CHEMICAL KINETICS

Activation Energy

The energy barrier every reaction must climb, what catalysts really do, the Arrhenius picture, and a 10-question practice set with full solutions.

SUBJECT	LEVEL	INCLUDES
Chemistry	High school and early college	Practice set, answer key, revision sheet

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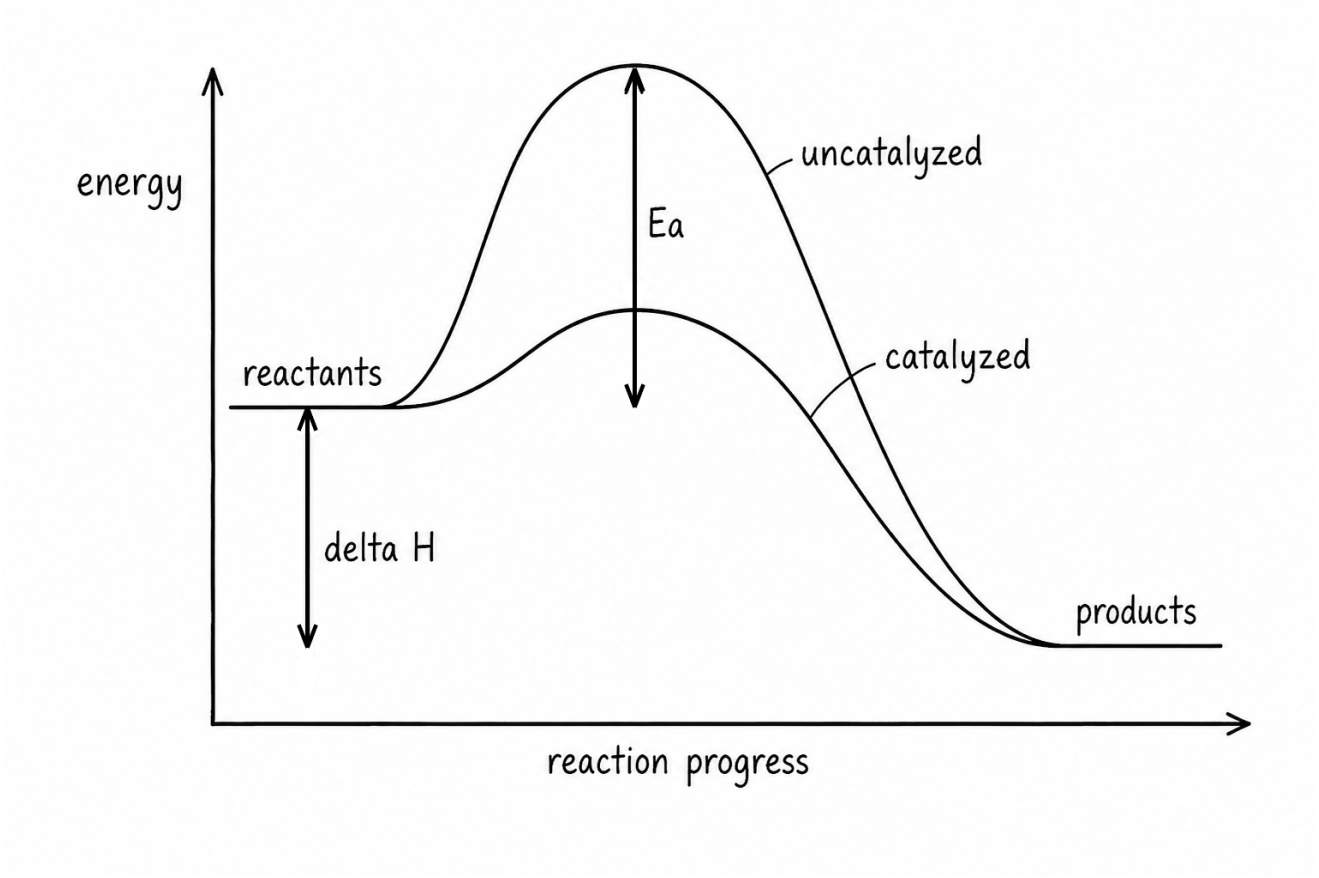
INSIDE

Activation energy is the minimum amount of energy that reactant molecules must have to undergo a chemical reaction. It's the height of the energy hill that reactants must climb before they can roll down to become products. The concept, formalized by Svante Arrhenius in 1889, explains why most reactions need heating to start, why catalysts speed reactions up, and why reaction rates depend exponentially on temperature.



Activation energy is the minimum energy barrier reactants must overcome. A catalyst lowers the barrier.

The Energy Barrier

In an energy diagram (energy vs reaction coordinate), reactants sit at one energy and products at another. Between them is a peak, the transition state, corresponding to the moment when old bonds are partially broken and new bonds are partially formed. The energy difference between reactants and this peak is the activation energy E_a.

For an exothermic reaction, the products are lower than the reactants, but molecules still must climb E_a first. Once over the barrier, energy is released as heat. For an endothermic reaction, products are higher than reactants, and E_a is even larger.

The Arrhenius Equation

The temperature dependence of a reaction rate constant k is:

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

where:

- A is the pre-exponential factor (frequency of collisions with correct orientation).
- E_a is the activation energy in joules per mole.
- $R = 8.314\ \text{J/(mol·K)}$ is the gas constant.
- T is absolute temperature in kelvins.

The exponential $e^{-E_a/(RT)}$ is the fraction of molecules with enough energy to react. At low temperature, very few molecules clear the barrier; at high temperature, many more do. A 10°C rise often doubles or triples a reaction rate, exactly because that exponential is so sensitive to T.

Catalysts: Lowering the Barrier

A catalyst speeds a reaction by providing an alternative pathway with a lower activation energy. The reactants and products are unchanged; the catalyst itself emerges unaltered at the end of each cycle. The lower E_a means a much larger fraction of molecules can react at any given temperature.

Example: hydrogen peroxide decomposes very slowly on its own ($E_a \approx 76\ \text{kJ/mol}$), but with catalase enzyme present, E_a drops to ~8 kJ/mol. The rate at body temperature rises by roughly $e^{(76-8)\cdot 10^3/(8.314\cdot 310)} \approx 10^{11}$, eleven orders of magnitude faster.

Worked Example: Rate Doubling with Temperature

A reaction has $E_a = 50$ kJ/mol. By how much does the rate change going from 25°C (298 K) to 35°C (308 K)?

Using $\ln(k_2/k_1) = -E_a/R \cdot (1/T_2 - 1/T_1)$:

$$\ln(k_2/k_1) = -\frac{50000}{8.314}\left(\frac{1}{308} - \frac{1}{298}\right) = -6015 \cdot (-1.089 \times 10^{-4}) = 0.655$$

So $k_2/k_1 = e^{0.655} \approx 1.93$. A 10°C rise nearly doubles the rate, the classic 'rule of thumb' that holds when E_a is in the 50 kJ/mol range.

Why Activation Energy Exists

For a chemical reaction to occur, reactant molecules must collide with enough kinetic energy to break or distort existing bonds, and they must collide in an orientation that allows new bonds to form. Most molecular collisions in everyday conditions fail one or both tests. The activation energy is the energetic price of getting past the unstable transition-state geometry.

Some reactions have negligible E_a, for example, radical recombination in a flame or acid-base proton transfers in solution. These proceed at the collision rate. Most reactions have E_a values between 40 and 200 kJ/mol, which is why heating is often required.

Applications

- **Reaction rate prediction.** Measuring rate at two or more temperatures and plotting $\ln k$ versus $1/T$ gives a straight line whose slope is $-E_a/R$, the standard way activation energies are determined experimentally.
- **Industrial catalysis.** Nearly every large-scale chemical (ammonia, sulfuric acid, ethylene, gasoline, polyethylene) is made using catalysts that drop E_a and let reactions run at practical temperatures.
- **Enzymes.** Living cells use thousands of protein enzymes, each tailored to lower the E_a of a specific

reaction. Without enzymes, life's chemistry would be too slow to support metabolism.

- **Food preservation.** Refrigeration slows spoilage reactions exponentially through the Arrhenius equation; a 10°C drop typically cuts microbial growth rate by 2-3×.
- **Combustion ignition.** Petrol needs a spark and diesel needs compression heating to clear the activation energy barrier; once over it, combustion releases enough energy to keep the reaction going.

Related study notes: Catalysis, Enthalpy, Redox Reactions, Chemical Kinetics.

Practice Questions

Work each question before reading its solution. The set runs from direct recall and substitution to the applied questions that exams actually use to separate grades.

Q1. Define activation energy, and explain why thermodynamically favorable reactions can still be slow.

SOLUTION

E_a is the minimum energy colliding molecules need to reach the transition state and react. Favorability ($\Delta G < 0$) says where equilibrium lies, not how fast it arrives: gasoline plus oxygen is enormously favorable yet sits inert until a spark supplies the barrier energy. Rate and destination are separate questions.

Q2. Sketch in words the energy profile of an exothermic reaction, marking E_a and ΔH.

SOLUTION

The curve starts at the reactant level, climbs to the transition-state peak, and descends to a products level below the start. E_a is the climb from reactants to the peak; ΔH is the net drop from reactants to products. The two are independent: a very exothermic reaction can still carry a huge barrier.

Q3. What does a catalyst change on the energy profile, and what does it leave untouched?

SOLUTION

It provides an alternative pathway with a lower peak, cutting E_a; enzymes and industrial catalysts both work this way. It does not change reactant or product energies, so ΔH and the equilibrium constant stay exactly as they were, and the catalyst emerges unconsumed. Catalysts change the speed of arriving at equilibrium, never its position.

Q4. Why does a 10°C temperature rise roughly double many reaction rates, when the average molecular energy rises only about 3%?

SOLUTION

Reacting molecules come from the high-energy tail of the Boltzmann distribution, the fraction $e^{-E_a/RT}$. A small shift of the whole distribution swells that exponential tail disproportionately: for $E_a \approx 50$ kJ/mol near room temperature, the tail roughly doubles per 10°C. The average barely moves; the eligible minority explodes.

Q5. State the Arrhenius equation and identify each factor.

SOLUTION

$k = Ae^{-E_a/RT}$. The exponential is the fraction of collisions energetic enough to clear the barrier; A, the pre-exponential factor, counts collision frequency and how often geometry is right; R is the gas constant and T absolute temperature. Taking logarithms gives $\ln k = \ln A - E_a/RT$, a straight line in $1/T$ whose slope measures E_a.

Q6. A reaction's rate constant doubles between 300 K and 310 K. Estimate E_a.

SOLUTION

$\ln 2 = \frac{E_a}{R}\left(\frac{1}{300} - \frac{1}{310}\right)$. The bracket is $\frac{10}{93{,}000} = 1.075 \times 10^{-4}$, so $E_a = \frac{0.693 \times 8.314}{1.075 \times 10^{-4}} \approx 53{,}600$ J/mol ≈ 54 kJ/mol. The rule-of-thumb doubling corresponds to a barrier near 50 kJ/mol, exactly why the rule is only a rule of thumb.

Q7. Why does a refrigerator slow food spoilage, in activation-energy terms?

SOLUTION

Spoilage is chemistry: microbial enzyme reactions and oxidation, all with barriers. Cooling from 25°C to 4°C shrinks the $e^{-E_a/RT}$ fraction severalfold for typical barriers, so the same reactions run several times slower. The food is not sterilized, merely decelerated, which is why refrigeration buys days, not months.

Q8. Enzymes accelerate reactions by up to 10^{17}. Translate a rate gain of 10^6 into an effective E_a reduction at 310 K.

SOLUTION

Rate ratio $= e^{\Delta E_a/RT}$, so $\Delta E_a = RT \ln 10^6 = 8.314 \times 310 \times 13.8 \approx 35{,}600$ J ≈ 36 kJ/mol. A modest-sounding barrier cut of 36 kJ/mol delivers a millionfold speedup: the exponential converts small energy engineering into enormous rate leverage.

Q9. A match ignites paper, yet the fire continues without more matches. Explain the energetics.

SOLUTION

The match supplies the initial E_a for combustion. Once burning starts, the exothermic reaction releases far more energy than the barrier requires, and that heat activates neighboring molecules in a chain: the reaction pays its own activation bill thereafter. Self-sustaining combustion is a reaction financing its E_a from its ΔH.

Q10. Two reactions have identical $\Delta H = -100$ kJ/mol, but reaction 1 has $E_a = 40$ and reaction 2 has $E_a = 120$ kJ/mol. Compare their behavior at room temperature and their responses to heating.

SOLUTION

Reaction 1 proceeds readily; reaction 2 is effectively frozen, its rate smaller by a factor $e^{80{,}000/RT} \approx 10^{14}$. Heating helps reaction 2 proportionally more, since the rate's temperature sensitivity grows with E_a: high-barrier reactions are the ones that "switch on" dramatically when heated. Same destination, wildly different journeys.

Answer Key

QUICK ANSWERS

Q1 the barrier height; kinetics ≠ thermodynamics **Q2** E_a to the peak; ΔH between levels **Q3** lowers E_a only; ΔH and K unchanged **Q4** the $e^{-E_a/RT}$ tail grows exponentially **Q5** $k = Ae^{-E_a/RT}$

Q6 $\approx$ 54 kJ/mol **Q7** lower T starves the barrier-clearing tail **Q8** $\approx$ 36 kJ/mol shaved **Q9** released heat keeps clearing the barrier **Q10** 2 is $\sim 10^{14}\times$ slower but more heat-sensitive

Revision Sheet

The idea

E_a = climb to the transition state. Sets speed; ΔH sets the drop; they are independent.

Catalysts

New pathway, lower peak. ΔH, equilibrium, and the catalyst itself unchanged.

Arrhenius

$k = Ae^{-E_a/RT}$; $\ln k$ vs $1/T$ is linear with slope $-E_a/R$.

Everyday cases

Spark ignites fuel; fridge slows spoilage; enzymes cut barriers by tens of kJ/mol.

Rules of thumb

~50 kJ/mol barrier: rate doubles per 10°C. Bigger E_a, stronger T-sensitivity.

Self-sustaining fire

Exothermic release re-pays E_a for the neighbors: combustion chains.

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