Reaction kinetics

A-Level Chemistry

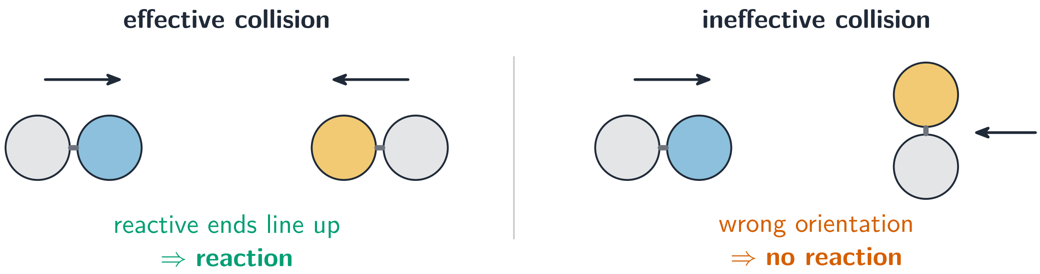
Rate of reaction

The **rate of reaction** 反应速率 is how fast reactants turn into products. We measure it as the change in concentration (or amount) in each unit of time.

To react, particles must **collide** 碰撞. The **collision frequency** 碰撞频率 is how often the particles hit each other. But not every collision leads to a reaction:

- an **effective collision** 有效碰撞 has enough energy *and* the correct direction, so a reaction happens.
- a **non-effective collision** 无效碰撞 does not have enough energy, or the particles hit at the wrong angle, so nothing happens.

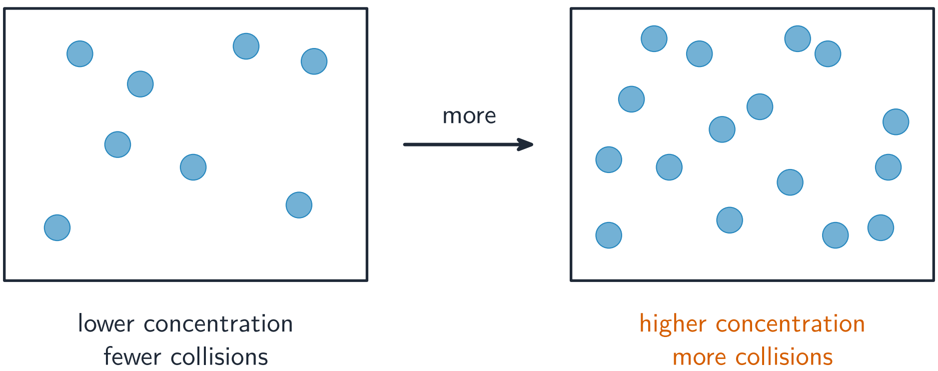
So the rate depends on the **frequency of effective collisions** —how many useful collisions happen each second.



A collision only reacts with the right orientation and enough energy (coloured ends = the reactive part)

Concentration and pressure

If you increase the concentration of a solution (or the pressure of a gas), the particles are packed closer together. They collide more often, so there are more effective collisions each second, and the rate goes up.



More particles in the same volume collide more often, so the rate rises

You can calculate a rate from experimental data —for example, the volume of gas made divided by the time taken.

Temperature and activation energy

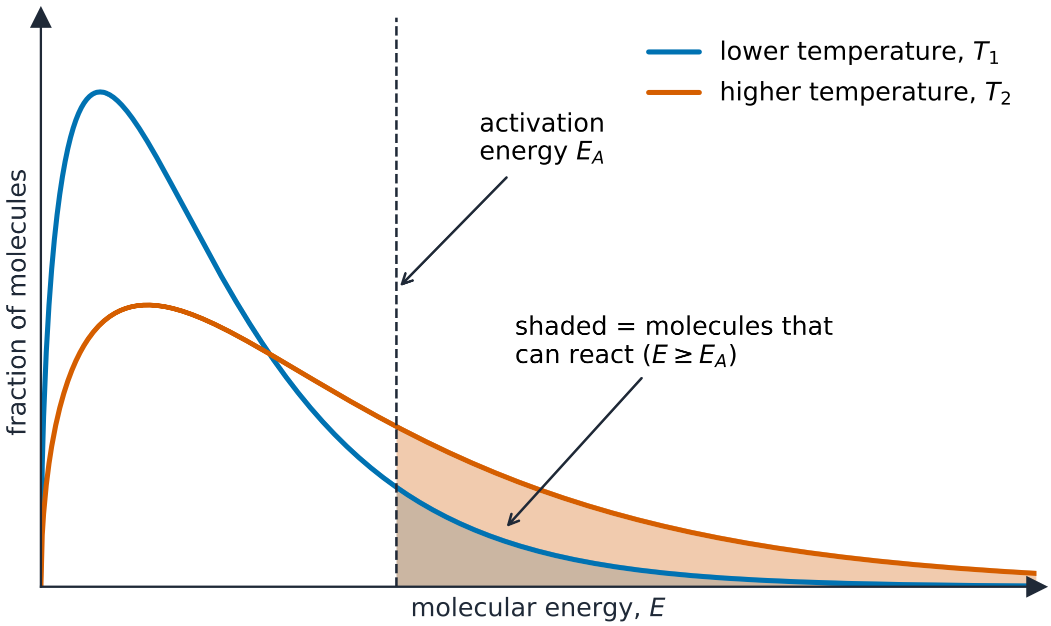
The **activation energy** 活化能 (E_A) is the minimum energy a collision needs in order to be effective.

The **Boltzmann distribution** 玻尔兹曼分布 is a graph showing how the energies of the molecules are spread out at one temperature. The curve starts at the origin, rises to a peak, then falls away in a long tail. The total area under the curve is the total number of molecules. Only the molecules to the right of E_A have enough energy to react.

When you raise the temperature:

- the curve flattens and spreads to the right, so a **much larger fraction** of molecules now have energy greater than E_A.
- the molecules also move faster and collide more often.

The first effect is the bigger one. This is why a small rise in temperature gives a large rise in rate.



At higher temperature the curve spreads to the right, so a larger fraction of molecules can react

Catalysts

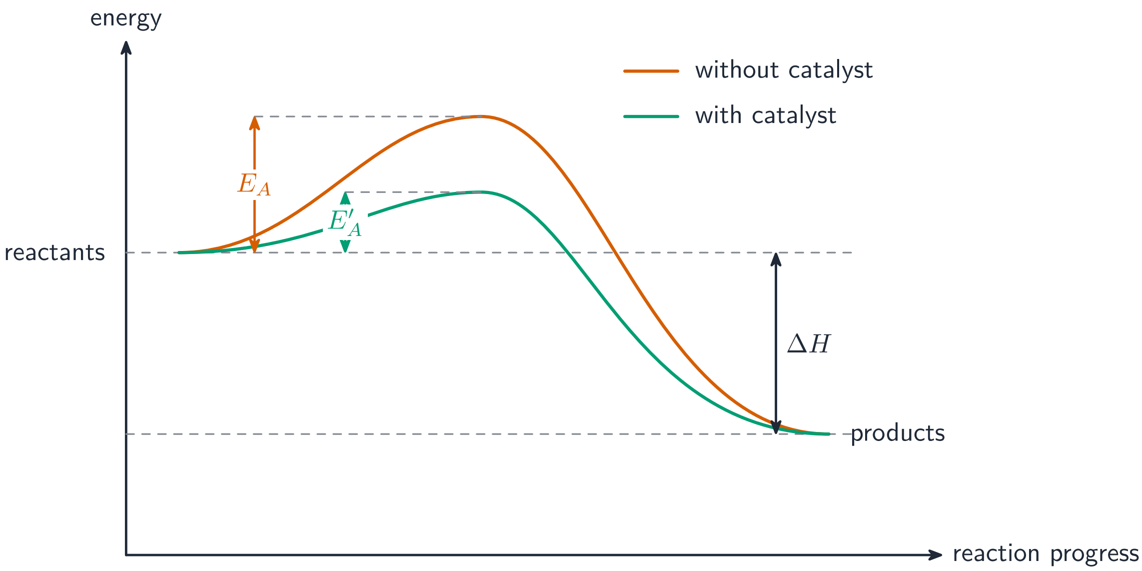
A catalytic converter speeds up the reactions that clean a car's exhaust gases.

Image: Unknown authorUnknown author, Public domain (commons.wikimedia.org)

A **catalyst** 催化剂 speeds up a reaction but is not used up itself. **Catalysis** 催化作用 is the name for this action.

A catalyst works by giving the reaction a different **reaction mechanism** 反应机理—a new route with a **lower** activation energy. On the Boltzmann distribution, lowering E_A moves the line to the left, so more molecules now have enough energy. This means more effective collisions each second, and a faster rate.

On a **reaction pathway diagram** 反应路径图, the catalysed route has a lower energy "hill". The enthalpy change of the reaction, ΔH, is **not** changed by the catalyst.



A catalyst gives a route with lower activation energy; the enthalpy change is unchanged

There are two types:

- a **homogeneous catalyst** 均相催化剂 is in the **same** physical state as the reactants —for example, an acid catalyst dissolved in a solution of liquids.
- a **heterogeneous catalyst** 多相催化剂 is in a **different** state from the reactants — for example, solid iron speeding up the reaction of gases in the Haber process.

A car's catalytic converter is a heterogeneous catalyst; its honeycomb gives a huge surface area

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