16–18 years

Collision theory and Maxwell–Boltzmann distribution curves

ROYAL SOCIETY OF CHEMISTRY

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Learning objectives

By the end of this lesson, you will be able to:

- Understand reaction kinetics in terms of collision theory and energy profile diagrams.
- Draw and interpret Maxwell–Boltzmann distribution curves.
- Use Maxwell–Boltzmann distribution to explain how a change in temperature affects the rate of reaction.
- Use Maxwell–Boltzmann distribution to help explain the action of a catalyst on reaction rate.

Learning objectives

Collison theory of chemical reactions

The rate of a chemical reaction is a measure of how fast a reaction takes place. Chemical reactions take place when reactant particles successfully collide.

For particles to react they must collide:

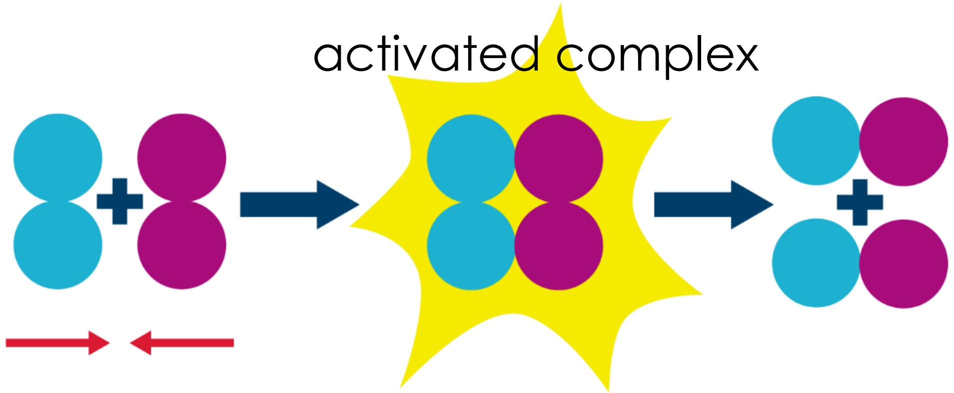
- with sufficient kinetic energy
- at the correct orientation.

The kinetic energy is usually needed to break the relevant chemical bonds in the reactants. The **activation energy** is the minimum amount of energy required for particles to react.

The **rate of reaction** is a measure of the number of successful collisions per unit time.

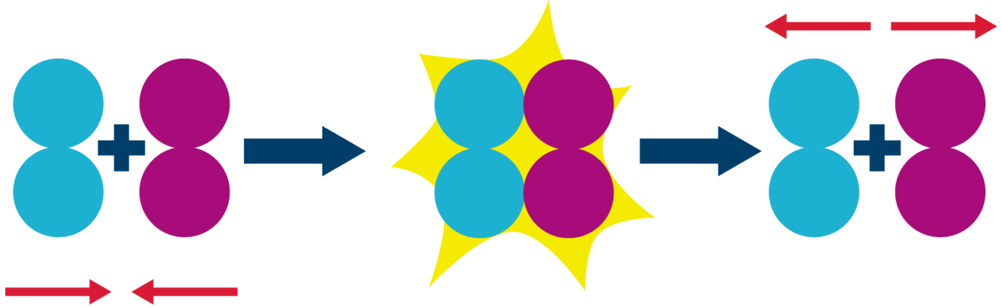
Collison theory: particle energy

A successful collision:



- Reactant particles collide with sufficient kinetic energy.
- Bonds break in the reactant molecules and new bonds form in product molecules.
- The collision is successful.
- Products are formed (provided the particle orientation is correct).

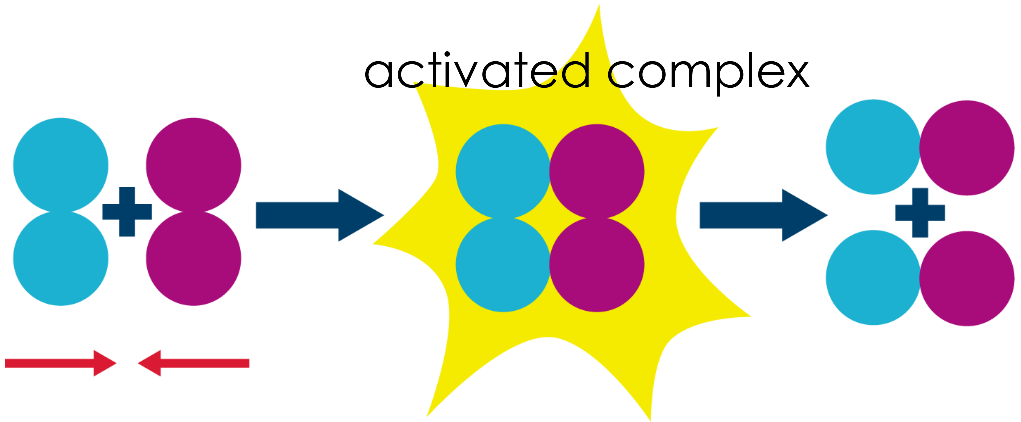
An unsuccessful collision:



- Reactant particles collide with insufficient kinetic energy.
- Bonds do not break in the reactant molecules.
- The collision is unsuccessful.
- The reactant particles move away from each other.
- Products are not formed.

Collision theory: particle orientation

A successful collision:



- Reactants particles collide with the correct orientation.
- Products are formed (provided they collide with sufficient kinetic energy).

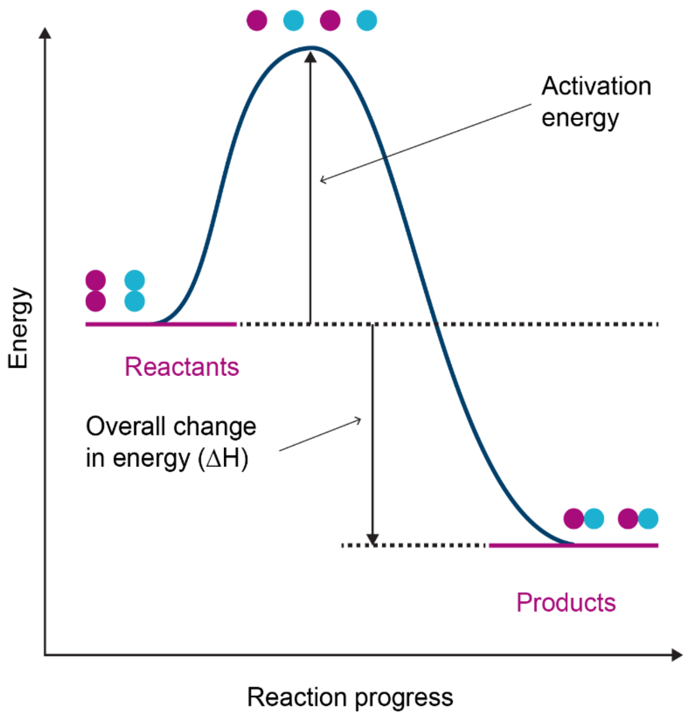
An unsuccessful collision:

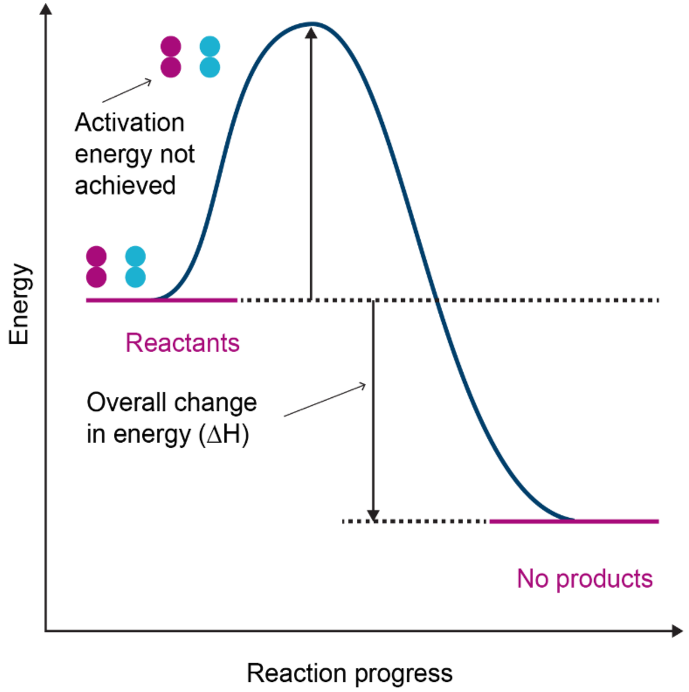
- Reactants particles collide with the incorrect orientation.
- The reactants move away from each other.
- Products are not formed.

Energy profile diagrams

An energy profile diagram shows the energy levels of reactants, intermediates and products as the reaction progresses.
These diagrams show the energy profile of a successful reaction (left) and an unsuccessful reaction (right).

Explain why one reaction is successful and the other isn't.



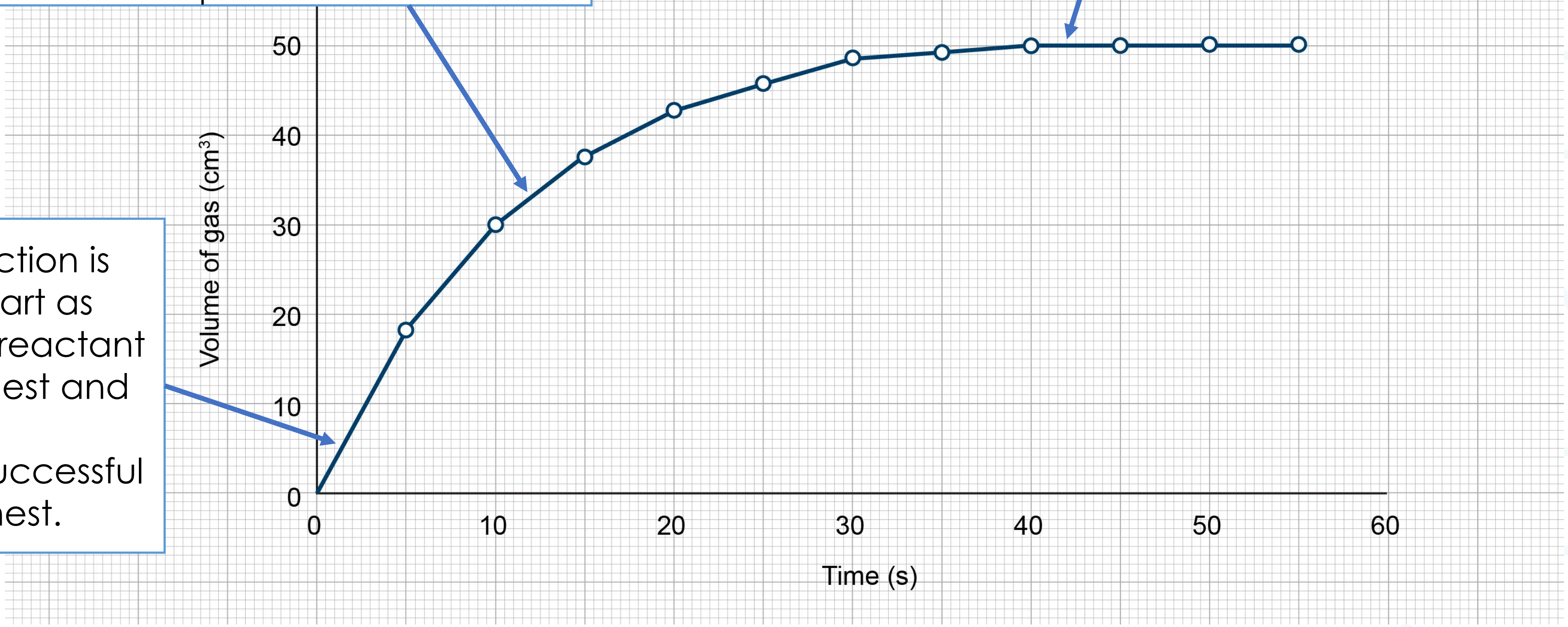


Interpreting experimental data

This graph shows typical data collected when investigating the reaction of magnesium and acid.

The rate of reaction decreases as the reaction progresses, as many of the reactant particles have been used up so fewer successful collisions are possible.

The limiting reagent has been used up – therefore no more successful collisions are possible.



The rate of reaction is fastest at the start as the number of reactant particles is highest and therefore the frequency of successful collisions is highest.

Introducing the Maxwell–Boltzmann distribution

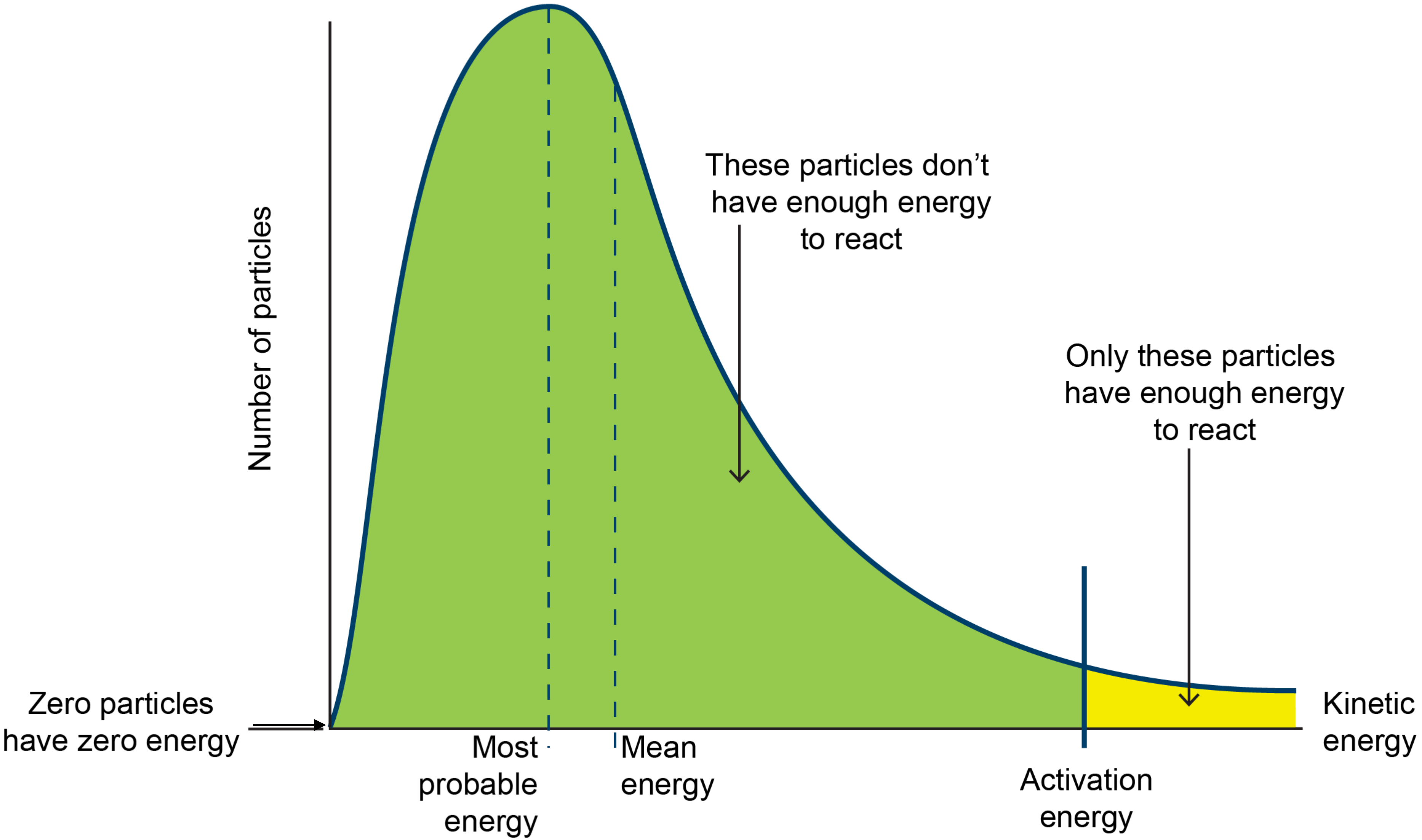
Not all particles in a substance have the same amount of energy. Their energies are distributed in a pattern called the Maxwell–Boltzmann distribution. The area under the curve represents the total number of particles.

This online simulation shows a dynamic animation of particle movement. The difference in kinetic energy of the particles is shown by colour coding.

The Maxwell–Boltzmann distribution is plotted as a histogram, below the particles.

Changing the reaction conditions affects the particle energy distribution. The number of particles remains the same.

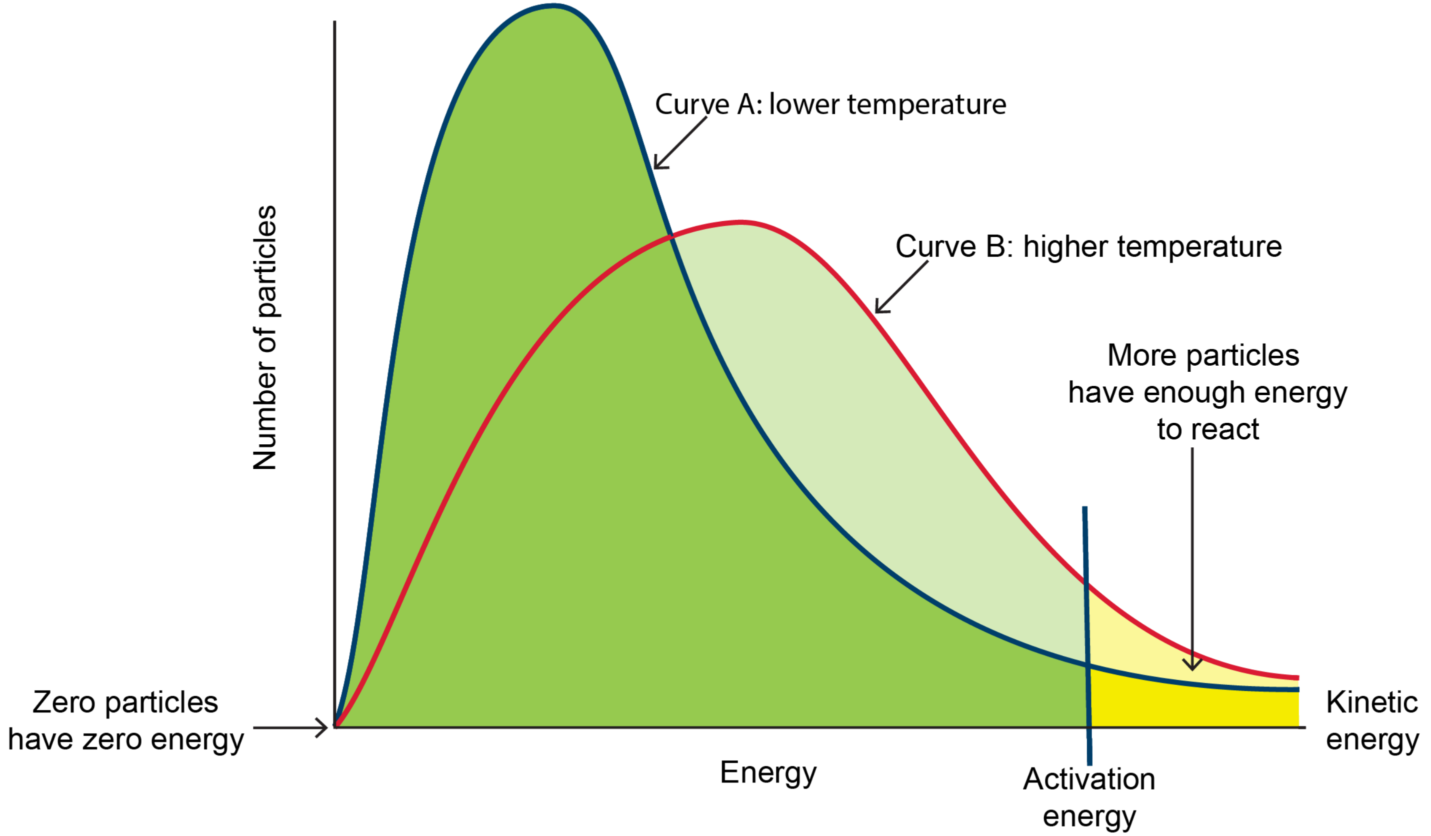
Maxwell–Boltzmann distribution



Maxwell–Boltzmann distribution and temperature

Increasing the temperature shifts the energy distribution to the **right**, i.e. towards having more particles with higher energies.

Collisions will occur more often and more will be successful because more particles have energy greater than, or equal to, the activation energy.



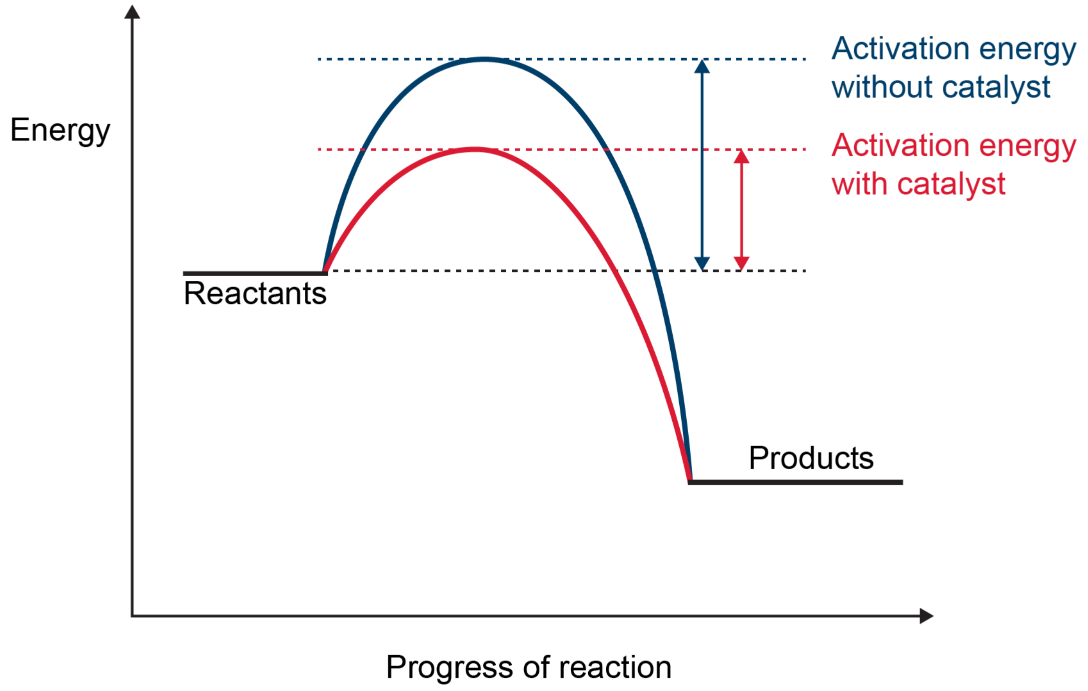
Therefore, the rate of reaction will increase as there are more frequent successful collisions.

Energy profile diagram: catalysts

A **catalyst** is a substance that speeds up a chemical reaction without taking part.

A catalyst works by providing an alternative reaction pathway with lower activation energy.

There are many economic and environmental benefits of using catalysts, as reactions can take place at lower temperatures and pressures.



Maxwell–Boltzmann distribution and catalysts

When a catalyst is added, the shape of the Maxwell–Boltzmann distribution stays the same.

The position of the activation energy is shifted to the **left**.

A greater proportion of particles have sufficient energy to react.

Therefore, the rate of reaction will increase as there are more frequent successful collisions.