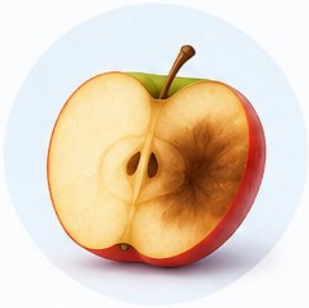
Reaction Rates:
What Makes Reactions Faster



- **Core question:** why do some reactions finish in milliseconds, others in years?
- **Reaction rate:** change in concentration per unit time
- **Four levers preview:** concentration/pressure, temperature, surface area, catalysts
- **Everyday anchors:** rusting iron, burning fuel, food spoilage, cold packs, catalytic converters
- **Roadmap:** every factor acts through collision frequency, collision energy, or the energy barrier

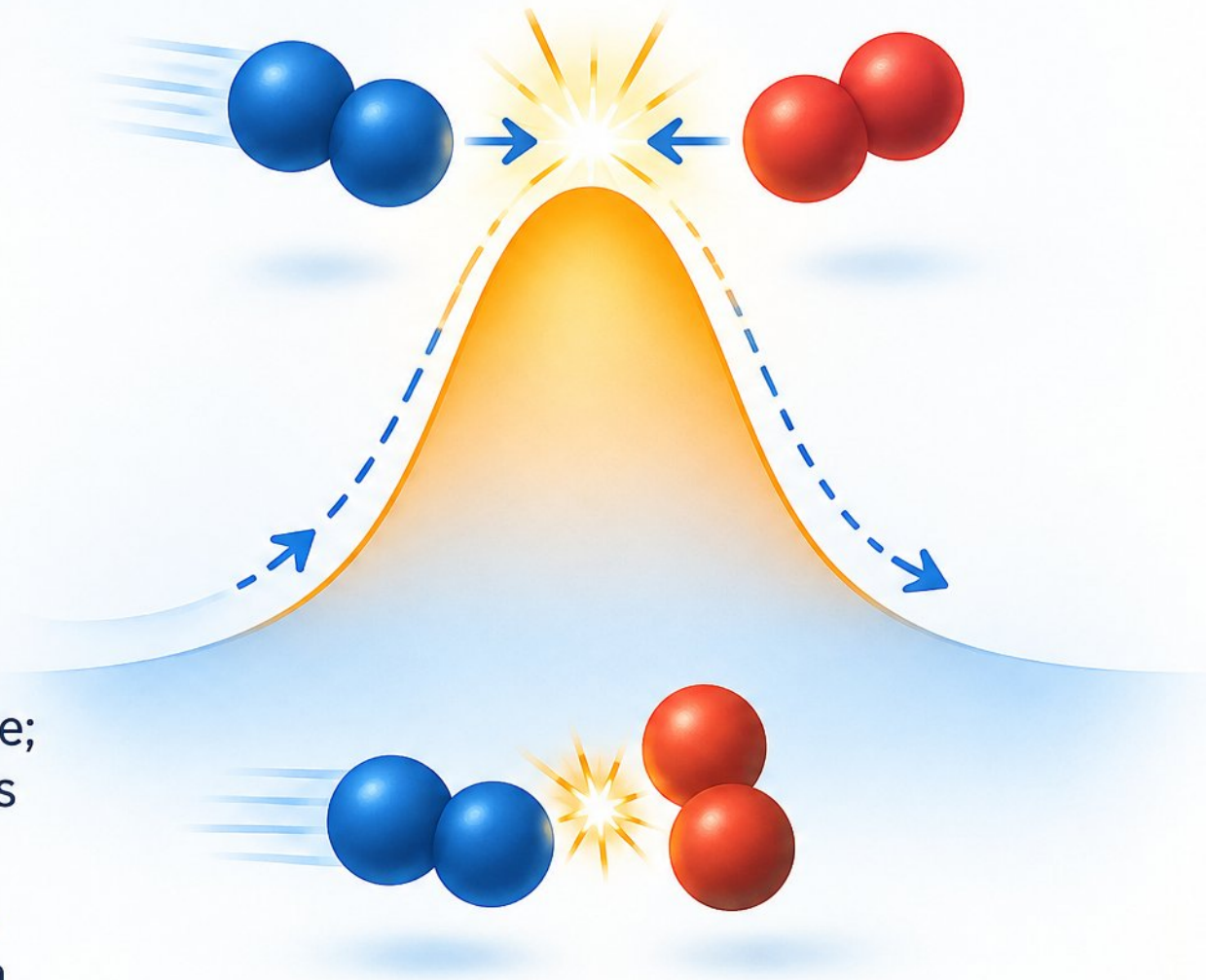






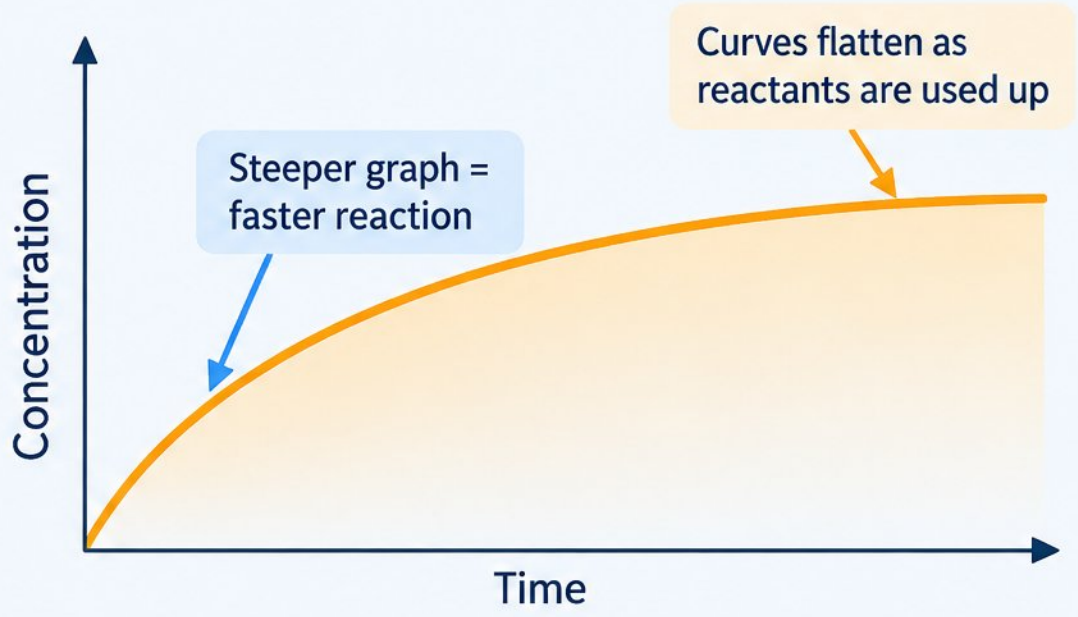
Collision Theory: The Molecular Basis of Rate

- Particles must collide for a reaction to occur; more collisions mean faster rate
- Only collisions with correct orientation and enough energy succeed
- Activation energy (Ea) and the transition state act as the gatekeeper
- Most collisions are ineffective; only a fraction yield products
- Each 'faster' factor changes collision frequency, collision energy, or the barrier



Reading the Rate: Measuring and Comparing Speeds

- Track gas volume, mass loss, color, pH, conductivity, or pressure
- Initial rate (start), average rate (interval), instantaneous rate (slope at a point)
- Steeper graph = faster reaction; curves flatten as reactants are used up
- Different species change at different rates — always state which substance you track
- Classroom methods: gas syringe, mass balance, colorimetry; each has limits





Gas syringe



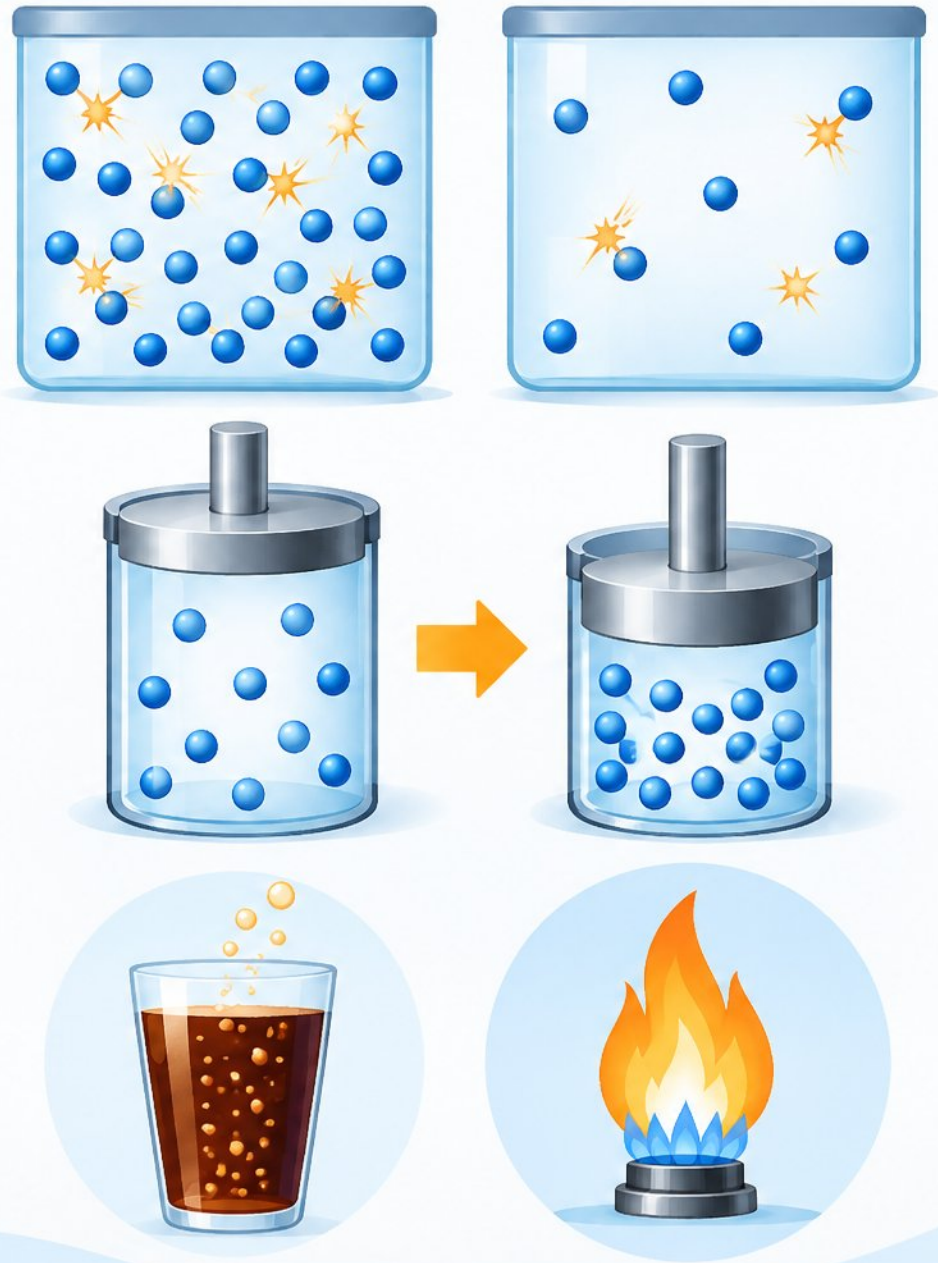
Mass balance



Colorimetry

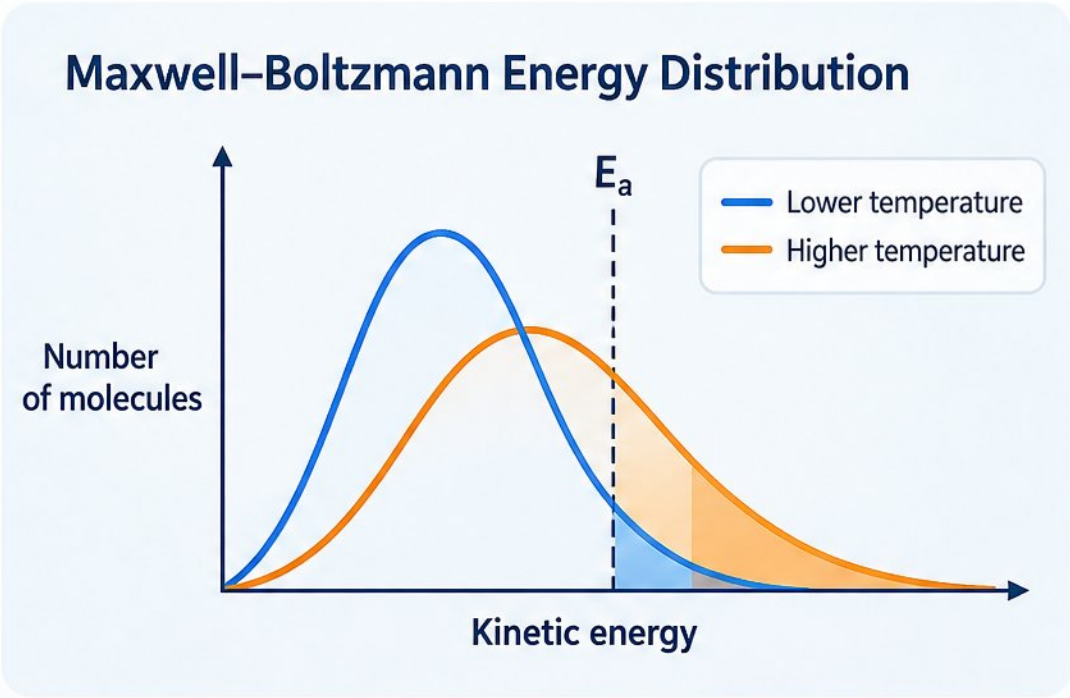
Concentration and Pressure: More Particles, More Collisions

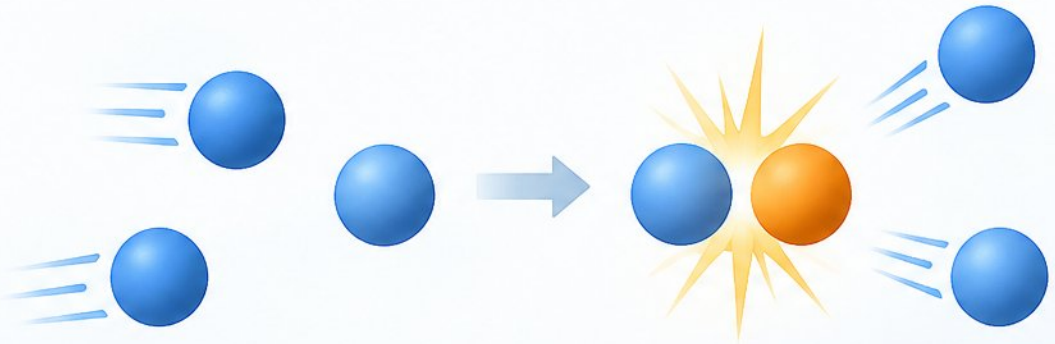
- More particles per unit volume → **more frequent collisions** → generally **faster reaction**
- In gases, **higher pressure** squeezes particles closer — it acts like **higher concentration**
- Reaction order shows how strongly rate responds; rate laws come from **experiment, not coefficients**
- **Not always**: some zero-order reactions keep a constant rate as concentration falls
- **Everyday**: fizzy drinks, and fuel burning in oxygen-rich versus oxygen-poor air



Temperature: More Energy per Collision

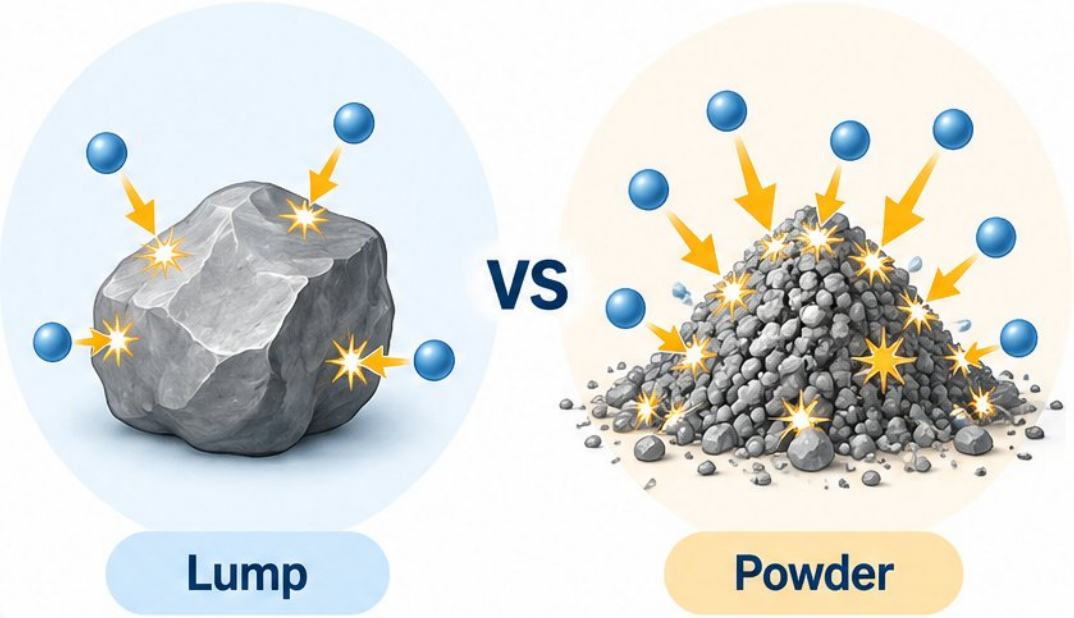
- Heat raises average kinetic energy and collision frequency
- Bigger effect: more collisions exceed the activation energy
- Maxwell–Boltzmann curve shifts right; shaded area beyond E_a grows
- Rule of thumb: rate roughly doubles per 10 °C rise
- Temperature changes the rate constant, not equilibrium position
- Refrigeration slows food spoilage by lowering k

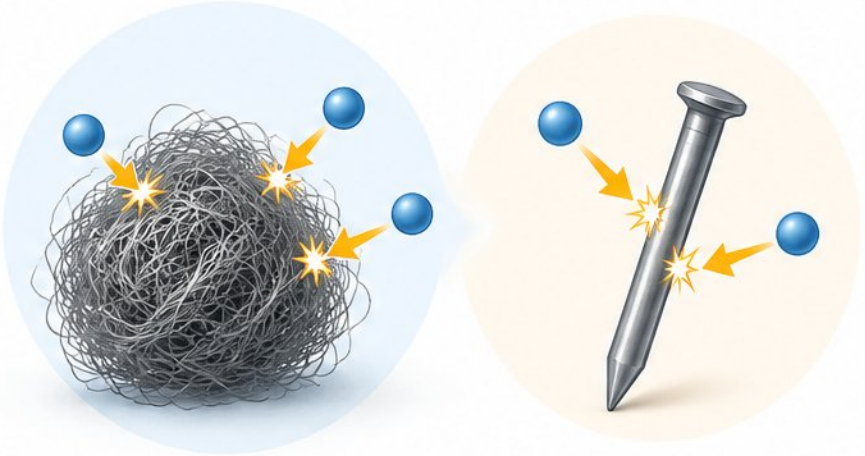


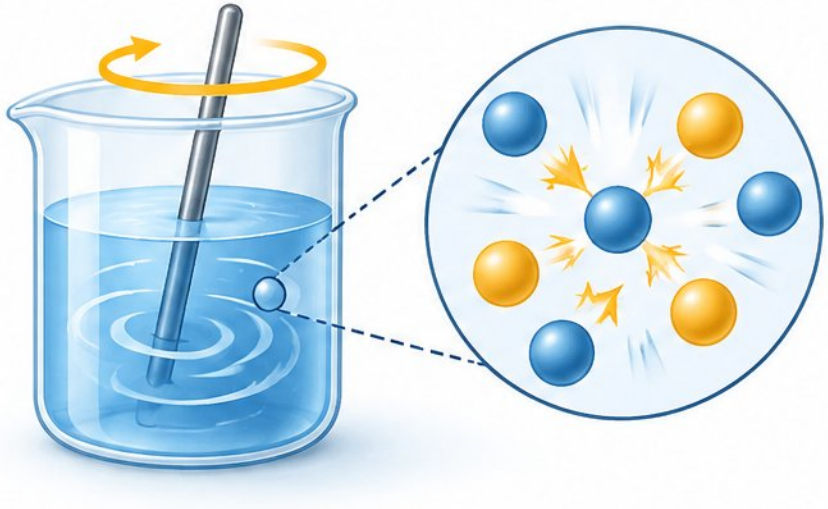


Surface Area and Mixing: Contact Matters

- Reactions occur at interfaces when reactants are in different phases
- More exposed surface area means more collision sites and faster reaction
- Classic demos: powder versus lump, steel wool versus nail
- Stirring and mixing reduce local depletion and improve contact
- Heterogeneous catalysts rely on surface area; metal powders pose dust-explosion risks

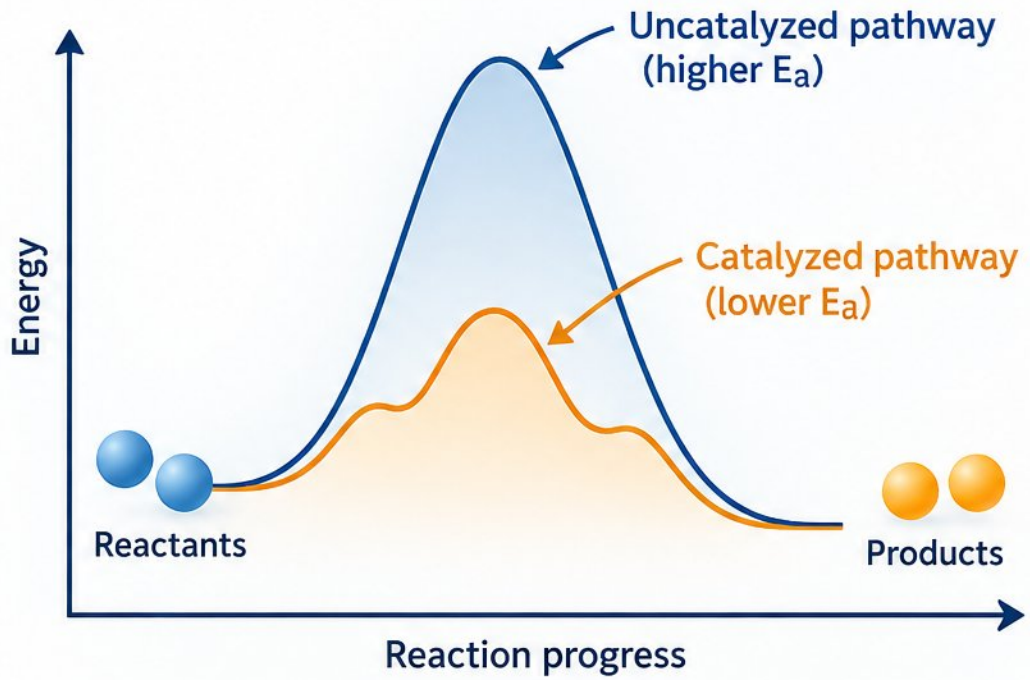






Catalysts: Lowering the Barrier

- Catalyst: alternative pathway with lower activation energy (E_a)
- Not consumed overall — reusable, but participates in the mechanism
- Changes rate, not equilibrium yield or Gibbs free energy difference
- Homogeneous (same phase) vs heterogeneous (different phase)
- Examples: catalytic converters, H_2O_2 decomposition, enzymes in digestion



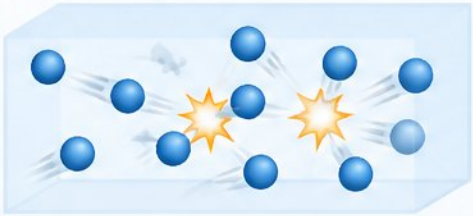


Connecting the Factors: Frequency, Energy, or Barrier

Collision Frequency

More collisions give more opportunities to react.

Concentration and pressure:
more particles per volume, more frequent collisions



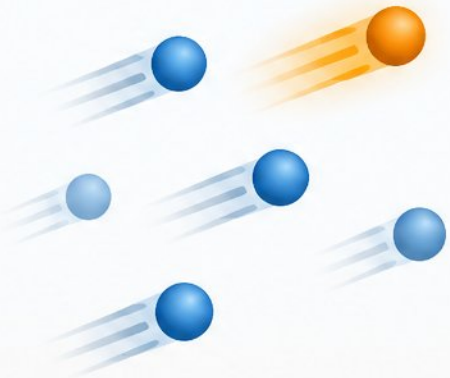
Surface area:
more exposed contact sites, more frequent collisions



Collision Energy

More energetic collisions help overcome the barrier.

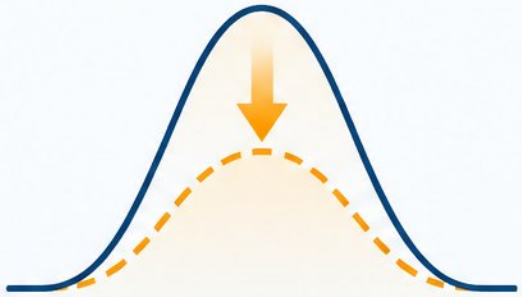
Temperature:
more energetic collisions; a larger fraction exceeds the activation energy



Barrier Height

A lower barrier makes reaction easier.

Catalysts:
an alternative pathway with a lower activation energy barrier



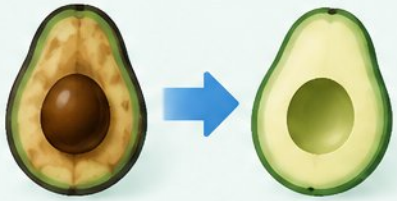
Match the lever to the reaction:
solid–gas needs surface area; gas-phase needs pressure

Putting It Together: Designing Faster Reactions

Reaction Examples: Using Rate Levers



Speeding up rust removal: add acid, warm it, use steel wool, stir

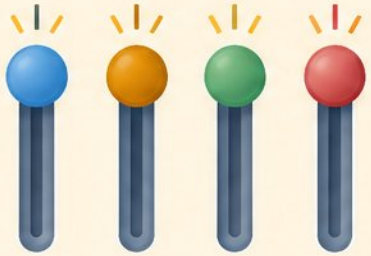


Slowing food oxidation: refrigerate, limit surface exposure, add antioxidants

Choose a reaction



Select the best rate lever



Check trade-offs

- safety
- cost
- selectivity
- side reactions

Predict-then-test

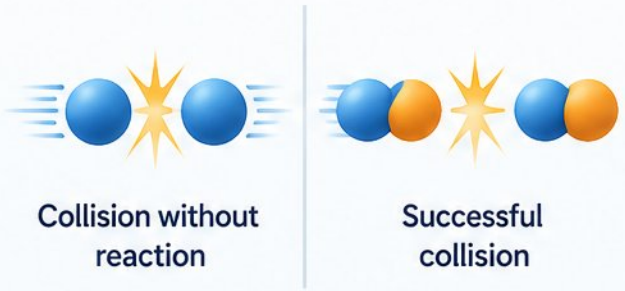
change one factor, predict the effect, test it



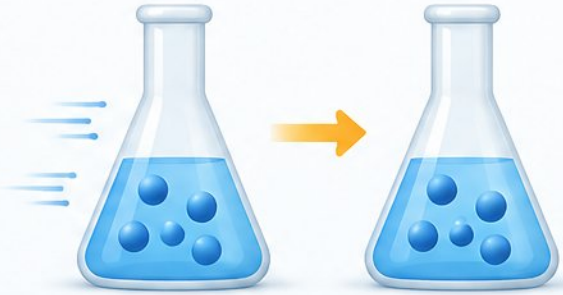
Self-check: Which lever acts on collision frequency versus energy?

Common Misconceptions and How to Surface Them

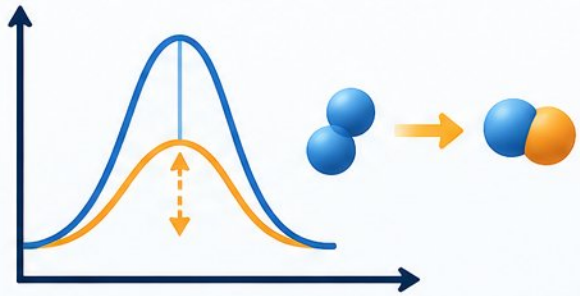
1. "Any collision reacts" — not all collisions are successful; repeat this often



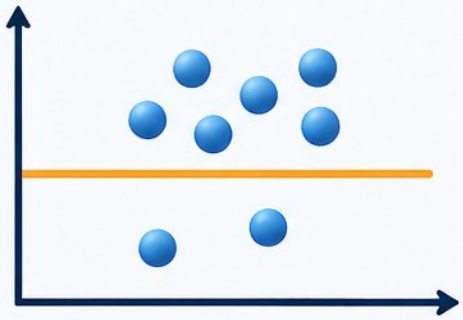
2. Faster reaction ≠ more product: rate changes speed, not final yield



3. Catalysts lower activation energy; they do not add energy or raise yield



4. "Rate allways rises with concentration" — zero-order reactions stay constant



5. Rate laws come from experiment, not stoichiometry; show varied examples



6. Use questioning and conceptual-change activities to surface these ideas



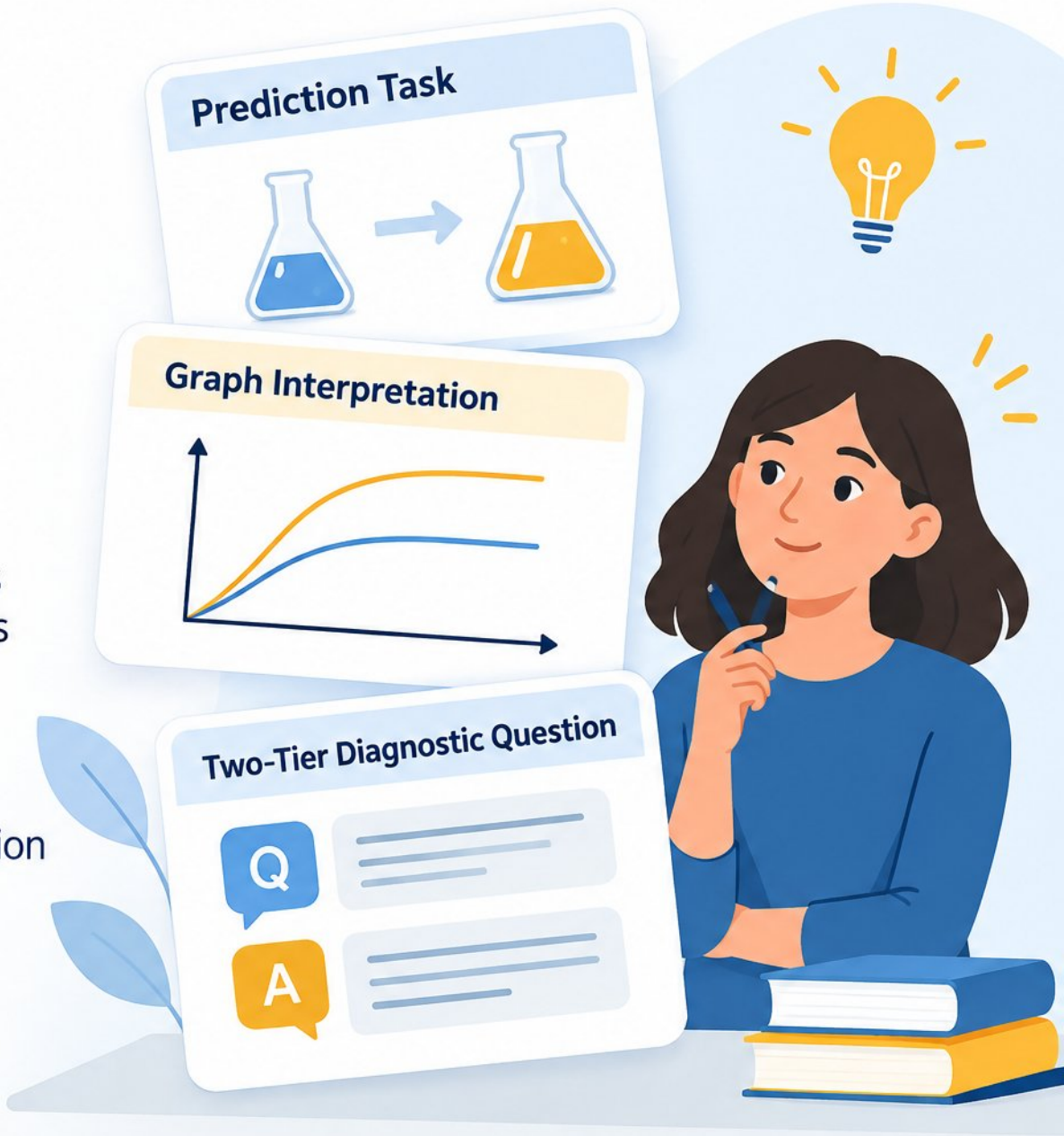
Teaching with Demonstrations and Inquiry

- Low-cost, safe investigations: gas syringe, mass loss, color change
- Choose concentration ranges so all reactions finish on time
- 5E inquiry plus analogies boosted conceptual understanding in studies
- Cooperative and conceptual-change teaching beat traditional instruction
- Differentiate: school students, adult learners, and teachers need different supports



Assessment and Feedback: Checking Understanding

- Use graph interpretation, prediction tasks, and experimental design
- Ask learners to interpret and explain reaction curves, not recall
- Two-tier diagnostic questions reveal alternative conceptions
- Target known trouble spots: reaction order and catalyst action
- Use results to reteach, not just to grade



Synthesis: The Four Levers of Reaction Rate

Concentration / pressure



Raises collision frequency

Surface area



Raises collision frequency

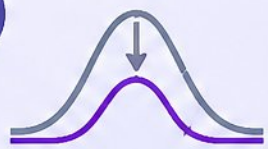


Temperature



Boosts collision energy

Catalysts



Lowers the activation barrier via an alternative pathway

- Four levers: concentration/pressure, temperature, surface area, catalysts — each speeds reactions
- Concentration/pressure and surface area raise collision frequency; temperature boosts collision energy
- Catalysts lower the activation barrier via an alternative pathway
- Key formulas: rate = Δ[concentration]/Δt; $k = A \cdot e^{(-Ea/RT)}$; rate = $k[A]^m[B]^l$

Everyday anchors: rusting, food spoilage, cold packs, catalytic converters, enzymes



rusting



food spoilage



cold packs



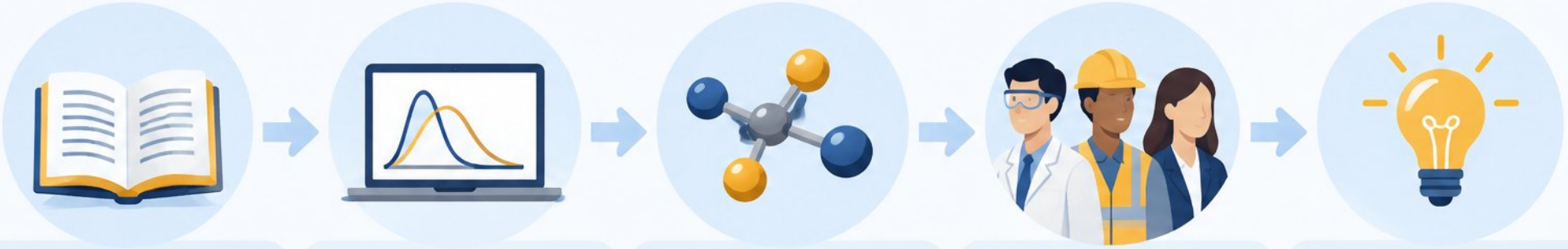
catalytic converters



enzymes

Learners: predict, then test. **Teachers:** revisit misconceptions about order and catalysts.

Further Resources and Next Steps



- CPD articles on rates, collision theory, and catalysis from RSC Education
- Interactive simulations for Maxwell–Boltzmann distributions and collision models
- Next topics: reaction mechanisms, rate laws, and industrial kinetics
- Use up-to-date contexts and careers to capture learner interest
- Confirm why some reactions finish in milliseconds while others take years



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