form a temporary, high-energy state.



The Potential Energy Profile



The Energy Hill:

The reaction coordinate represents the energy profile. Molecules must "climb" the activation energy (E_a) hill to successfully react.

Enthalpy (ΔH):

If products have lower potential energy than reactants (as shown in the graph), energy is released into the surroundings (ΔH is negative/exothermic).

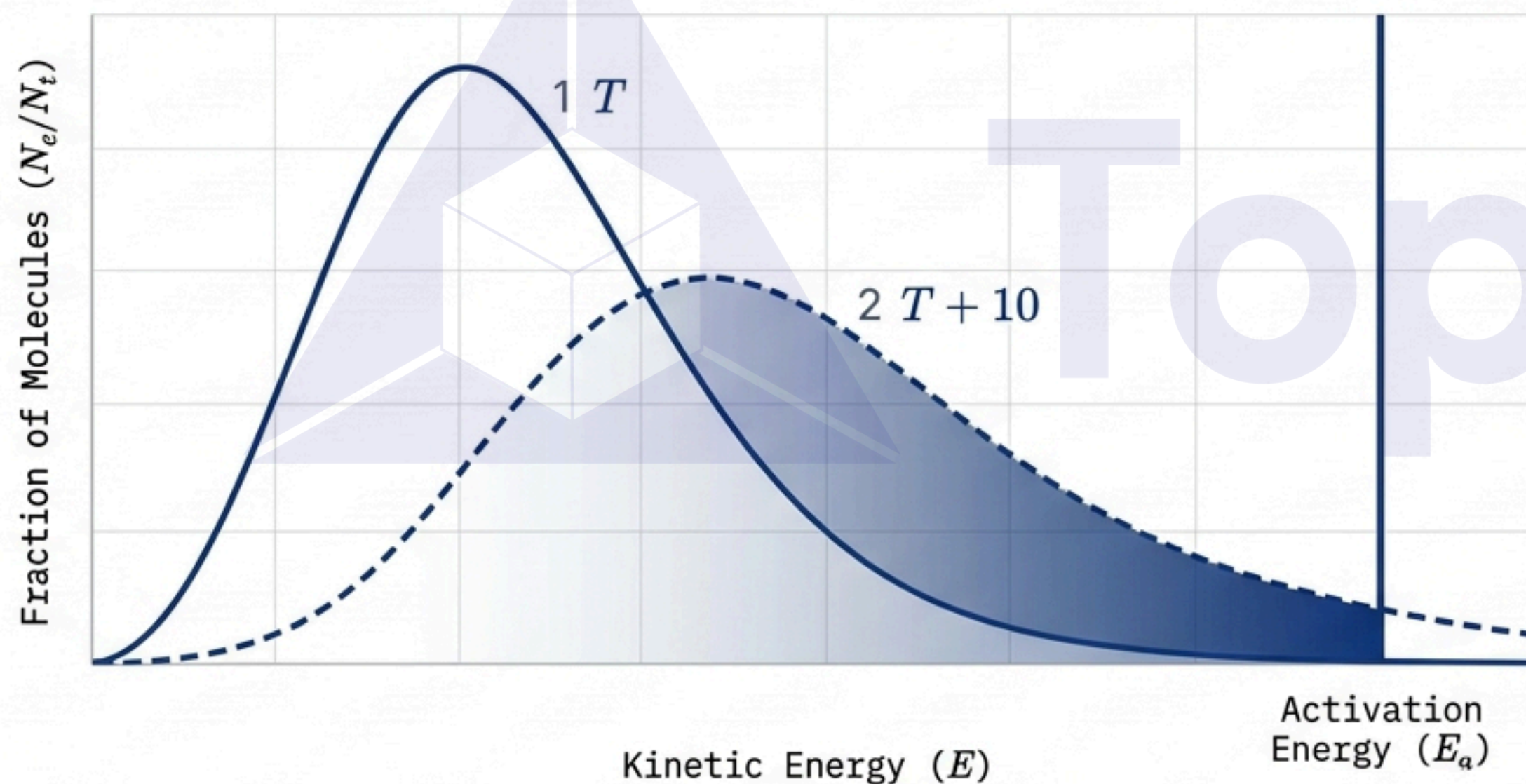
Exam Tip

Catalyst Effect:

A catalyst does NOT change the starting or ending energy. It accelerates the reaction by providing an alternate pathway, lowering the E_a peak.

Why a 10° Rise Doubles the Rate

Not all molecules possess the same energy. The Maxwell-Boltzmann distribution visualizes the fraction of molecules at varying kinetic energies.

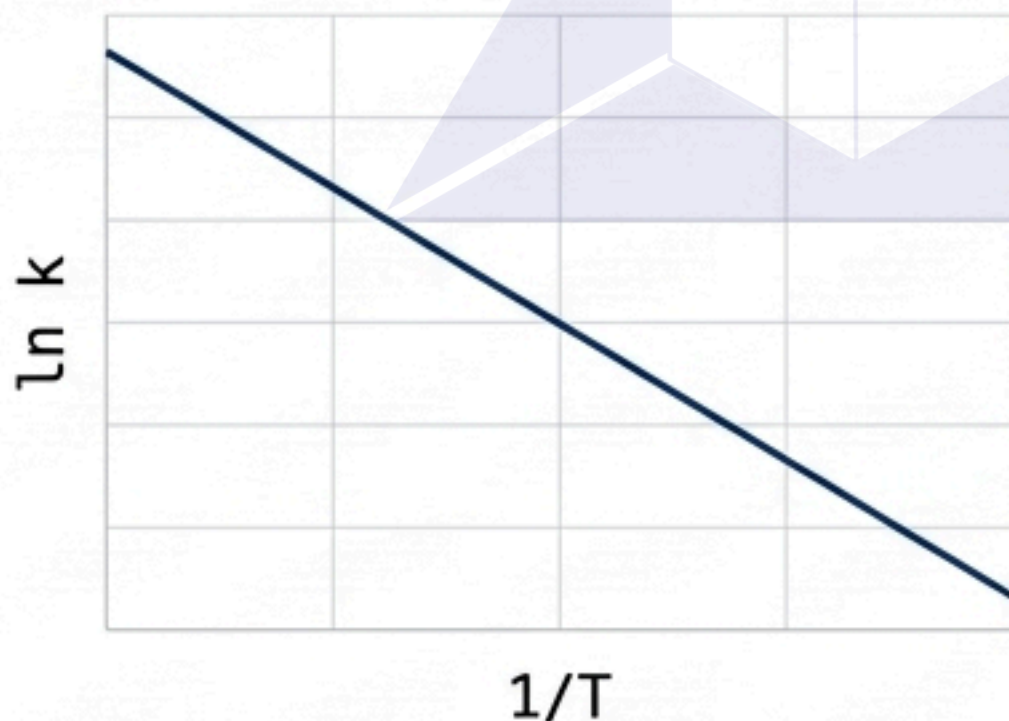


****The Area Under the Curve****
When temperature rises from T to $T+10$, the peak shifts right and broadens. The area representing molecules with energy $\geq E_a$ ****doubles****. The Arrhenius exponent ($e^{(-E_a/RT)}$) calculates this exact shaded fraction!

Determining E_a and A Graphically

By taking the natural logarithm ($\ln$) of the Arrhenius equation, we transform an exponential curve into a linear equation ($y = mx + c$).

$$\begin{aligned} y &= m \cdot x + c \\ \ln k &= - (E_a / R) \cdot (1/T) + \ln A \end{aligned}$$



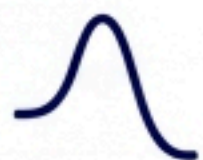
Slope (m): $= -E_a / R$

By measuring the downward slope of experimental data, we can easily calculate the specific Activation Energy.

Y-Intercept (c): $= \ln A$

Where the line crosses the y-axis gives us the frequency factor.

The Core Physics of Reaction Rates



The Barrier: Activation Energy (E_a)

Reactions require an initial energy investment to form an unstable activated complex. Lowering this barrier (via catalyst) accelerates the rate.



The Statistics: Maxwell-Boltzmann

Temperature dictates the distribution of kinetic energy. Heating the system flattens the curve, pushing a greater fraction of molecules past the E_a threshold.



The Math: Arrhenius Equation

$k = A e^{(-E_a/RT)}$
mathematically unites the macroscopic temperature (T) with the microscopic barrier (E_a) to predict reaction speed precisely.

Temperature does not lower the energy barrier; it increases the fraction of molecules energetic enough to cross it.



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