

Activation Energy and Catalysts

What is activation energy and how do catalysts speed up reactions?

●●○ Intermediate

Kinetics

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Before you start: it helps to understand [Reaction Rates and Rate Laws](#).

What is it?

For particles to react, they must **collide** with enough energy to break existing bonds and start forming new ones. The minimum energy needed is the **activation energy**, E_a (usually in kJ/mol).

An **energy profile** shows the energy of the reacting particles as the reaction proceeds. The highest point is the **transition state** (or activated complex), an unstable arrangement partway between reactants and products. E_a is the height of this “energy hill” above the reactants.

🔑 KEY IDEA

The **higher** the activation energy, the **slower** the reaction. Raising the temperature gives more molecules enough energy to get over the hill; a **catalyst** provides a different route with a **lower** hill.

Why does it matter?

- **Food and medicine.** Fridges slow spoilage, and medicines are stored cool, because lower temperatures slow reactions.
- **Industry.** Catalysts let chemical plants run faster at lower temperatures, saving energy (iron in the Haber process; platinum in catalytic converters).
- **Life.** Enzymes are biological catalysts that make the reactions of life fast enough at body temperature.

How does it work?

1. Collision theory

A collision leads to reaction only if the particles:

1. collide with energy **at least equal to** E_a , and
2. have a suitable **orientation**.

Most collisions fail. Anything that increases the number of *successful* collisions per second increases the rate.

2. Why temperature has such a large effect

The **Maxwell-Boltzmann distribution** shows how kinetic energy is spread among molecules. At a higher temperature the curve flattens and shifts to the right, and the area beyond E_a (the fraction of molecules that can react) grows **much more** than the average energy does. A rise of only 10 K can roughly double the rate of many reactions.

3. The Arrhenius equation

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

where A is the frequency factor (collision frequency and orientation), $R = 8.314 \text{ J}/(\text{mol}\cdot\text{K})$ and T is in kelvin. Comparing two temperatures:

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Always convert E_a to **J/mol** to match the units of R .

4. Catalysts

A **catalyst** speeds up a reaction by providing an **alternative pathway with a lower activation energy**. It takes part in the reaction but is **regenerated**, so it is not used up.

- It speeds up the forward **and** reverse reactions equally, so it does **not** change ΔH , the equilibrium position or K .
- **Homogeneous** catalysts are in the same phase as the reactants; **heterogeneous** catalysts (often solids such as iron or platinum) work at a surface.

∞ THINK OF IT LIKE THIS

Getting from one valley to the next over a high mountain pass is slow; few hikers have the energy. A tunnel (the catalyst) offers a lower route, so many more get through each hour, but both valleys stay at the same heights: the start and finish (reactants and products, and ΔH) don't change.

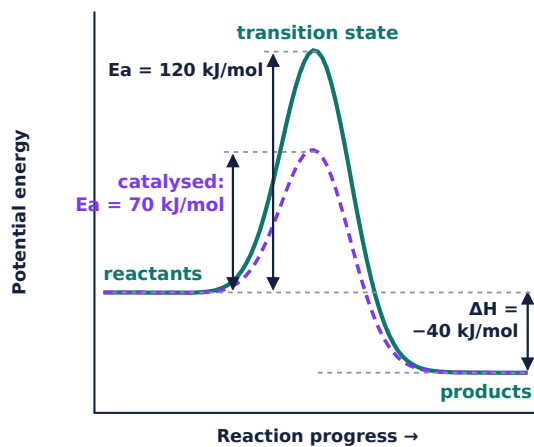
Q MORE PRECISELY

Taking logarithms gives $\ln k = \ln A - \frac{E_a}{R} \cdot \frac{1}{T}$, so a plot of $\ln k$ against $1/T$ (an **Arrhenius plot**) is a straight line with slope $-E_a/R$. That's how activation energies are measured from experimental rate constants.

Visualise it

Energy profile of an exothermic reaction

A catalyst gives a lower-energy pathway; ΔH is unchanged.



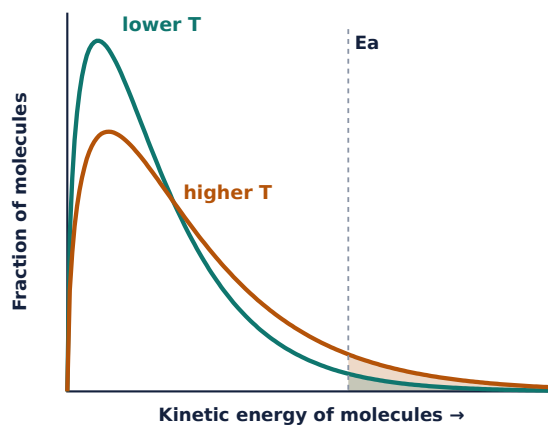
Solid line: uncatalysed. Dashed line: catalysed.

E_a (reverse) = $120 \text{ kJ/mol} + 40 \text{ kJ/mol} = 160 \text{ kJ/mol}$.

A catalyst lowers the activation energy but leaves ΔH unchanged.

Why heating speeds up reactions

Fraction of molecules with energy $\geq E_a$ at two temperatures.



The shaded area beyond E_a is much larger at the higher temperature (here about 2.9 times), so more collisions succeed.

Heating mainly increases the fraction of molecules with energy above E_a .

Worked example

 **WORKED EXAMPLE:** How much faster when 10 K warmer?

Question: A reaction has $E_a = 50.0$ kJ/mol. By what factor does k increase from 298 K to 308 K?

Step 1 Convert: $E_a = 50.0 \text{ kJ/mol} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 5.00 \times 10^4 \text{ J/mol}$

Step 2 Substitute:

$$\begin{aligned} \ln \frac{k_2}{k_1} &= \frac{5.00 \times 10^4 \text{ J/mol}}{8.314 \text{ J/(mol}\cdot\text{K)}} \\ &\quad \times \left(\frac{1}{298 \text{ K}} - \frac{1}{308 \text{ K}} \right) \\ &= 0.655 \end{aligned}$$

J/mol cancel, and $\text{K} \times \text{K}^{-1}$ cancel, so the logarithm is a pure number.

Step 3 $\frac{k_2}{k_1} = e^{0.655} = \mathbf{1.93}$: the rate almost doubles.

 WORKED EXAMPLE: Finding the activation energy

Question: A rate constant is $1.00 \times 10^{-3} \text{ s}^{-1}$ at 300. K and $4.00 \times 10^{-3} \text{ s}^{-1}$ at 320. K. Find E_a .

Step 1 Rearrange: $E_a = \frac{R \ln(k_2/k_1)}{\frac{1}{T_1} - \frac{1}{T_2}}$

Step 2 $\ln \frac{4.00 \times 10^{-3} \text{ s}^{-1}}{1.00 \times 10^{-3} \text{ s}^{-1}} = \ln 4.00 = 1.3863$ (the units of s^{-1} cancel)

Step 3 $\frac{1}{300. \text{ K}} - \frac{1}{320. \text{ K}} = 2.0833 \times 10^{-4} \text{ K}^{-1}$

Step 4 Numerator: $R \ln \frac{k_2}{k_1} = 8.314 \text{ J}/(\text{mol}\cdot\text{K}) \times 1.3863 = 11.526 \text{ J}/(\text{mol}\cdot\text{K})$
(keep extra digits until the end)

Step 5 Divide:

$$\begin{aligned} E_a &= \frac{11.526 \text{ J}/(\text{mol}\cdot\text{K})}{2.0833 \times 10^{-4} \text{ K}^{-1}} \\ &= 5.53 \times 10^4 \text{ J/mol} \\ &= 55.3 \text{ kJ/mol} \end{aligned}$$

K and K^{-1} cancel, leaving J/mol.

Common mistake

 COMMON MISTAKE: Mixing kJ and J

$R = 8.314 \text{ J}/(\text{mol}\cdot\text{K})$ is in **joules**. Using E_a in kJ/mol without converting makes the exponent 1000 times too small.

 COMMON MISTAKE: Using °C in the Arrhenius equation

Temperatures must be in **kelvin**. $\frac{1}{25 \text{ °C}}$ has no physical meaning; use 298 K.

 COMMON MISTAKE: Thinking a catalyst changes ΔH or the yield

A catalyst lowers E_a for both directions. It does not change ΔH , the equilibrium position or K; it only gets the system to equilibrium faster.

NOTATION NOTE

- E_a may be written per mole (kJ/mol) or per molecule (J); chemistry courses use kJ/mol.
- “Activated complex” and “transition state” are used for the top of the energy profile.

Remember this

REMEMBER THIS

- Reaction needs collisions with energy $\geq E_a$ and the right orientation.
- Higher temperature: many more molecules exceed E_a (Maxwell-Boltzmann), so the rate rises steeply.
- $k = Ae^{-E_a/RT}$; use J/mol and kelvin; $\ln(k_2/k_1) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$.
- Catalysts lower E_a , are regenerated, and don't change ΔH or K.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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