Topic 9: Kinetics I

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

9.1 Understand, in terms of collision theory, the effect of changes in concentration, temperature, pressure and surface area on rate.

For a reaction to occur, particles must collide with sufficient energy (at least equal to the activation energy, Ea) AND with the correct orientation.

Increasing concentration (solutions): more particles per unit volume → more frequent collisions → faster rate.

Increasing pressure (gases): molecules are closer together → more frequent collisions → faster rate. (Same effect as increasing concentration.)

Increasing temperature: particles move faster → more frequent collisions AND a greater proportion of collisions have energy ≥ Ea → much faster rate. Temperature has a larger effect than concentration because it increases both collision frequency AND the proportion of successful collisions.

Increasing surface area (solids): more particles exposed at the surface → more collisions possible per unit time → faster rate.

9.2 Understand that reactions only take place when collisions have sufficient energy, known as activation energy.

Activation energy (Ea) is the minimum energy that colliding particles must possess for a collision to result in a reaction. Only collisions with energy ≥ Ea lead to product formation. This is why not every collision between reactant particles is successful – most collisions simply bounce off without reacting.

9.3 Be able to calculate rate of reaction from: (i) time data, (ii) gradient of a graph.

(i) Rate from time data: rate = 1/time (for clock reactions where a fixed observable change occurs). Or: rate = change in quantity / time. E.g. if a precipitate forms after 50 s, rate ≈ 1/50 = 0.02 s^{-1}.

(ii) Rate from a concentration-time graph: draw a tangent to the curve at the required time and calculate the gradient of the tangent. Gradient = $\Delta y/\Delta x$ = change in concentration / change in time. For initial rate: draw the tangent at t = 0. Units of rate depend on axes (e.g. mol dm^{-3} s^{-1}).

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

9.4 Understand qualitatively, using the Maxwell-Boltzmann distribution, how temperature affects rate.

The Maxwell-Boltzmann distribution shows the spread of molecular energies in a sample of gas at a given temperature. Key features:

- The curve starts at the origin (no molecules have zero energy).
- It rises to a peak (the most probable energy) then falls with a long tail to the right.
- The area under the entire curve represents the total number of molecules (constant).

At higher temperature:

- The peak shifts to the right (higher most probable energy) and is lower (flatter curve, same total area).
- The area under the curve to the right of Ea increases significantly – a much greater proportion of molecules now have energy ≥ Ea.
- This is the primary reason rate increases so dramatically with temperature.

9.5 Understand the role of catalysts in providing alternative reaction routes of lower activation energy.

A catalyst increases the rate of reaction by providing an alternative reaction pathway with a lower activation energy. This means a greater proportion of molecules have energy ≥ the new (lower) Ea at the same temperature, so more successful collisions occur per unit time.

Important: the catalyst is not consumed in the overall reaction – it is regenerated. A catalyst does NOT:

- Change the enthalpy change (ΔH) of the reaction
- Change the position of equilibrium (for a reversible reaction)
- Appear in the overall equation

A catalyst only increases the rate at which equilibrium is reached.

9.6 Be able to draw reaction profiles for uncatalysed and catalysed reactions.

A reaction profile (energy profile diagram) shows:

- y-axis: enthalpy/energy
- x-axis: reaction progress (or reaction coordinate)
- Reactants at the start, products at the end
- A hump representing the activation energy barrier (Ea = difference between reactant energy level and the peak)
- ΔH = difference in energy between reactants and products

For the catalysed profile:

- The hump is lower (smaller Ea) but the start and end points are the same (same reactants, same products, same ΔH)
- Draw both profiles on the same axes to show the comparison clearly

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

9.7 Be able to interpret catalyst action using the Maxwell-Boltzmann distribution.

On a Maxwell-Boltzmann distribution diagram, show catalyst action by drawing a new, lower Ea line to the left of the original Ea line (while leaving the distribution curve unchanged). The area under the curve to the right of the new (lower) Ea is larger than for the original Ea, showing that a greater proportion of molecules now have sufficient energy to react. This explains why the rate is faster with a catalyst at the same temperature.

9.8 Understand the use of a solid (heterogeneous) catalyst for industrial gas-phase reactions.

A heterogeneous catalyst is in a different phase from the reactants (typically a solid catalyst with gaseous or aqueous reactants). Mechanism:

1. Reactant molecules adsorb onto the surface of the solid catalyst (bonds form between reactants and the surface).
2. The adsorbed molecules are held in close proximity, bonds within them are weakened.
3. Reaction occurs on the surface with a lower activation energy.
4. Products desorb from the surface, freeing the catalyst.

Industrial examples:

- Iron (solid) in the Haber process: $N_2(g) + 3H_2(g) \quad 2NH_3(g)$
- V_2O_5 (solid) in the Contact process: $2SO_2(g) + O_2(g) \quad 2SO_3(g)$
- Platinum/palladium/rhodium (solid) in catalytic converters: $CO + NO \rightarrow CO_2 + N_2$

9.9 Understand the economic benefits of catalysts in industrial reactions.

Catalysts provide significant economic benefits in industrial processes:

- Allow reactions to proceed at lower temperatures and pressures → reduces energy costs and capital costs (cheaper, less robust equipment needed).
- Increase the rate of reaction → more product can be made per unit time → greater productivity.
- Catalysts are not consumed → they can be reused many times, reducing raw material costs.
- Lower activation energy → reactions viable at commercially useful rates under mild conditions.
- Can improve selectivity → fewer unwanted by-products, improving atom economy and reducing waste disposal costs.

ABOUT THESE NOTES

Notes written by Chloe Archard (MSc Biochemistry, Oxford-trained chemist, 20+ years teaching experience at top UK independent schools). For 1-1 tuition, revision courses, and more free resources: chemistrywithchloe.com

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.