Reactivity 2.2 Answers

IB CHEMISTRY SL

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

Reactivity 2.2.1

Understandings:

- The rate of reaction is expressed as the change in concentration of a particular reactant/product per unit time.

Learning outcomes:

- Determine rates of reaction.

Additional notes:

- Calculation of reaction rates from tangents of graphs of concentration, volume or mass against time should be covered.

Linking questions:

- Tool 1, 3, Inquiry 2 Concentration changes in reactions are not usually measured directly. What methods are used to provide data to determine the rate of reactions?
- Tool 1 What experiments measuring reaction rates might use time as i) a dependent variable ii) an independent variable?

Rate of reaction

- Rate of reaction is defined as the change in concentration of reactants or products per unit time.
- The unit for rate of reaction is mol dm^{-3} s^{-1}.

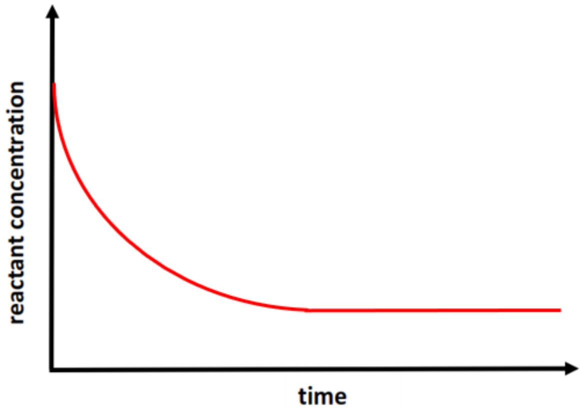
$$\textbf{rate of reaction} = \frac{\textbf{increase in product concentration}}{\textbf{change in time}}$$

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

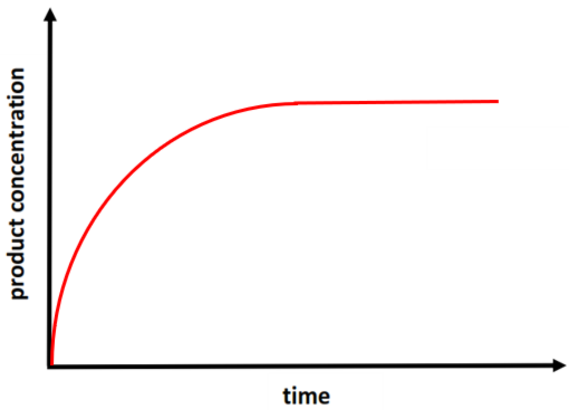
$$\text{rate of reaction} = \frac{\Delta[\text{P}]}{\Delta \text{t}} \text{ or } -\frac{\Delta[\text{R}]}{\Delta \text{t}}$$

Exercise: On the axis below sketch the following graphs.

1) The change in concentration of reactants against time.

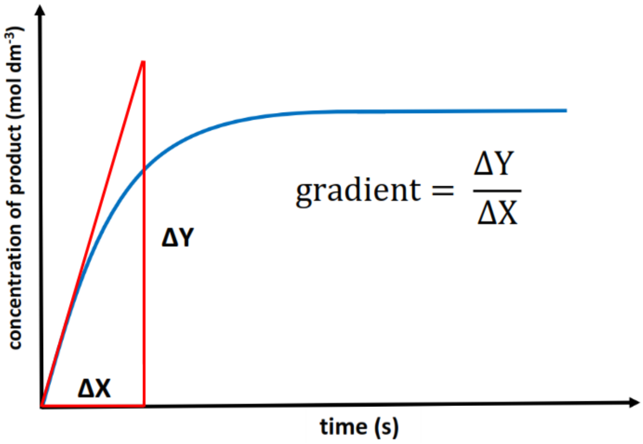


2) The change in concentration of products against time.



Rates of reaction graphs

- The instantaneous rate of reaction can be determined graphically as shown below.
- The rate of reaction decreases with time as the concentration of reactant particles decreases which results in a lower frequency of collisions between the reactant particles.
- This can be seen in the graph below, where the initial rate is fastest (the line is the steepest) but the line becomes less steep (rate of reaction is decreasing). Once the limiting reactant is consumed, no more product is formed and the line becomes horizontal.

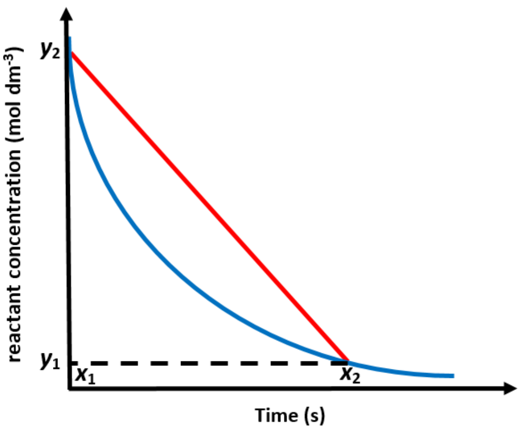


Determine the rate of reaction from a graph

- 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 line (change in y / change in x)
- The gradient of the line = the rate of reaction.

Average rate of reaction

- The average rate of reaction is the rate of reaction averaged over a certain time period.



$$\text{Average rate of reaction} = \frac{y_2 - y_1}{x_2 - x_1}$$

Exercises:

1) Outline what is meant by the term *rate of reaction.*
Change in concentration of reactant or product per unit time.

2) Outline how the initial rate of reaction can be determined from a graph.
Draw tangent at curve at time = 0
Determine gradient of line (ΔY / ΔX); the gradient is the rate of reaction.

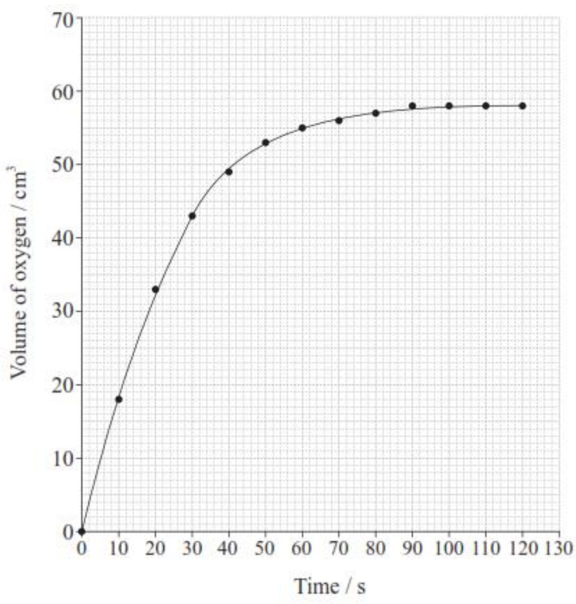
3) Explain how the rate of reaction changes with time.
Initial rate of reaction is faster due to higher concentration of reactant particles which results in a higher frequency of collisions between reactant particles. Rate of reaction decreases over time as the concentration reactant particles decreases and less frequent collisions between reactant particles.

4) Use the graph below to determine:

(a) the rate of formation of $O_2(g)$ at 40 seconds.

(b) the average rate of reaction between 0 and 60 seconds.

$$\text{Average rate of reaction} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{55-0}{60-0} = 0.92\ \text{cm}^3\ \text{s}^{-1}$$



Reactivity 2.2.2

Understandings:

- Species react as a result of collisions of sufficient energy and proper orientation.

Learning outcomes:

- Explain the relationship between the kinetic energy of the particles and the temperature in kelvin, and the role of collision geometry.

Linking questions:

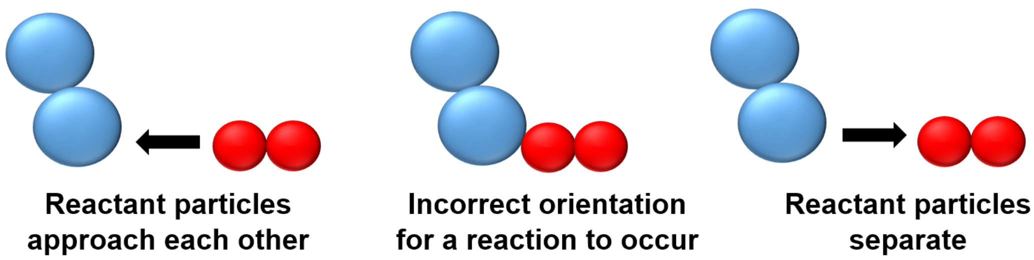
- Structure 1.1 What is the relationship between the kinetic molecular theory and collision theory?

Collision theory

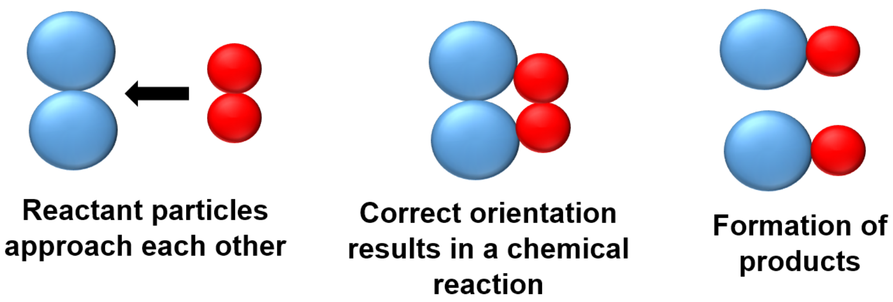
- Collision theory states that to react:
  1. reactant particles must collide with energy equal to, or greater than, the activation energy, ($E \geq E_a$).
  2. reactant particles must collide with the correct orientation (or geometry).
- Collisions that result in a chemical reaction are known as successful collisions.
- Not all collisions between reactant particles are successful – only those with the correct orientation and $E \geq E_a$.

Example of an unsuccessful and a successful collision between reactant particles

Unsuccessful collision

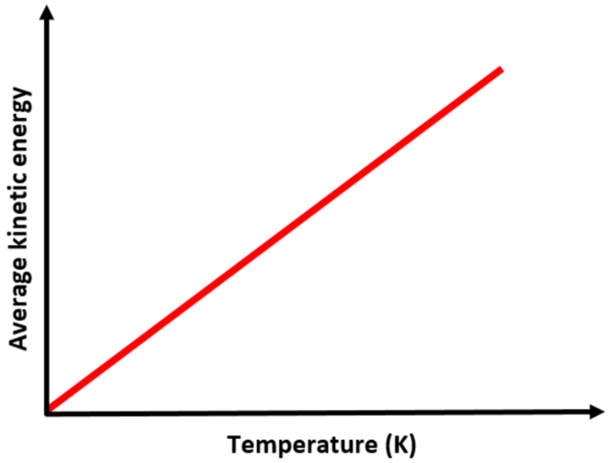


Successful collision



Temperature and kinetic energy

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



- For example, a doubling of the absolute temperature in kelvin (300 K to 600 K) results in a doubling of the average kinetic energy of the particles.
- Particles at a higher temperature have higher average kinetic energy than those at a lower temperature.
- As temperature increases, the average kinetic energy of the particles increases and as temperature decreases the average kinetic energy of the particles decreases.

Reactivity 2.2.4

Understandings:

- Activation energy, E_a, is the minimum energy that colliding particles need for a successful collision leading to a reaction.

Learning outcomes:

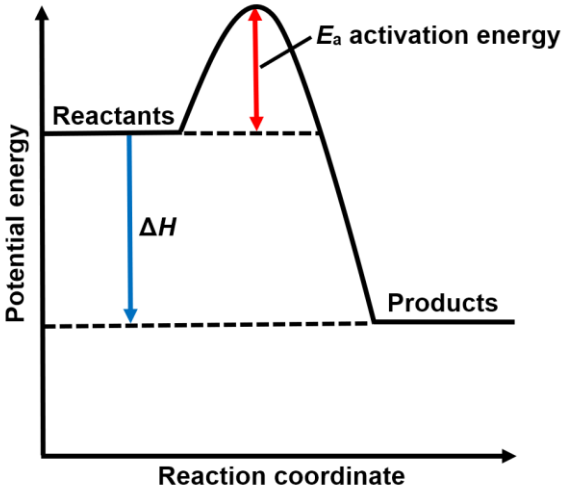
- Construct Maxwell–Boltzmann energy distribution curves to explain the effect of temperature on the probability of successful collisions.

Activation energy, E_a

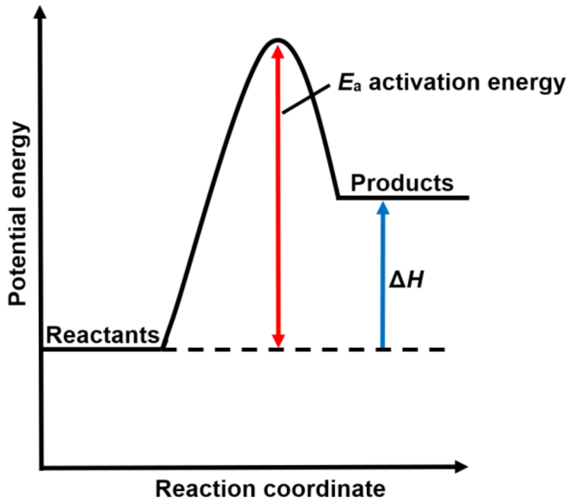
- For a chemical reaction to happen, the reactant particles must have energy equal to, or greater than, the activation energy (in addition to the correct orientation).
- Activation energy (E_a) is the minimum energy that colliding molecules need to have a successful collision leading to a chemical reaction.

Energy level profiles for an exothermic and endothermic reaction showing activation energy, E_a.

Exothermic reaction



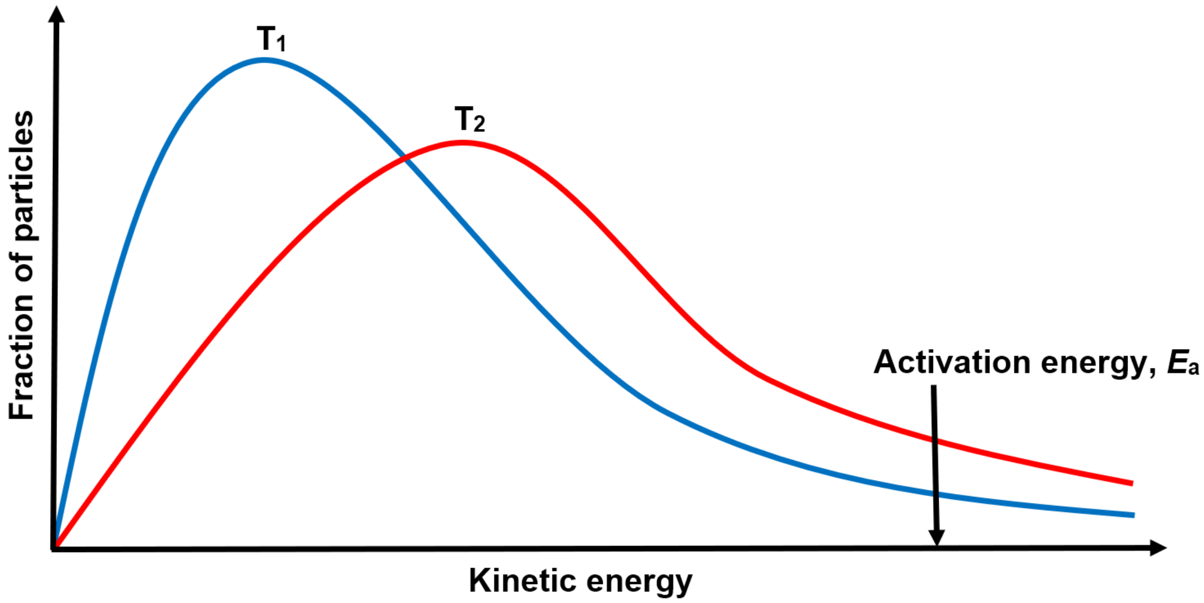
Endothermic reaction



Effect of temperature on the frequency of successful collisions

- At higher temperatures, the frequency of collisions between reactant particles increases (because of the increased kinetic energy of the particles).
- There is also an increase in the number of particles that have energy equal to, or greater than, the activation energy ($E \geq E_a$).
- These two factors combined result in an increase in the frequency of successful collisions at higher temperatures.

Maxwell- Boltzmann showing the distribution of kinetic energy in a sample of gas at two temperatures where $T_2 > T_1$.



- After the point of activation energy, the red curve has a greater area under the curve which corresponds to a greater number of particles that have $E \geq E_a$.
- This results in a higher frequency of successful collisions between reactant particles and an increase in the rate of reaction.

Exercises:

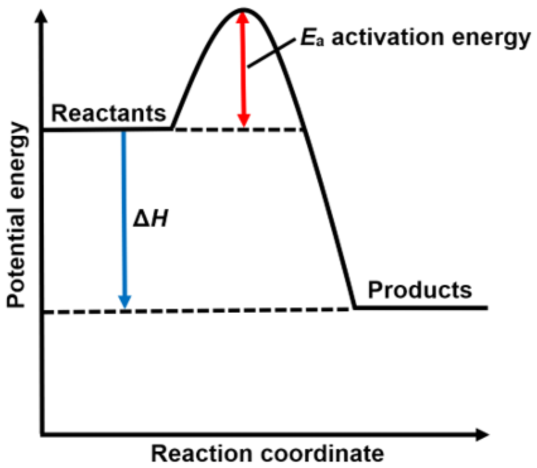
1) State two conditions necessary for a reaction to take place between reactant particles.
Reactant particles must collide with energy equal to, or greater than, the activation energy, and reactant particles must collide with the correct or orientation.

2) Outline what is meant by the term *activation energy*, E_a.

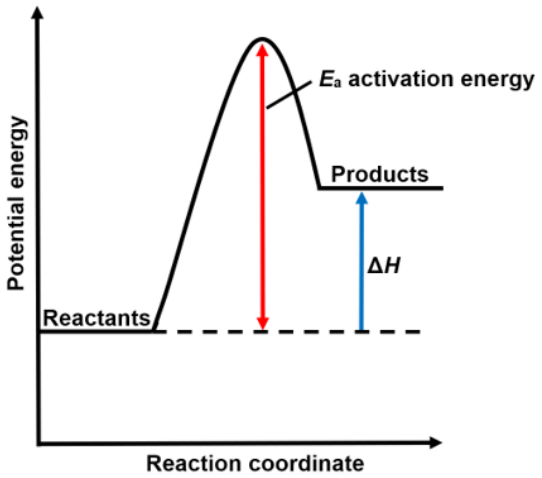
The minimum energy required by reactant particles to have a successful collision leading to a chemical reaction.

3) Sketch an energy profile for an exothermic and an endothermic reaction, labelling the activation energy, E_a.

Exothermic reaction



Endothermic reaction



Reactivity 2.2.3

Understandings:

- Factors that influence the rate of a reaction include pressure, concentration, surface area, temperature and the presence of a catalyst.

Learning outcomes:

- Predict and explain the effects of changing conditions on the rate of a reaction.

Factors that affect the rate of reaction

Pressure

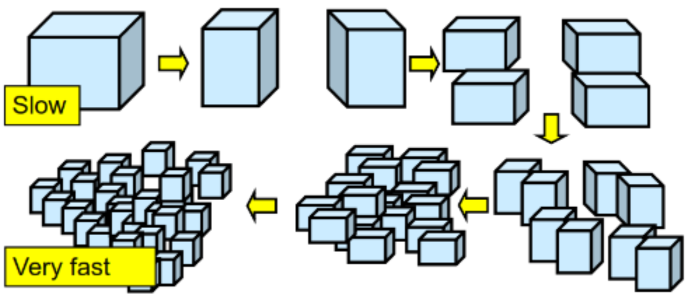
- For reactions involving gases, increasing the pressure increases the rate of reaction.
- At higher pressures, there are a greater number of reactant particles per unit volume which leads to an increase in the frequency of successful collisions between the reactant particles.

Concentration

- Increasing the concentration of a reactant increases the rate of a reaction.
- As the concentration of reactant particles increases, the frequency of successful collisions between the reactant particles also increases.

Surface area

- Decreasing the size of the reactant particles (for solid reactants) increases the surface area per unit volume.
- This results in a increase in the frequency of successful collisions between the reactant particles.
- For this reason, powdered reactants react faster than large lumps.



Temperature

- Increasing the temperature increases the rate of reaction.
- At higher temperatures, the frequency of collisions between reactant particles increases (because of the increased kinetic energy of the particles).
- There is also an increase in the number of particles that have $E \geq E_a$.
- These two factors combined result in an increase in the frequency of successful collisions at higher temperatures.

Exercises:

1) State and explain the effect of the following changes on the rate of reaction for the following reaction:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

a) Increasing the temperature of the hydrochloric acid.
The rate of reaction increases. Increasing the temperature increases the average kinetic energy of the reactant particles resulting in a higher frequency of collisions between reactant particles. The number of reactant particles with energy equal to or greater than the activation energy also increases.

b) Increasing the concentration of the hydrochloric acid.
Increasing the concentration of the hydrochloric acid increases the frequency of collisions between reactant particles, therefore, the rate of reaction increases.

c) Increasing the particle size of the calcium carbonate.
Increasing the particle size will decrease the rate of reaction. Larger particles have a smaller surface area per unit volume which results in a lower frequency of collisions between reactant particles.

2) When magnesium ribbon is added to hydrochloric acid, the following reaction takes place:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

State and explain one change that could be made to the reactants to increase the rate of reaction and one change to decrease the rate of reaction.
Increase the rate of reaction:

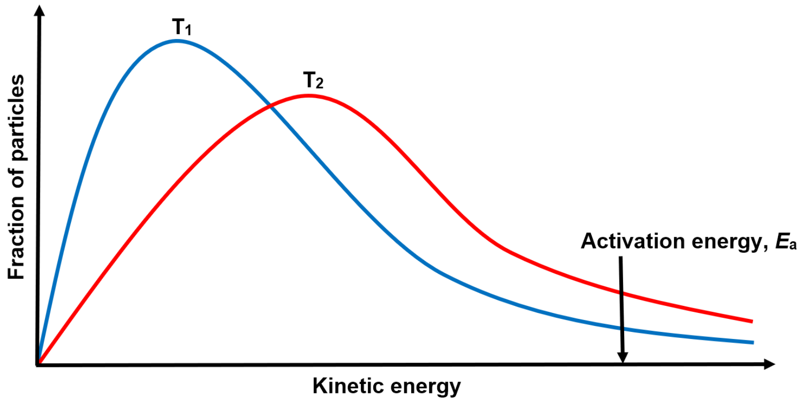
- Increase the concentration of the hydrochloric acid
- Increase the temperature of the hydrochloric acid
- Use powdered magnesium

Decrease the rate of reaction:

- Decrease the concentration of the hydrochloric acid
- Decrease the temperature of the hydrochloric acid

3) Graphs are an effective method to follow the rate of a chemical reaction.

a) Draw a graph to show the distribution of energies in a sample of gas molecules. Label the axis and label your curve T_1. Using the same axis, draw a second curve to represent the distribution of energies at a higher temperature. Label this curve T_2.



b) State and explain, with reference to your graph, what happens to the rate of a reaction when the temperature is increased.

Increasing the temperature increases the average kinetic energy of the reactant particles resulting in a higher frequency of collisions between reactant particles. The number of reactant particles with energy equal to or greater than the activation energy also increases.

Reactivity 2.2.5

Understandings:

- Catalysts increase the rate of reaction by providing an alternative reaction pathway with lower E_a.

Learning outcomes:

- Sketch and explain energy profiles with and without catalysts for endothermic and exothermic reactions.
- Construct Maxwell–Boltzmann energy distribution curves to explain the effect of different values for E_a on the probability of successful collisions.

Additional notes:

- Biological catalysts are called enzymes.
- The different mechanisms of homogeneous and heterogeneous catalysts will not be assessed.

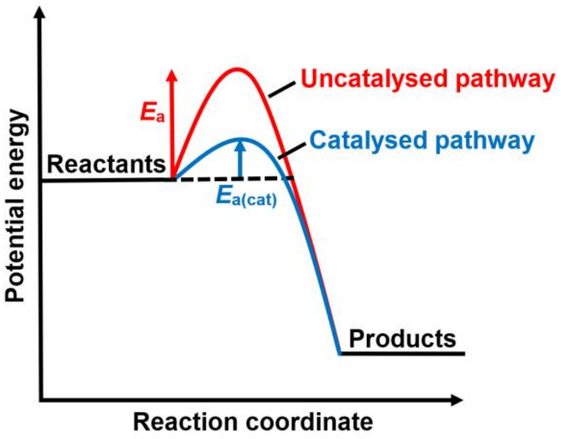
Linking questions:

- Reactivity 2.3 What is the relative effect of a catalyst on the rate of the forward and backward reactions?
- Structure 3.1 (HL) What are the features of transition elements that make them useful as catalysts?

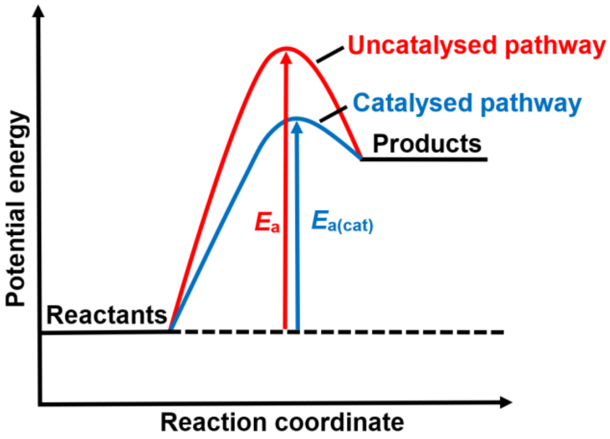
Catalysts

- A catalyst is a substance that increases the rate of a reaction without being consumed in the reaction.
- Catalysts increase provide an alternative reaction pathway for the reaction that has a lower activation energy.
- As can be seen in the energy level profiles below, the catalysed reaction pathway has a lower activation energy than the uncatalysed pathway.
- This results in a greater number of particles that have $E \geq E_a$.

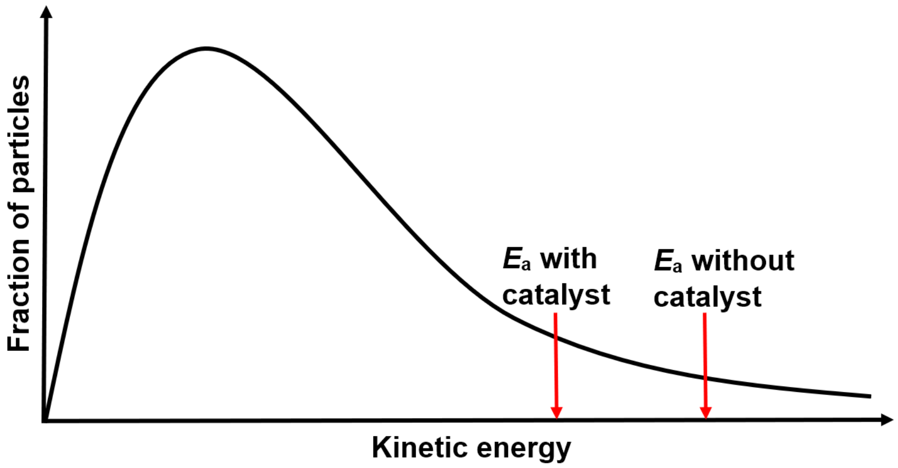
Catalysed exothermic reaction



Catalysed endothermic reaction



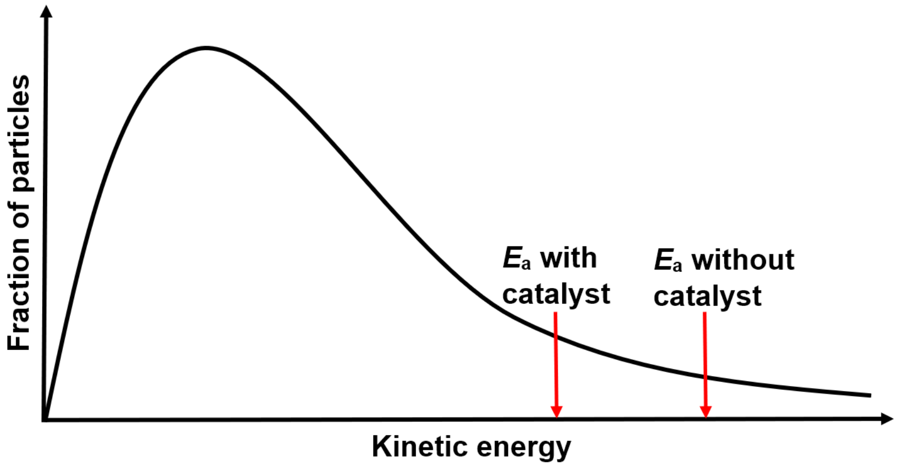
Maxwell Boltzmann curve demonstrating the effect of a catalyst on the fraction of particles with $E \geq E_a$.



- As can be seen from the diagram, the lower the activation energy, the greater the number of particles with $E \geq E_a$.

Exercises:

1) Annotate the Maxwell–Boltzmann distribution curve showing the activation energies, E_a, for the catalysed and uncatalysed reactions.



2) Explain how a catalyst increases the rate of reaction.

A catalyst increases the rate of reaction by providing an alternative reaction pathway with a lower activation energy. The number of particles with $E \geq E_a$ increases.