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Topic 6 Kinetics SL

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Rate of reaction

Rate of reaction

The rate of reaction is the change in concentration of reactants or products per unit time.

$$\text{rate of reaction} = \frac{\text{increase in product concentration}}{\text{change in time}}$$

$$\text{rate of reaction} = \frac{\text{decrease in reactant concentration}}{\text{change in time}}$$

$$\text{rate} = -\frac{\Delta[\text{R}]}{\Delta t}$$

$$\text{rate} = \frac{\Delta[\text{P}]}{\Delta t}$$

Rate of reaction

The rate of reaction is the change in concentration of reactants or products per unit time.

$$\text{rate of reaction} = \frac{\text{increase in product concentration}}{\text{change in time}}$$

$$\text{rate of reaction} = \frac{\text{decrease in reactant concentration}}{\text{change in time}}$$

The unit for rate of reaction is mol dm^{-3} s^{-1}

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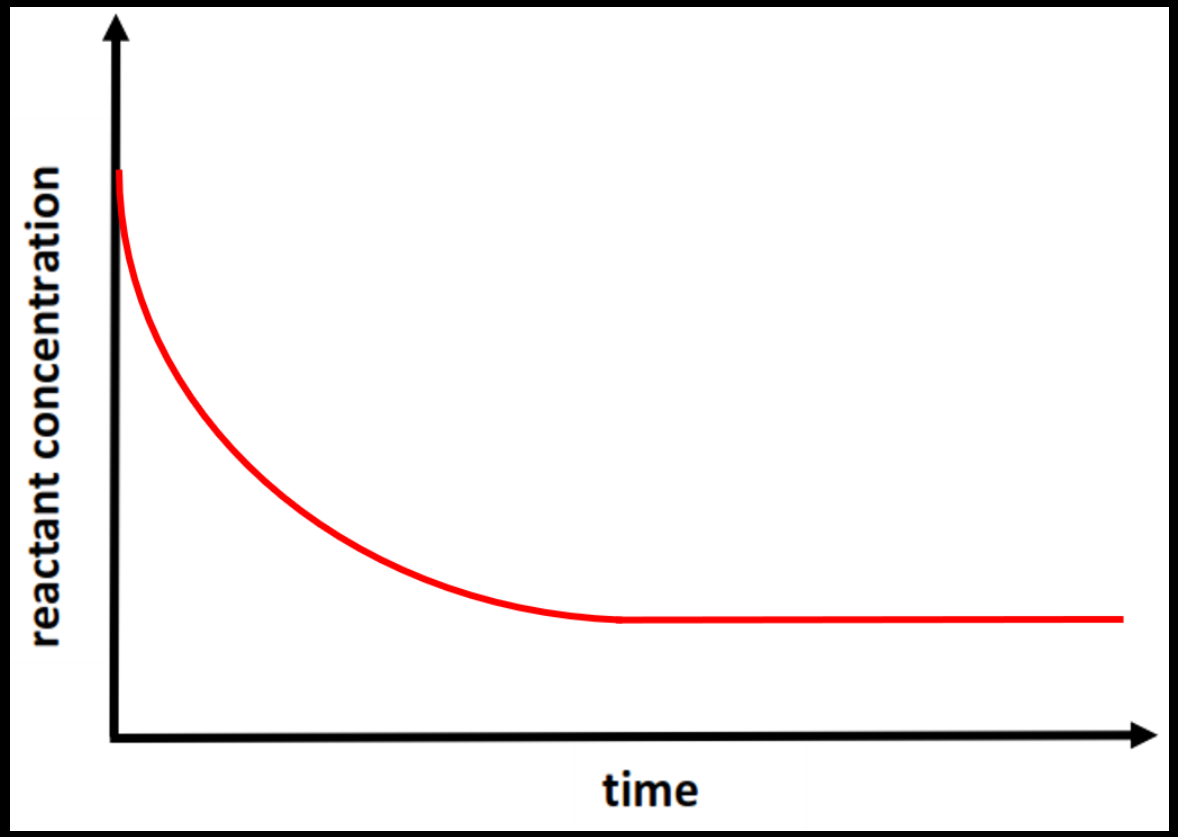
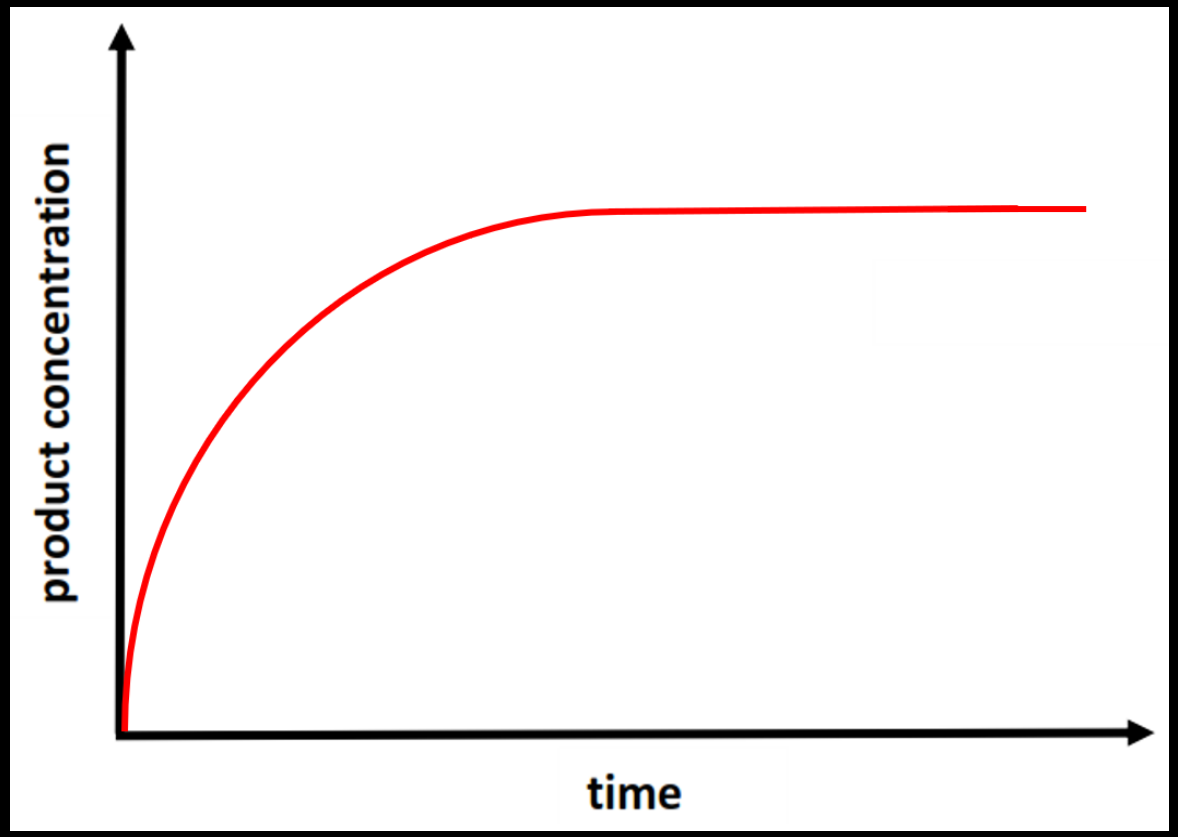
Rate of reaction

Change in reactant concentration over time

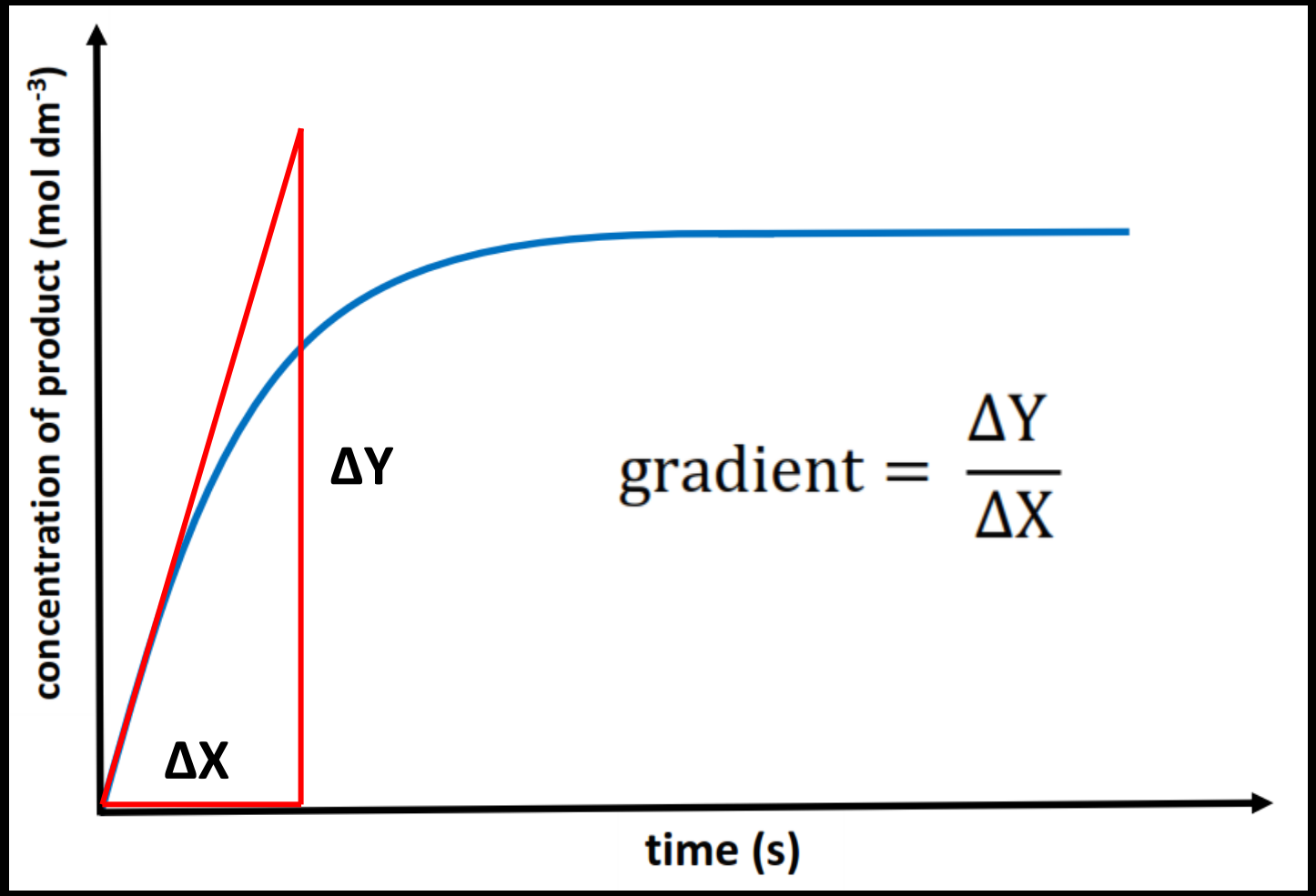


Change in product concentration over time



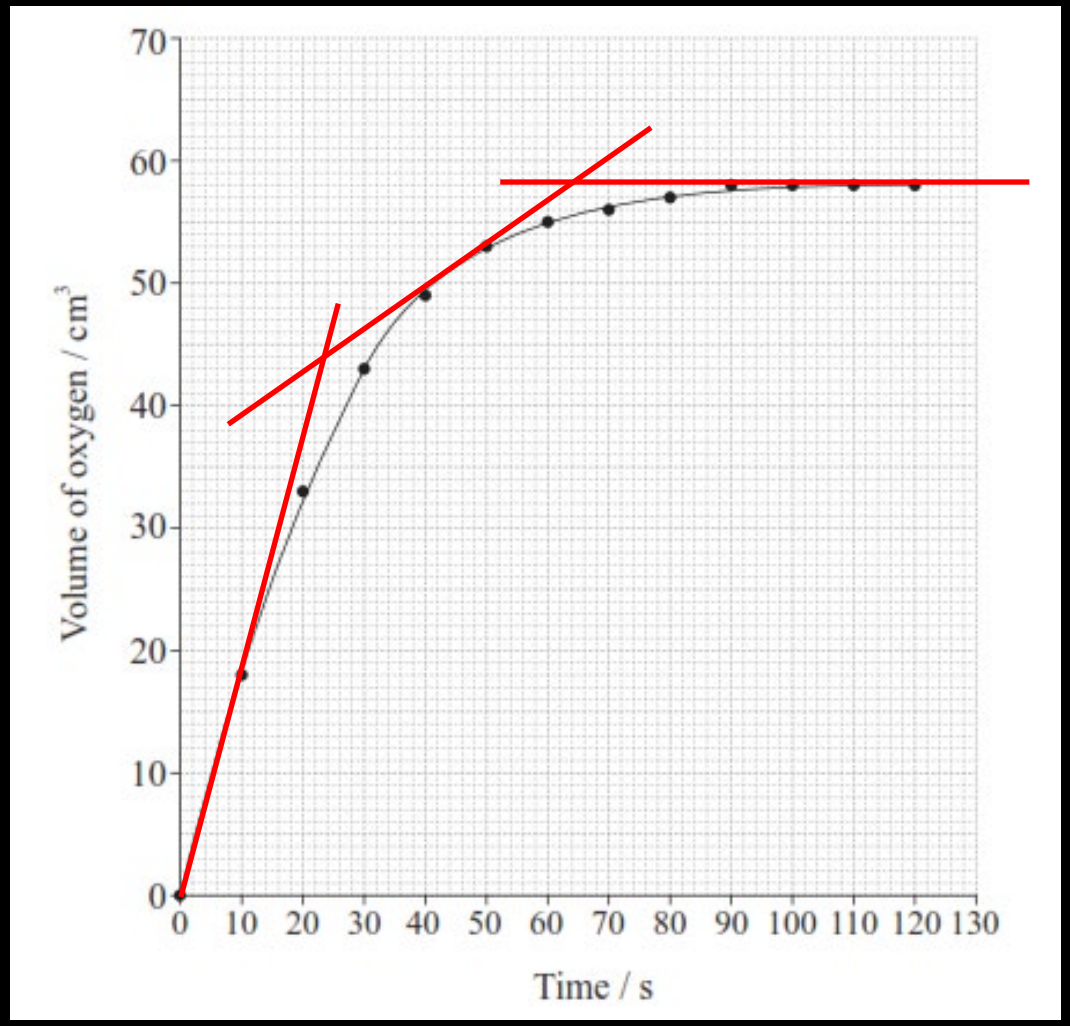
Rate of reaction

The instantaneous rate of reaction at time *t* can be calculated from a graph of concentration against time.



Draw a tangent to the curve at a specific time (at time = 0 s for the initial rate of reaction).
Measure the gradient of the tangent.
The gradient of the tangent = the rate of reaction.

Rate of reaction



Initial rate of reaction is fastest – high concentration of reactant particles results in a high frequency of collisions between reactant particles.

The rate of reaction decreases with time as the concentration of reactant particles decreases which results in a decrease in the frequency of collisions between reactant particles.

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Analysing rate of reaction graphs

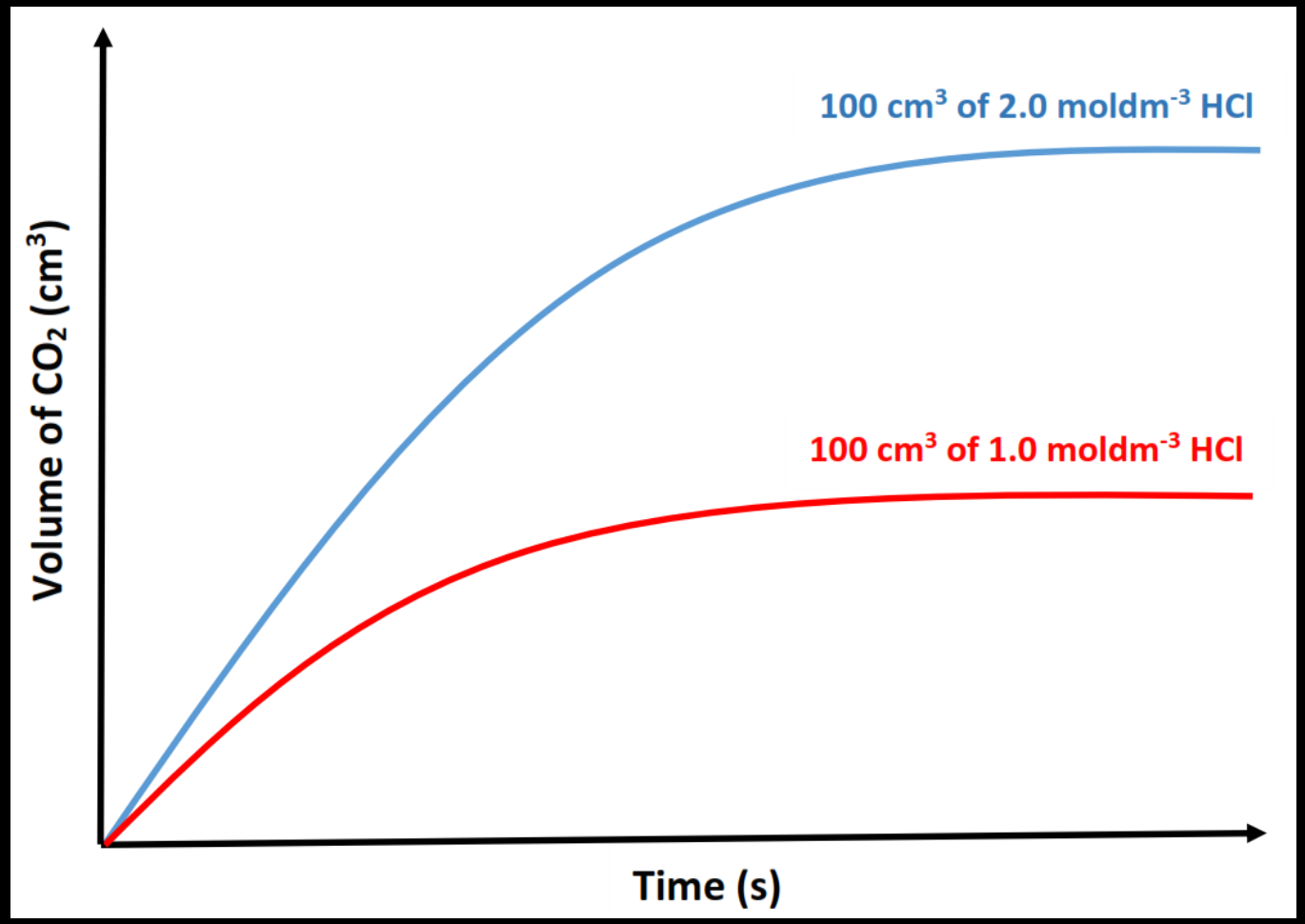
Rate of reaction graphs

In the following graphs of volume of CO_2 produced against time, the HCl is the limiting reactant, therefore it determines the amount (in mol) of CO_2 produced.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

The molar ratio of HCl to CO_2 is 2:1
Assume that all reactions are carried out at STP (100 kPa and 273 K).

Rate of reaction graphs



Blue curve

100 cm^3 of 2.0 mol dm^{-3} HCl

n = CV

2.0 × 100/1000 = 0.20 mol HCl

= 0.10 mol CO_2

Red curve

100 cm^3 of 1.0 mol dm^{-3} HCl

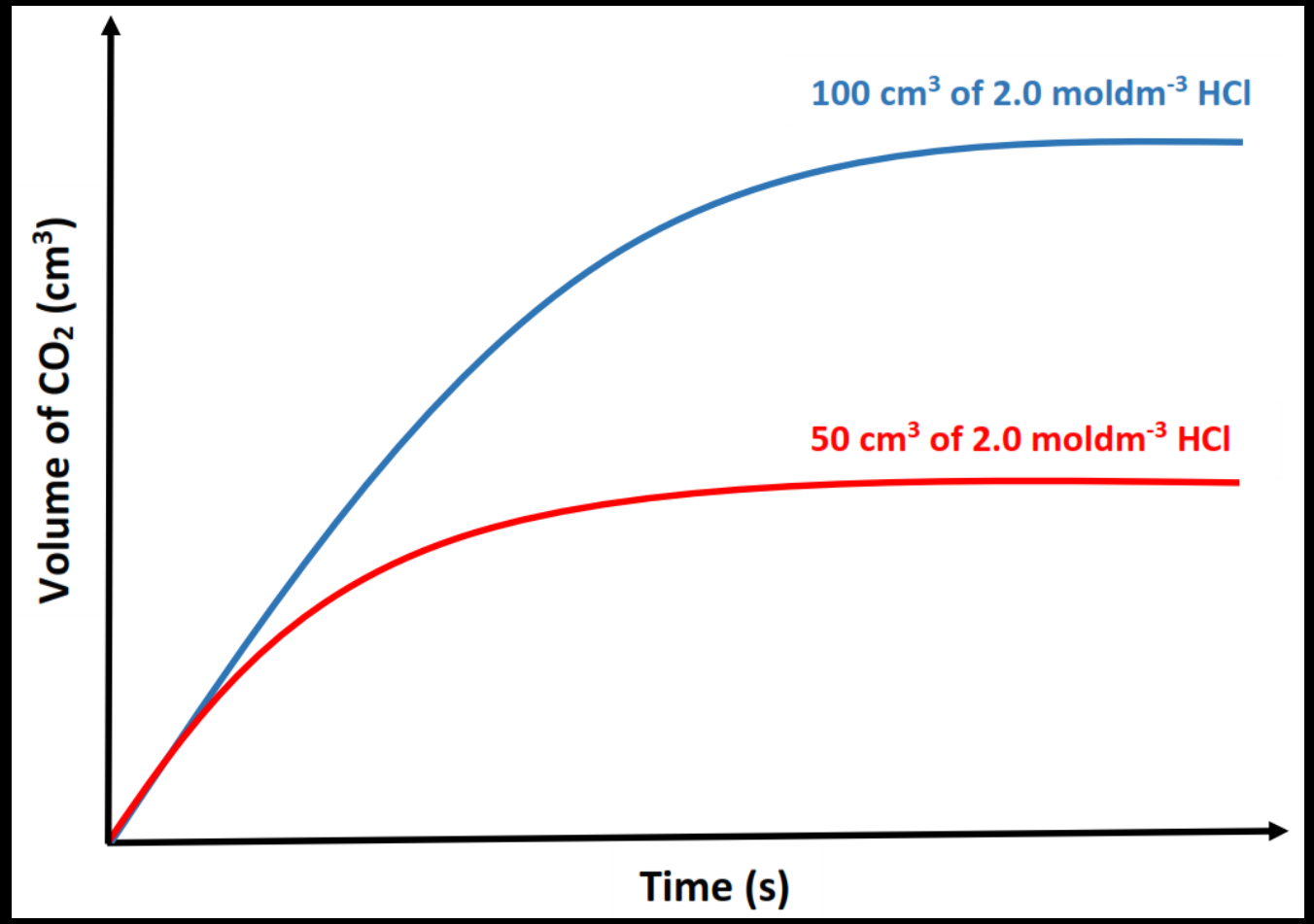
n = CV

1.0 × 100/1000 = 0.10 mol HCl

= 0.05 mol CO_2

Initial rate of reaction is higher and double the volume of gas is produced.

Rate of reaction graphs



Blue curve

100 cm_3 of 2.0 mol dm^{-3} HCl

n = CV

2.0 × 100/1000 = 0.20 mol HCl

= 0.10 mol CO_2

Red curve

50 cm_3 of 2.0 mol dm^{-3} HCl

n = CV

2.0 × 50/1000 = 0.10 mol HCl

= 0.05 mol CO_2

Initial rate of reaction is the same but half the volume of gas is produced.

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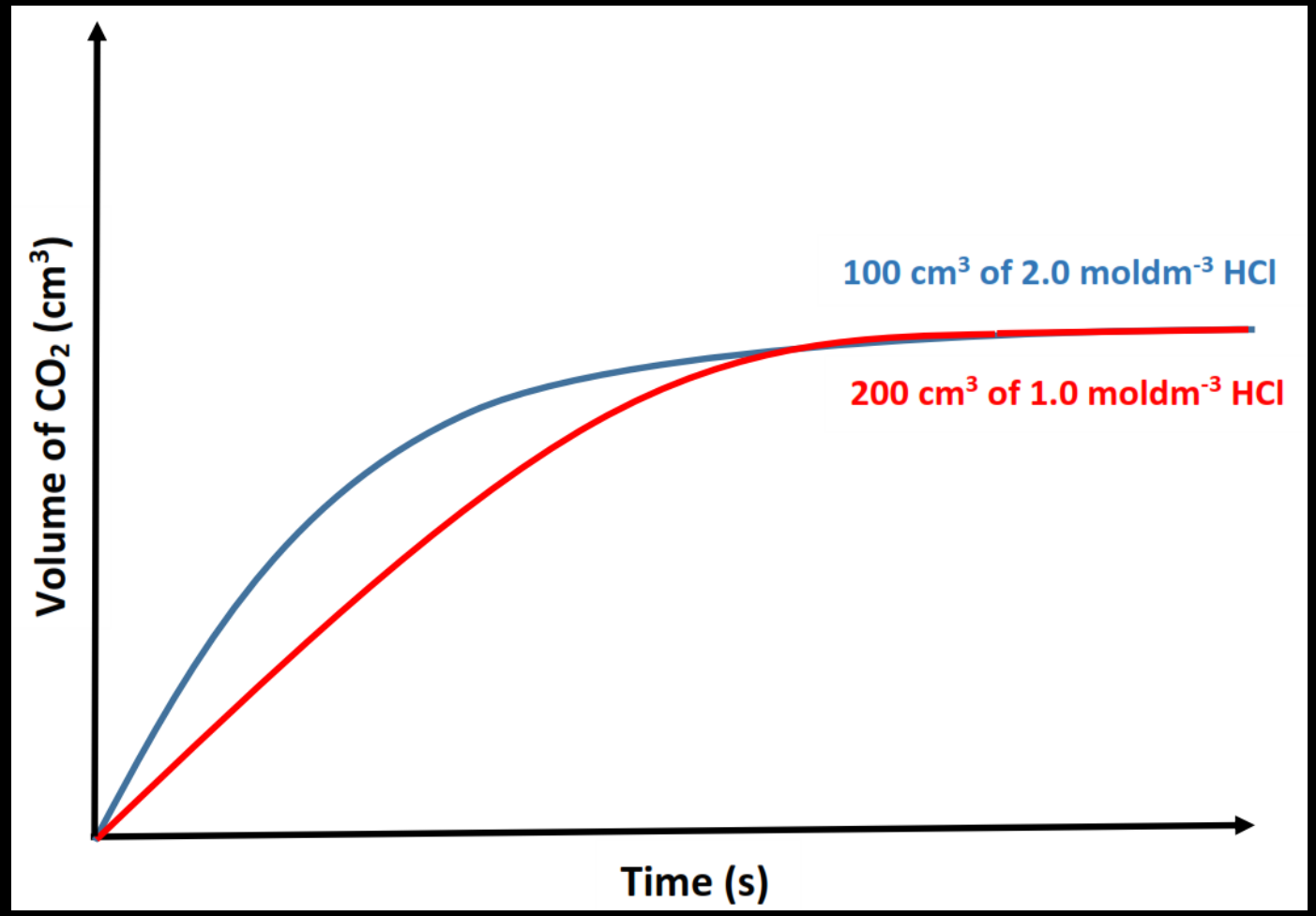
Rate of reaction graphs



Blue curve

100 cm_3 of 2.0 $moldm^{-3}$ HCl

n = CV

2.0 × 100/1000 = 0.20 mol HCl

= 0.10 mol CO_2

Red curve

200 cm_3 of 1.0 $moldm^{-3}$ HCl

n = CV

1.0 × 200/1000 = 0.20 mol HCl

= 0.10 mol CO_2

Initial rate of reaction is higher but the same volume of gas is produced.

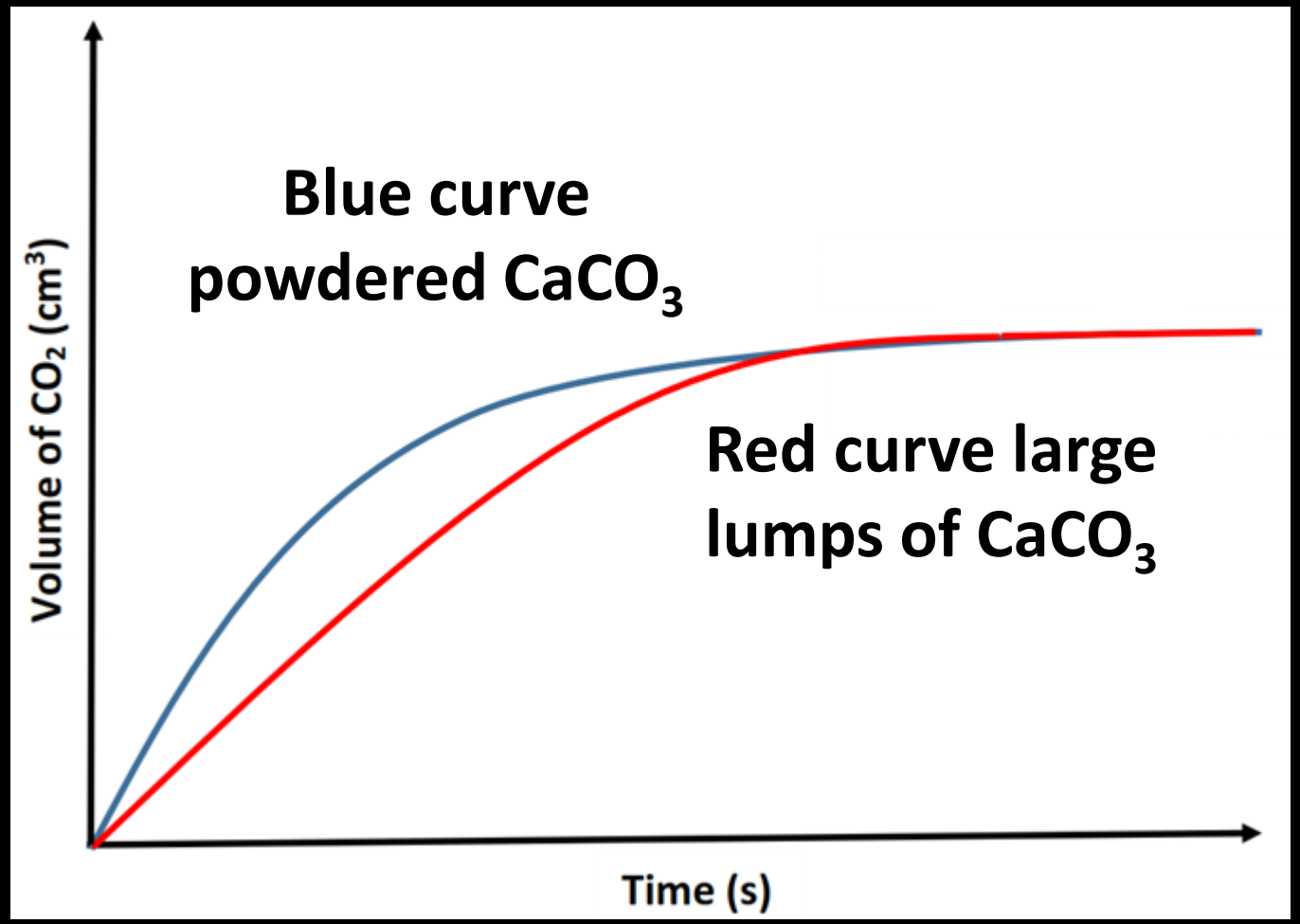
Rate of reaction graphs

The initial rate of reaction depends on the concentration of the HCl – the higher the concentration, the faster the initial rate and the steeper the gradient of the curve.

The volume of gas produced depends on the amount (in mol) of the limiting reactant.

Rate of reaction graphs

The same mass of powdered and large lumps of $CaCO_3$ were added to separate samples of excess hydrochloric acid.



Powdered solids react faster than large lumps due to the greater surface area.

As surface area increases, the frequency of collisions increases and the rate of reaction also increases.

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Collision theory

Collision theory

Reactant particles must collide with each other to have a chemical reaction.

Colliding particles must satisfy two conditions:

- They must collide with energy equal to, or greater than, the activation energy for the reaction.
- They must collide with the correct orientation (geometry).

Collisions between reactant particles that result in a chemical reaction are called successful, or effective collisions.

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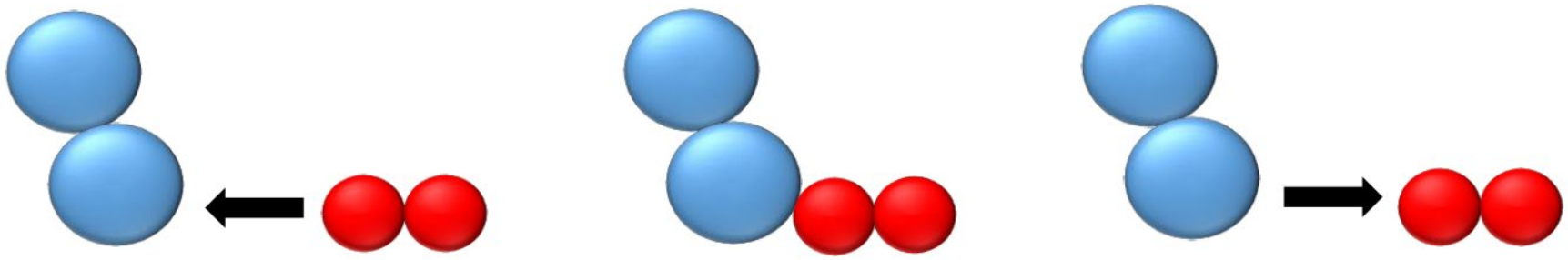
Collision theory

Unsuccessful collision

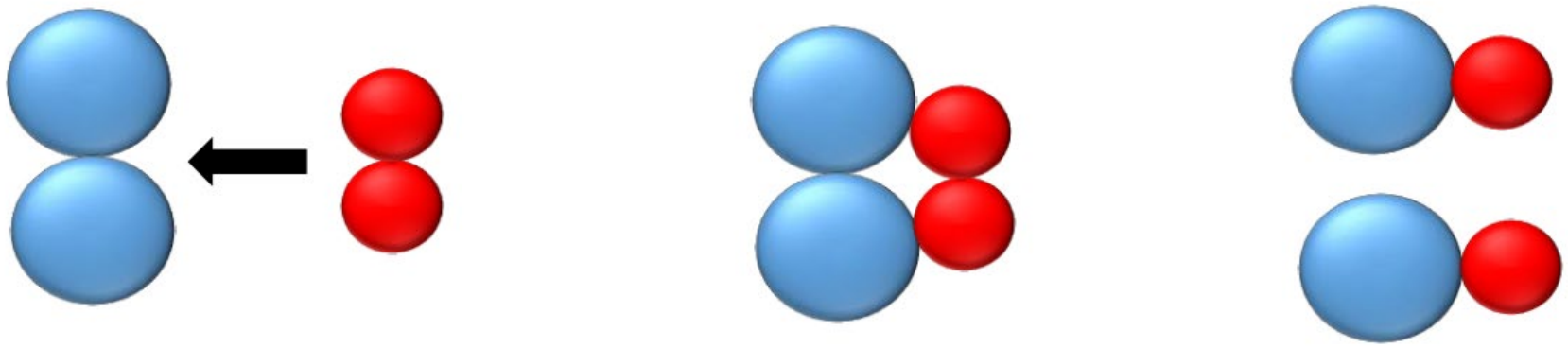


Reactant particles approach each other

They collide with the incorrect orientation

Reactant particles separate

Successful collision



Reactant particles approach each other

They collide with the correct orientation

Products formed

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Activation energy (E_a)

Collision theory

Reactant particles must collide with each other to have a chemical reaction.

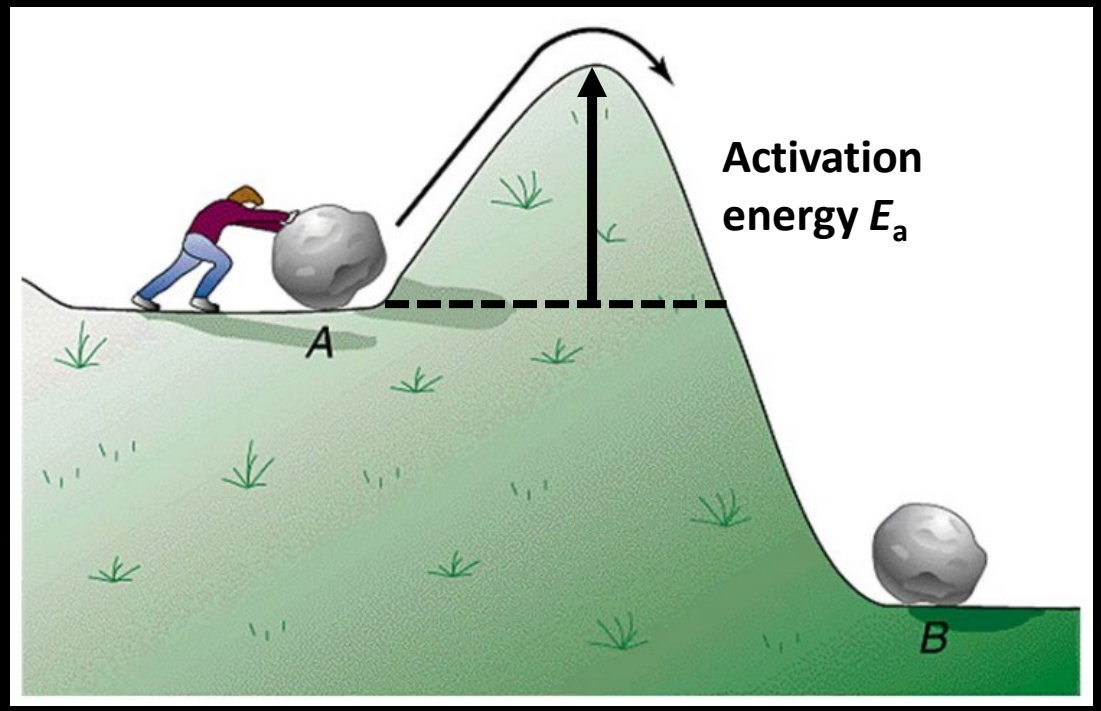
Colliding particles must satisfy two conditions:

- They must collide with energy equal to, or greater than, the activation energy for the reaction.
- They must collide with the correct orientation (geometry).

Activation energy

Activation energy (E_a) is the minimum energy that colliding particles need in order to have a successful collision that results in a chemical reaction.

It can be thought of as the energy barrier to a reaction.



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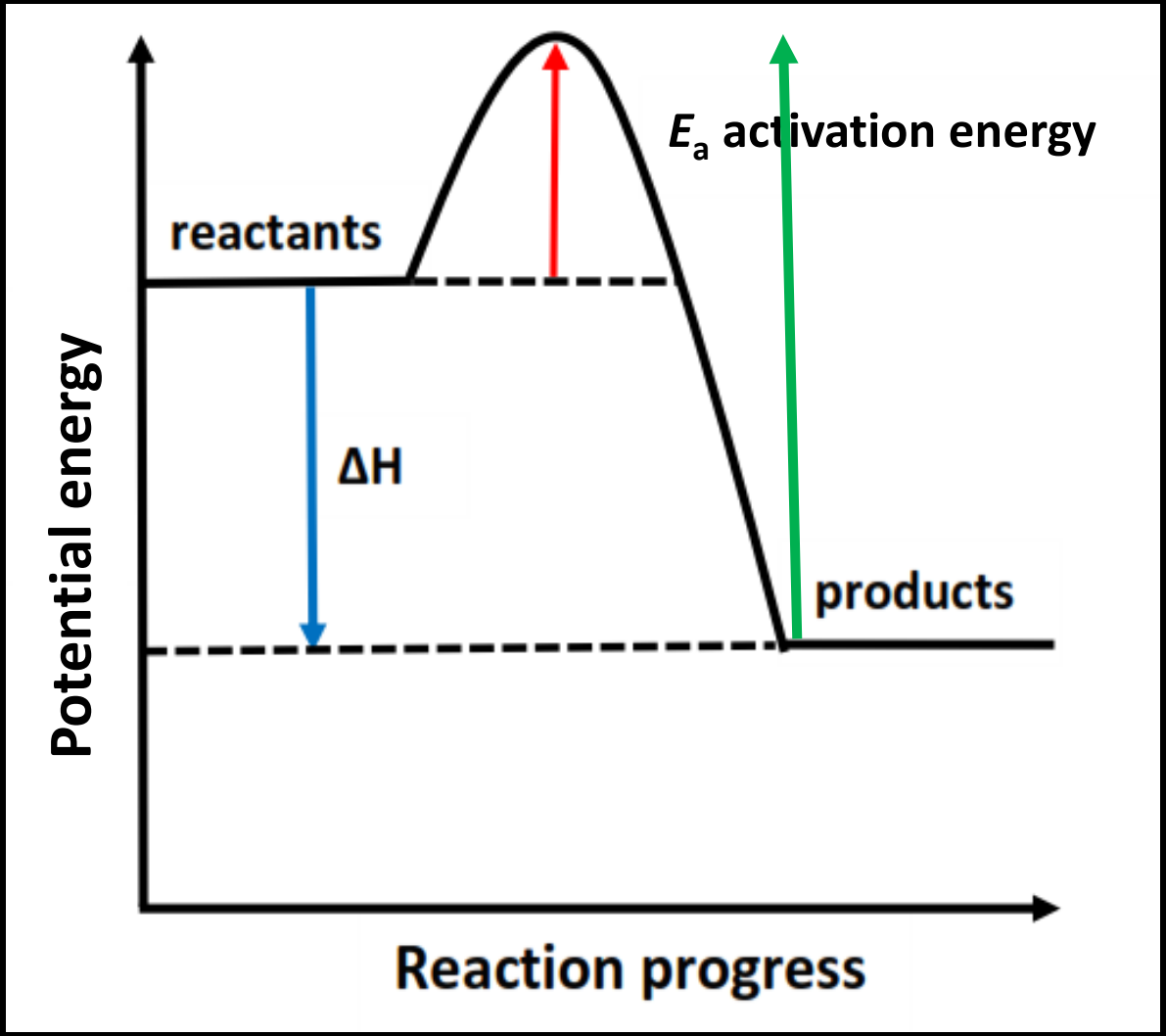
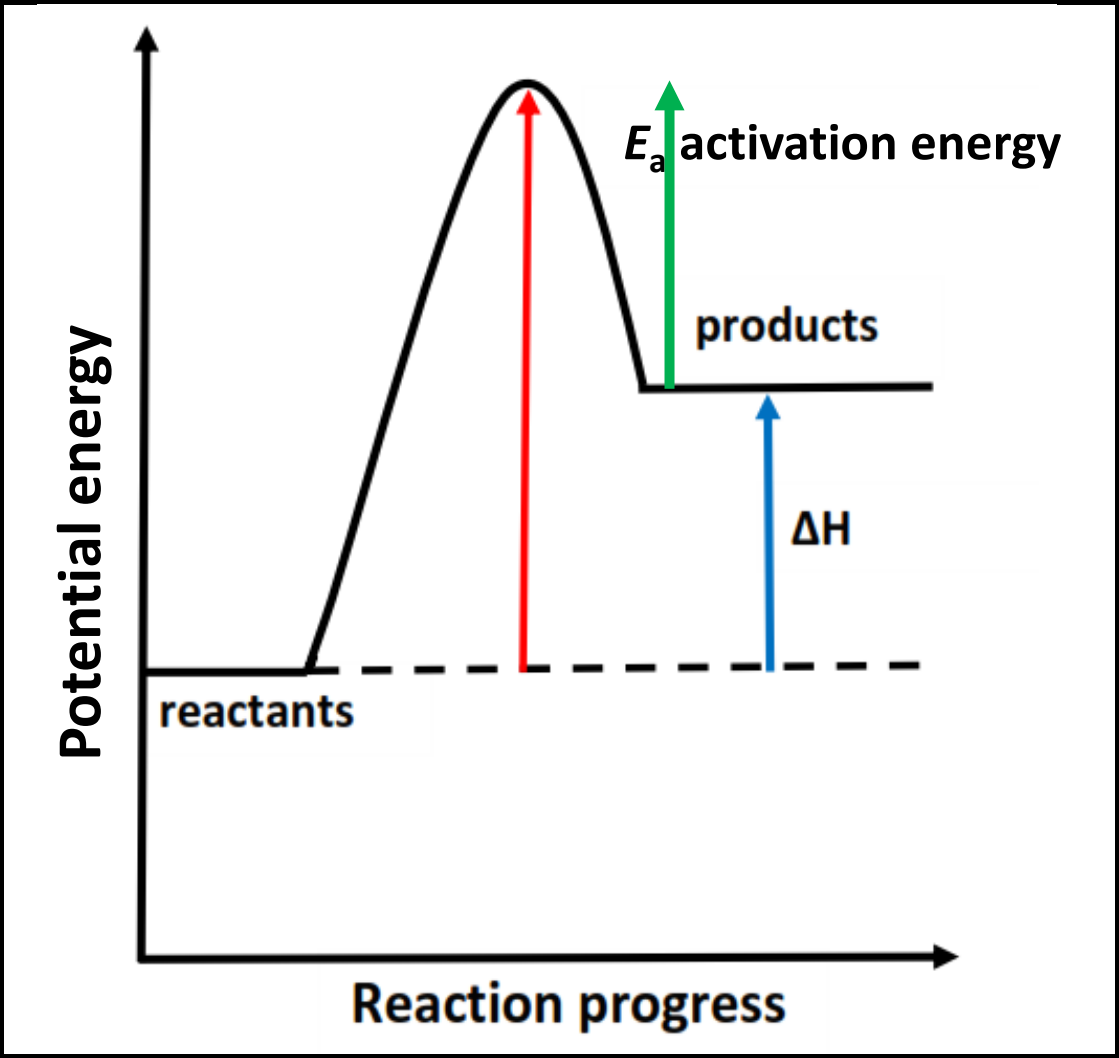
Activation energy

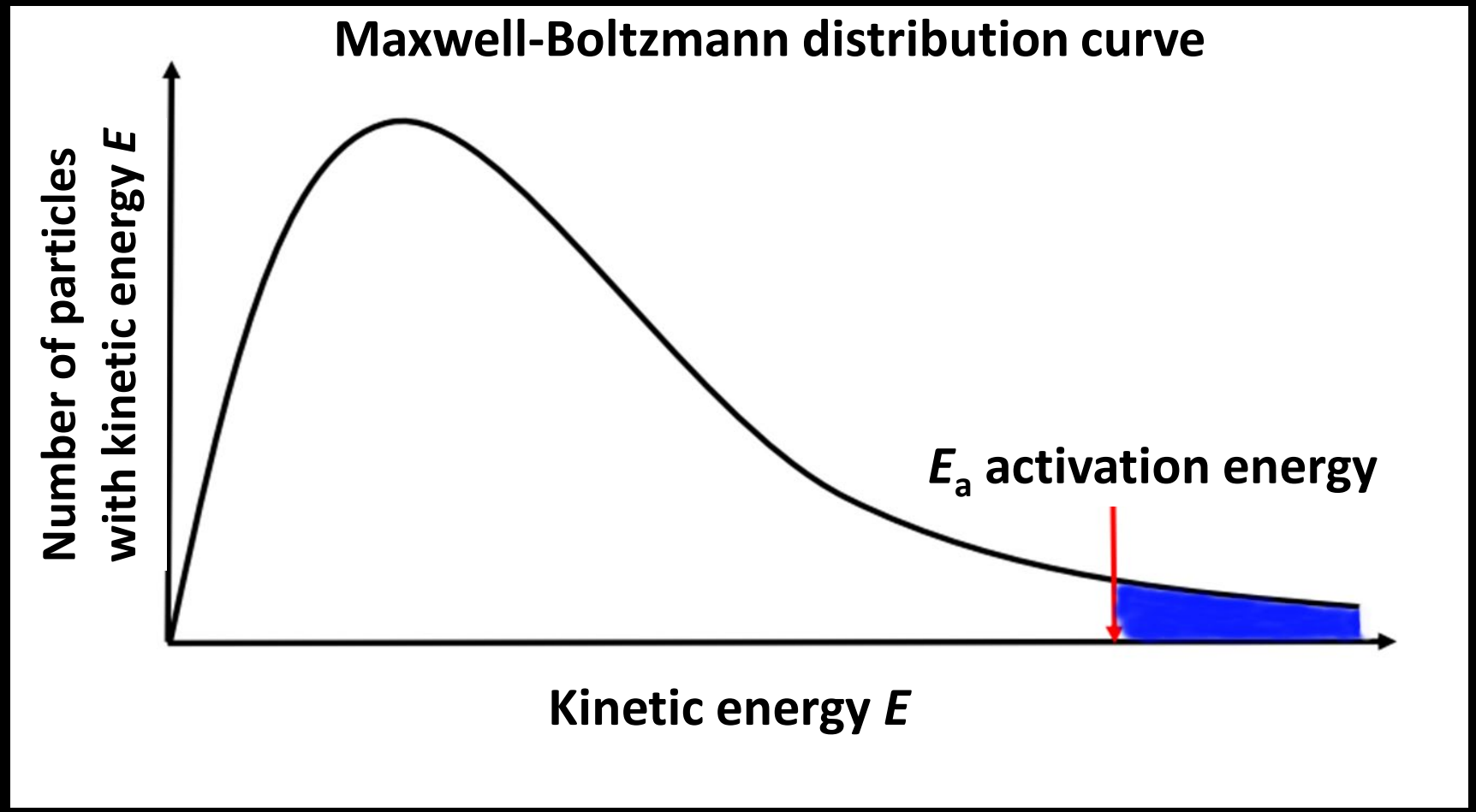
Exothermic reaction



Endothermic reaction



Activation energy



The area in white is the proportion of particles that have energy less than the activation energy ($E < E_a$).

The area in blue is the proportion of particles that have energy greater than the activation energy ($E > E_a$).

Activation energy

Temperature has no effect on the activation energy (E_a) of a reaction.

Slower reactions tend to have higher activation energies and faster reactions tend to have lower activation energies.

A catalyst provides an alternative reaction pathway with a lower activation energy – this results in an increased rate of reaction.

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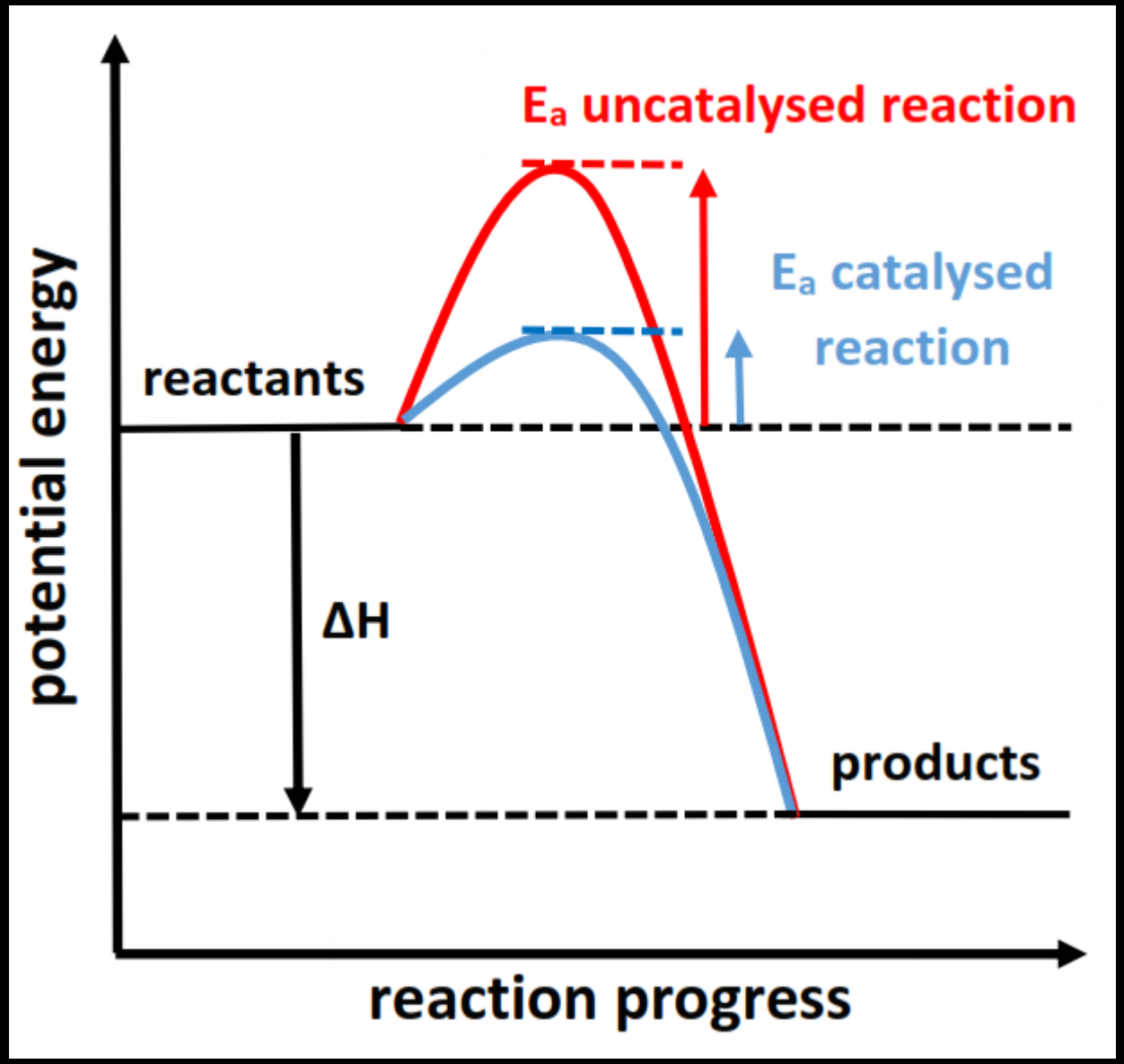
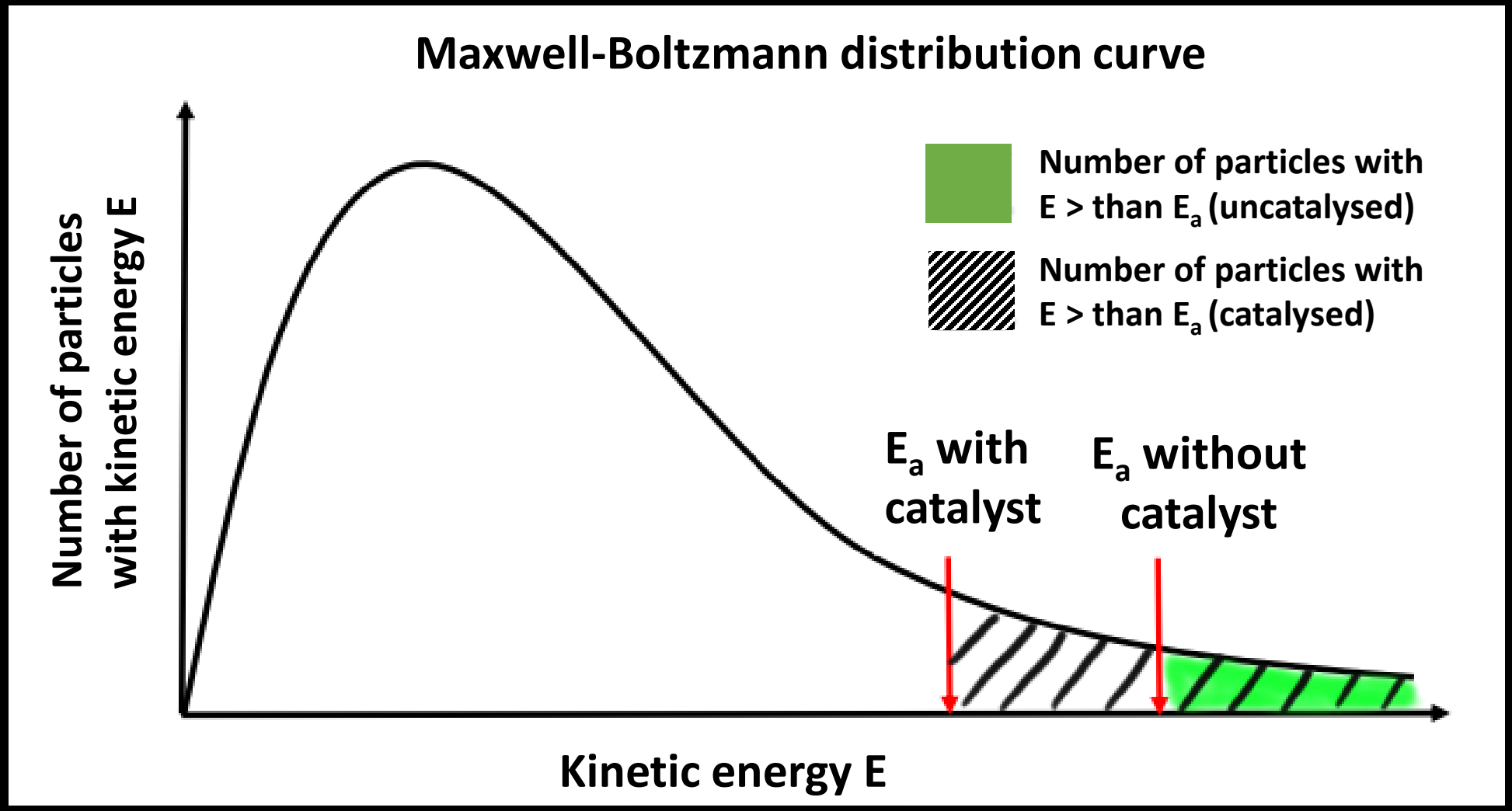
Catalysts

Catalysts



A catalyst increases the rate of a chemical reaction by providing an alternative reaction pathway with lower activation energy (E_a).

Catalysts



In the catalysed reaction, a greater proportion of particles have E > E_a, therefore the rate of reaction increases.

Catalysts

A catalyst increases the rate of a reaction by providing an alternative reaction pathway with lower activation energy.

At the same temperature, a greater proportion of reactant particles have energy greater than the activation energy. This increases the frequency of successful collisions between reactant particles, increasing the rate of reaction.

A catalyst does not undergo permanent change in a chemical reaction.

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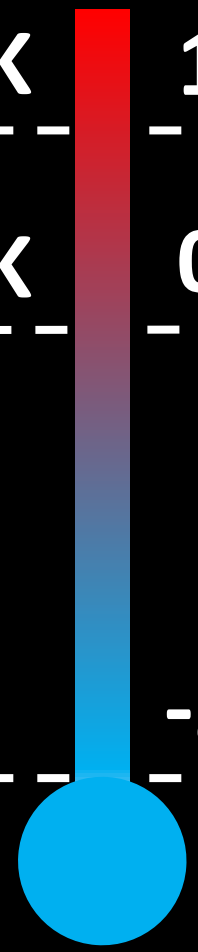
Temperature and kinetic energy

Temperature and kinetic energy

The Kelvin scale is an absolute temperature scale.



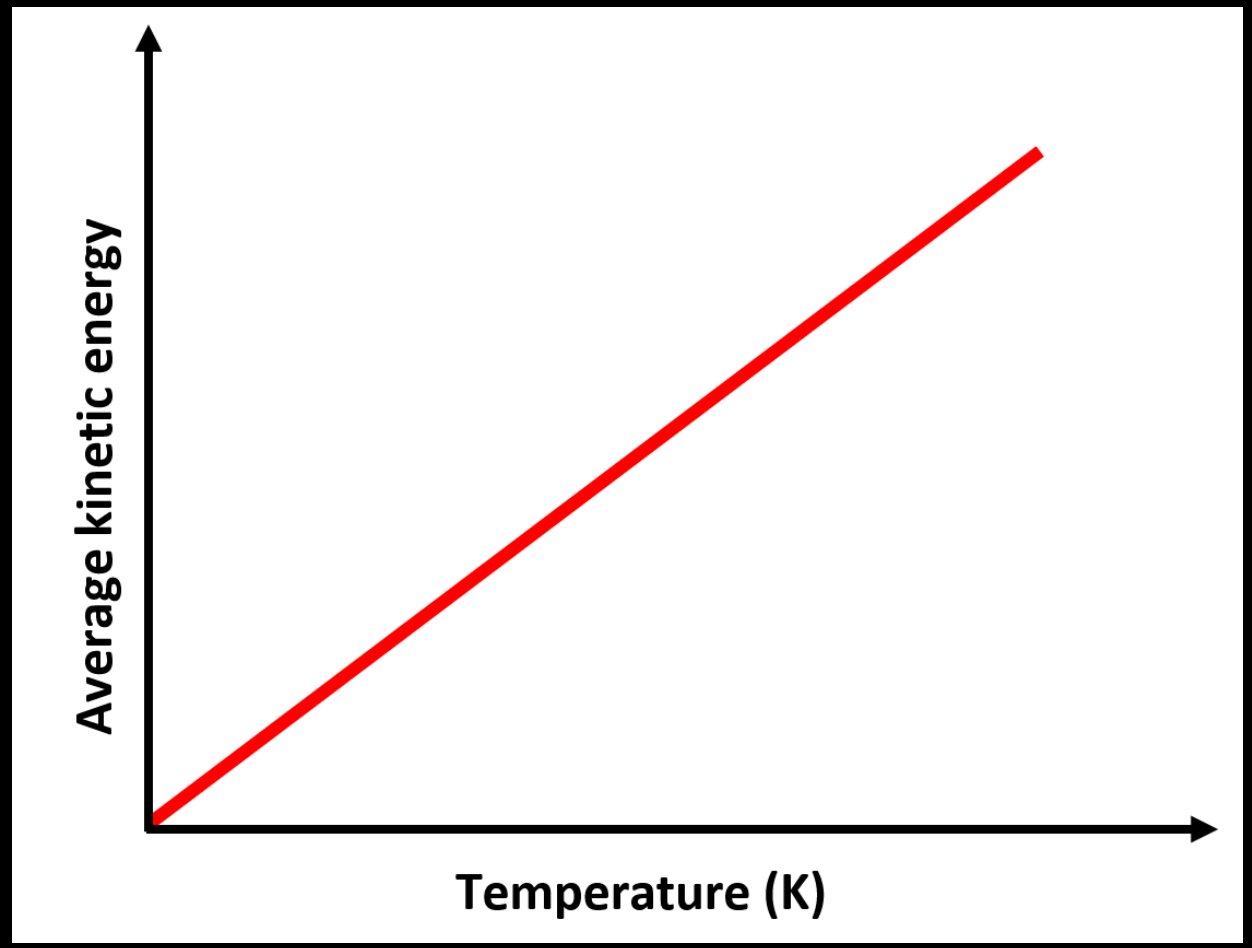
One degree on the kelvin scale is equal to one degree on the Celsius scale.

To convert from °C to K, add or subtract 273

25 °C = 298 K

Temperature and kinetic energy

Absolute temperature in kelvin (K) is directly proportional to the average kinetic energy of the particles in a substance.



$$\overline{E_k} = \frac{3}{2}\frac{R}{N_A}\mathrm{T}$$

$\overline{E_k}$ – average kinetic energy of one particle (J)

R – universal gas constant (8.31 J K^{-1} mol^{-1})

N_A – Avogadro constant (6.02×10^{23})

T – temperature in kelvin, K

The Kelvin scale

Calculate the average kinetic energy for a particle of gas at 300 K and at 600 K.

$$\overline{E_k} = \frac{3}{2}\frac{8.31}{6.02\times10^{23}}300 = 6.21 \times 10^{-21}\ \mathrm{J}$$

$$\overline{E_k} = \frac{3}{2}\frac{8.31}{6.02\times10^{23}}600 = 1.24 \times 10^{-20}\ \mathrm{J}$$

At 600 K the average kinetic energy is double that at 300 K.

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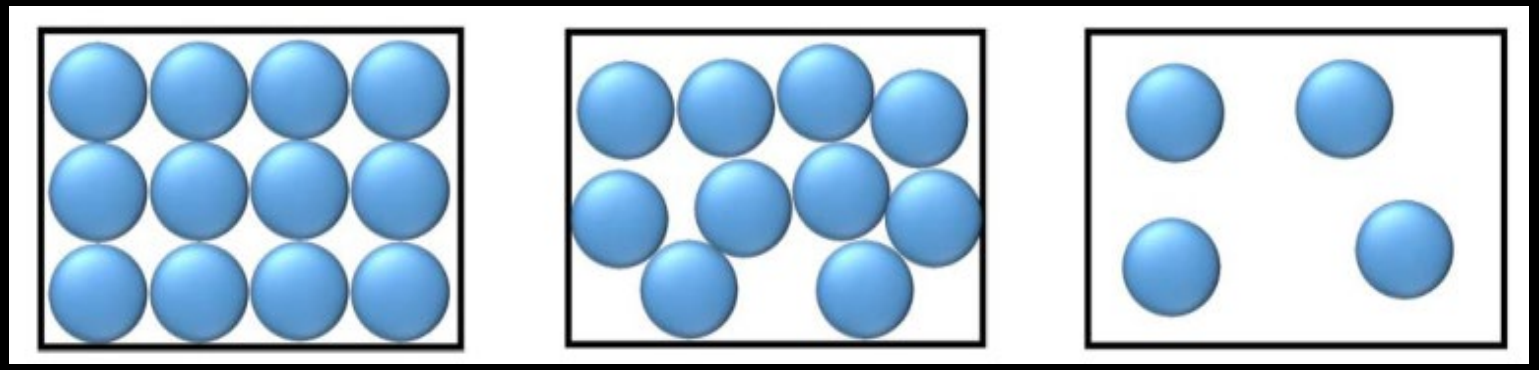
Temperature and kinetic energy

Temperature and kinetic energy

Absolute temperature is measured in kelvin (K).
It is proportional to the average kinetic energy of particles in a substance.

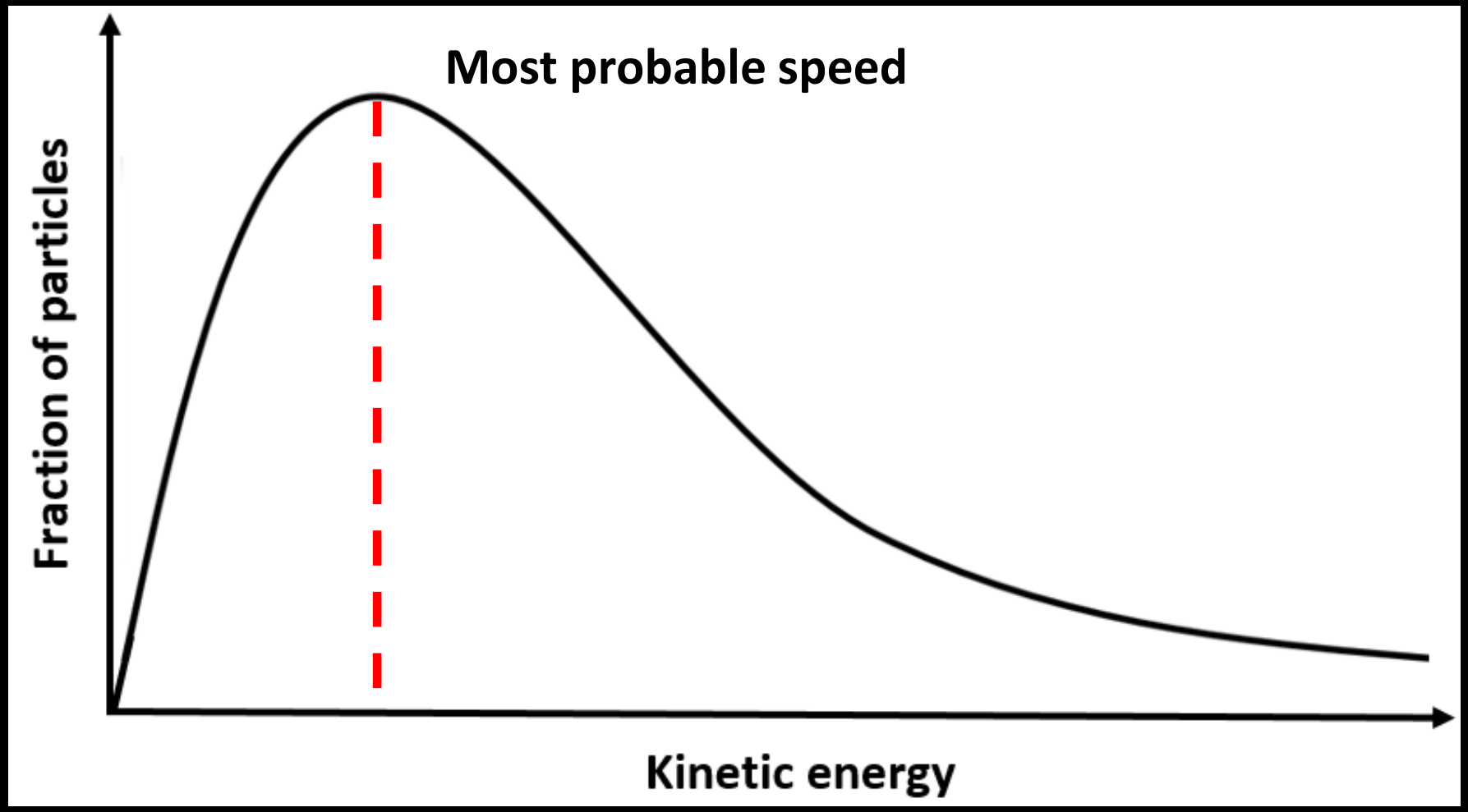


Increasing absolute temperature

Increasing average kinetic energy

Temperature and kinetic energy

A Maxwell-Boltzmann energy distribution curve shows the number of particles with a particular value of kinetic energy.



Temperature and kinetic energy

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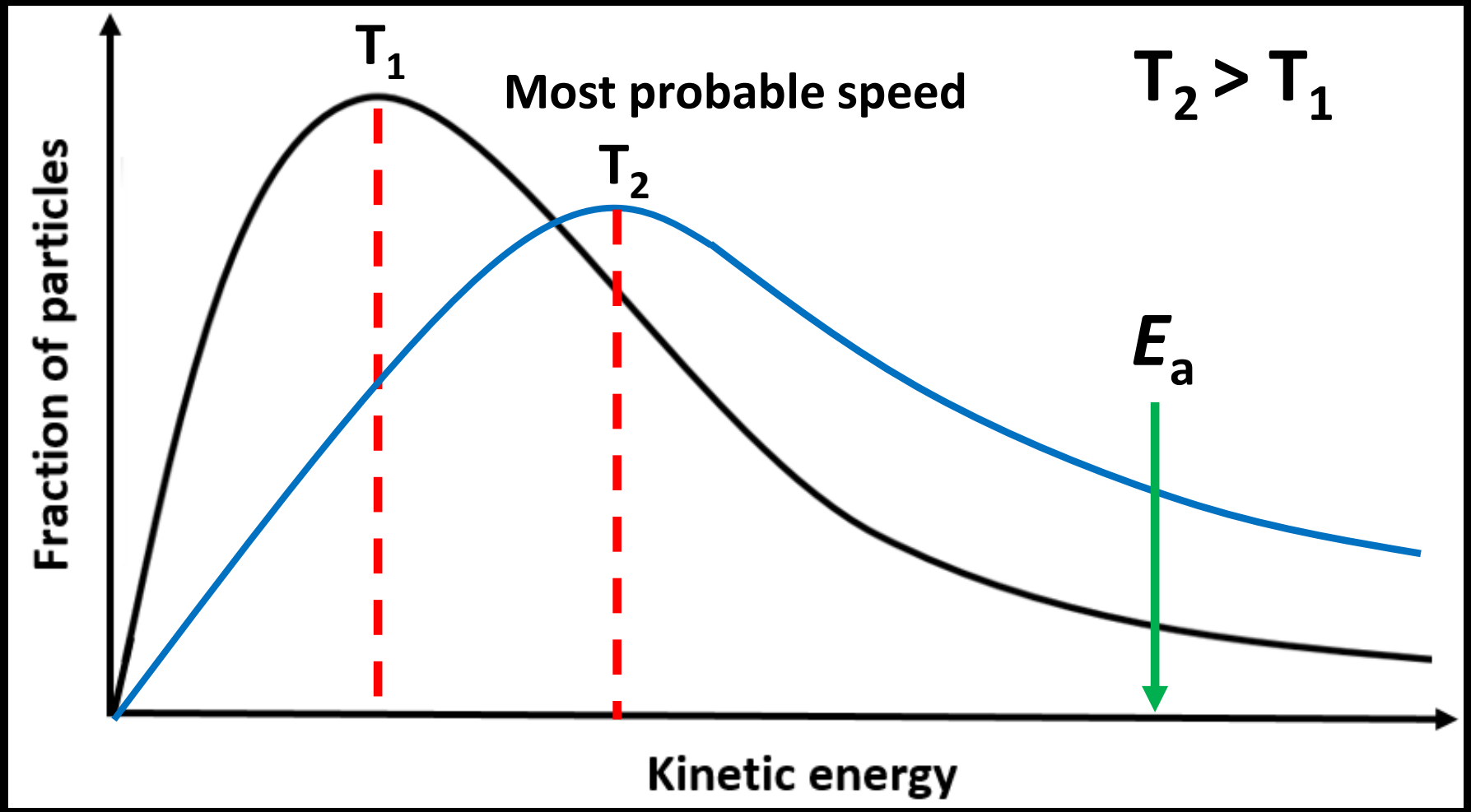

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Factors that affect the rate of reaction

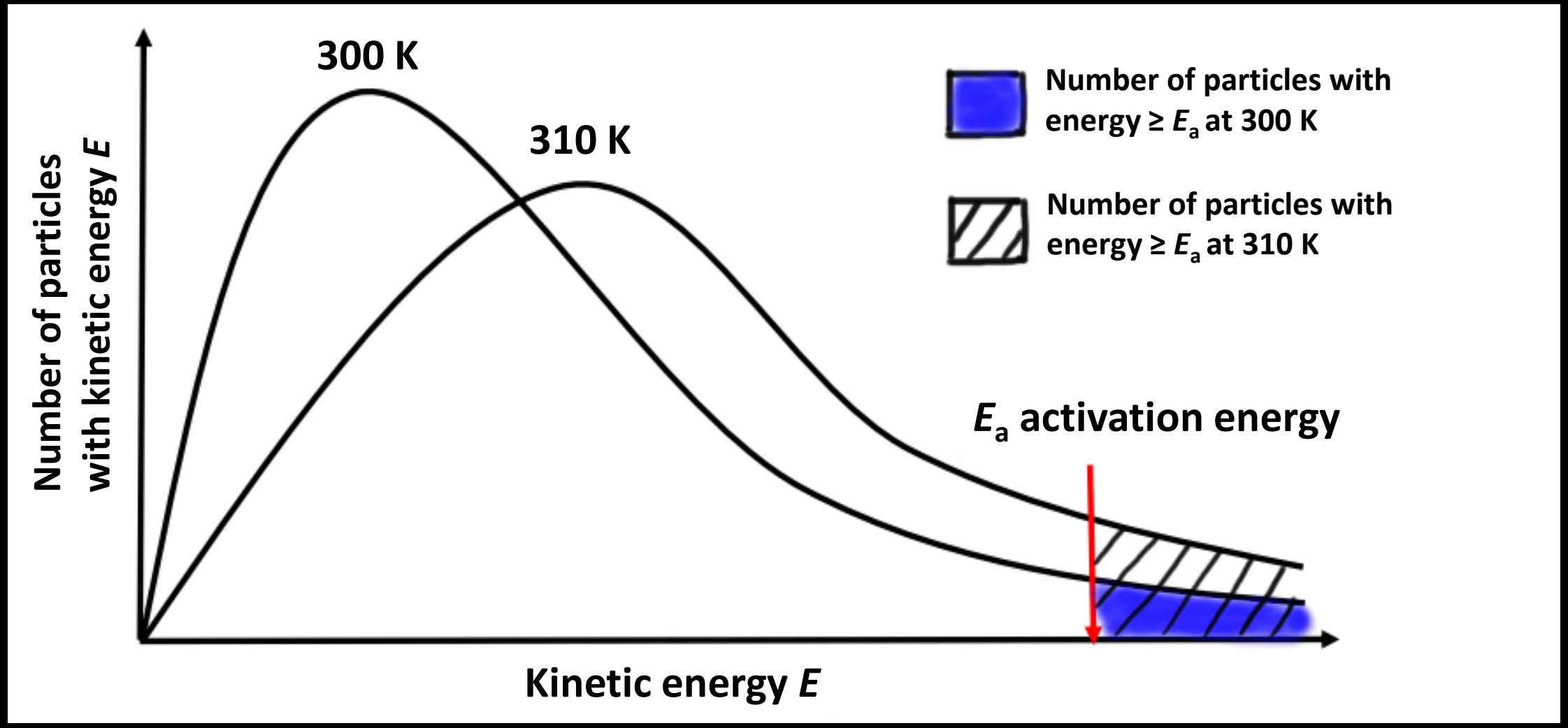
Factors that affect the rate of reaction

The rate of a chemical reaction is affected by:

- changes in temperature
- changes in concentration
- changes in particle size
- changes in pressure.

Changes in temperature

Increasing the temperature increases the rate of reaction.



Changes in temperature

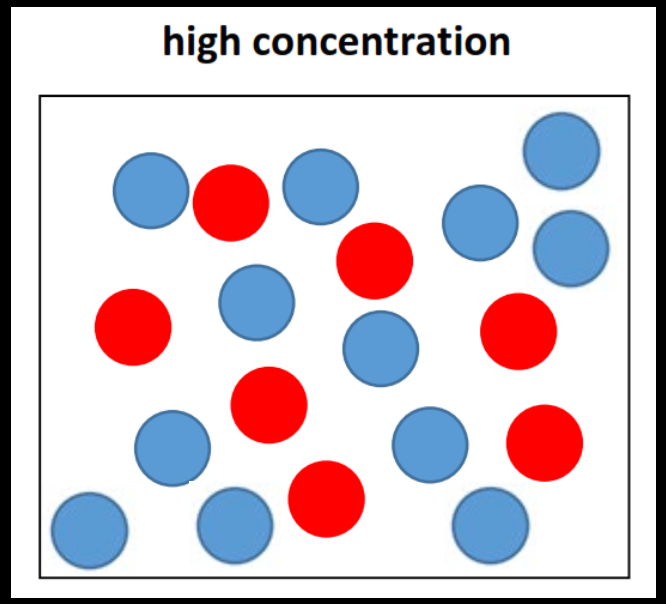
Increasing the temperature increases the average kinetic energy of the particles.

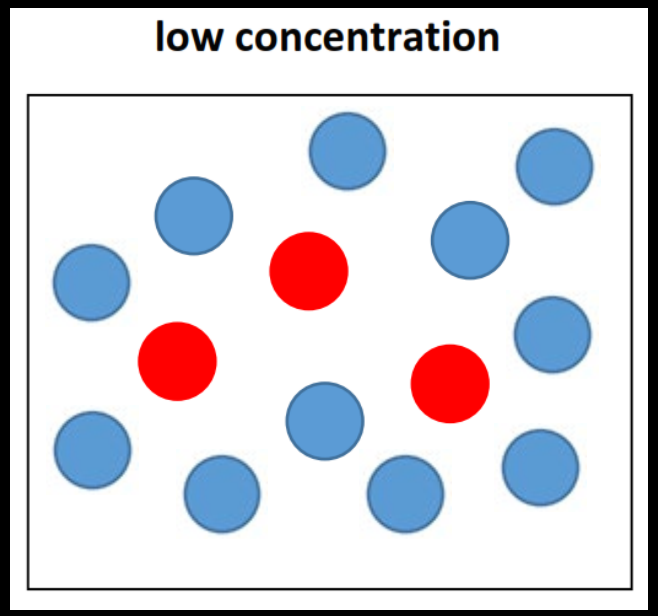
The frequency of collisions between reactant particles increases.

There is an increase in the proportion of particles that have energy equal to, or greater than, the activation energy.

Changes in concentration

Increasing the concentration of reactants increases the rate of reaction.

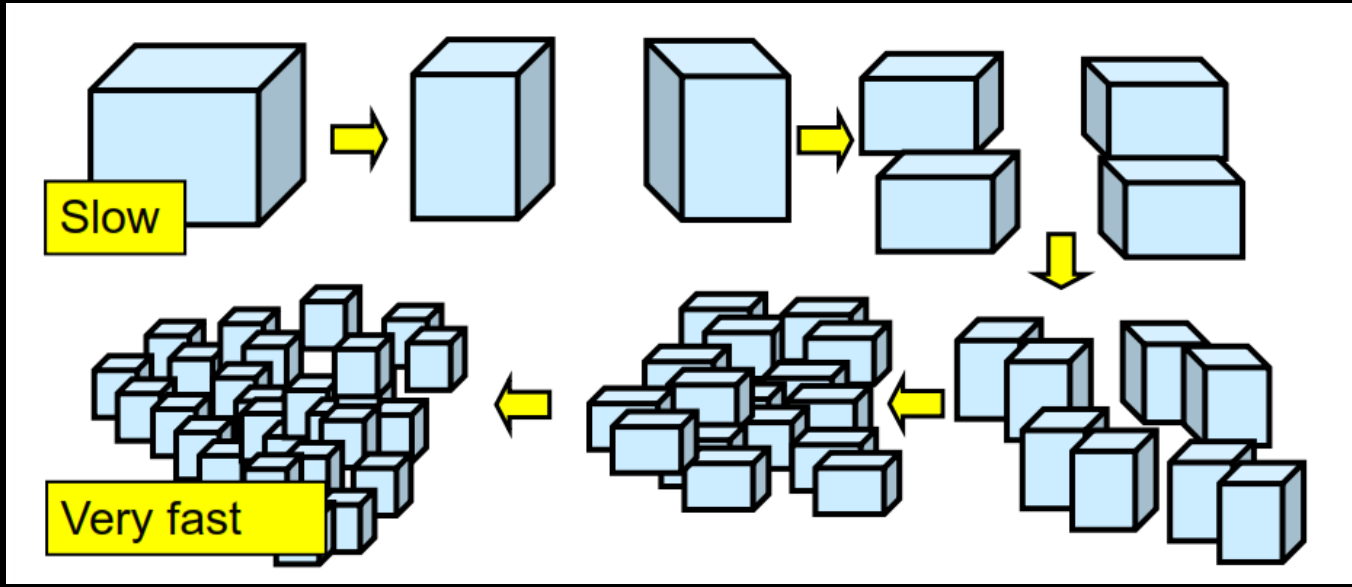




Increasing the concentration of the reactants increases the frequency of collisions between reactant particles.

Changes in surface area

As particle size decreases the rate of reaction increases.

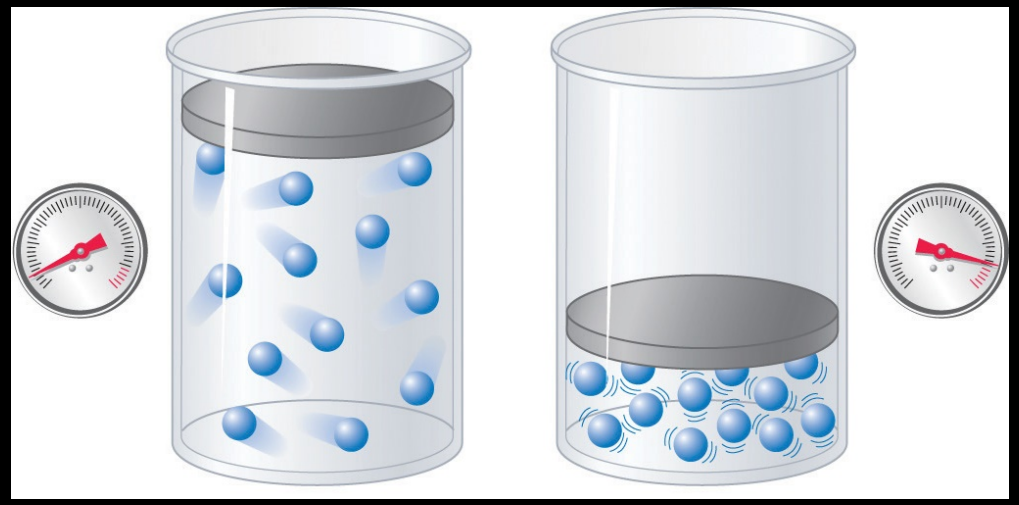


As particle size decreases, the surface area per unit volume increases which results in an increase in the frequency of collisions between reactant particles.

Changes in pressure

For reactions that involve gases, increasing the pressure increases the rate of reaction.

At higher pressures the gas is compressed which has the effect of increasing its concentration.



This results in an increase in the frequency of collisions between reactant particles.

Factors that affect the rate of reaction

Factor	Effect on collision frequency and rate of reaction
Increase temperature	Increase
Decrease temperature	Decrease
Increase concentration (pressure for gaseous reactions)	Increase
Decrease concentration (pressure for gaseous reactions)	Decrease
Increase particle size (large pieces)	Decrease
Decrease particle size (powders)	Increase