

CfE Higher Chemistry
Unit 1
Chemical changes and structures

Topic	Page
1 – Controlling the Rate	2
Minitest	11
2 – Periodicity	15
3 - Bonding in Elements	18
Minitest	22
4 – Bonding and Structure	23
Minitest	32
Glossary	33

Information sourced from BBC Bitesize – Higher Chemistry

1 - Controlling the Rate

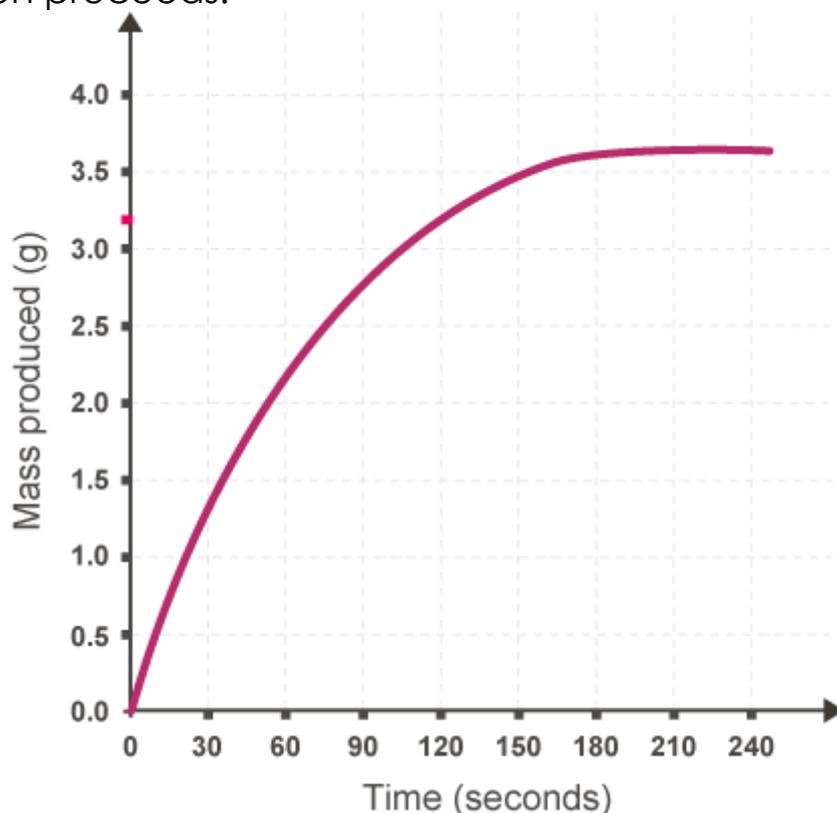
a) Reaction rates

It is important that chemists can control the rate of chemical reactions to ensure that processes are both economically viable (they will result in a good yield of products and profits for the company) and safe (the reaction does not progress too quickly potentially causing explosions).

The rate of a chemical reaction is proportional to concentration of reactants present. As reactants are used up during the process, the rate will decrease, and the reaction slows down.

By monitoring a chemical reaction and making measurements on how volume, concentration or mass change, the rate can be calculated.

The graph below shows how the rate in a chemical reaction changes as the reaction proceeds.



The average rate of reaction can be calculated by considering how the mass changes over a fixed period of time. This will give a rate measured in grams per second (g s^{-1}) using the formula:

$$\text{average rate} = \frac{\Delta \text{quantity}}{\Delta \text{time}}$$

The relative rate of reaction is the rate at any one particular point in time. This could not be measured using the results of an experiment, but since the rate of the reaction is proportional to time, relative rate can be given by the formula:

$$\text{relative rate} = \frac{1}{t}$$

For example, the relative rate of a reaction at 20 seconds will be $1/20$ or 0.05 s^{-1} , while the average rate of reaction over the first 20 seconds will be the change in mass over that period, divided by the change in time. Note that the units of relative rate are s^{-1} as no measurable change is being observed, whereas for average rate the unit used depends on the measurable quantity.

In the above graph, since a change in mass is measured in grams and a change in time is measured in seconds (in this example), the unit of rate would be grams per second (g s^{-1}).

Similarly, if a change in concentration is measured (in mol l^{-1}), then rate will have the unit moles per litre per second ($\text{mol l}^{-1} \text{ s}^{-1}$).

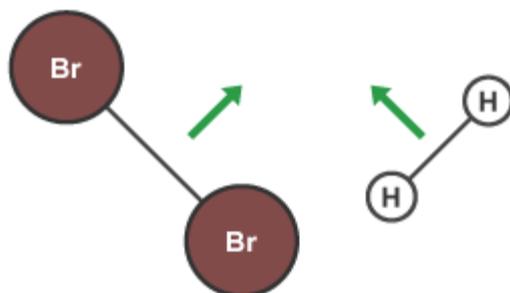
If a change in volume is measured (in cubic centimetres, cm^3), the unit of rate would be centimetres cubed per second ($\text{cm}^3 \text{ s}^{-1}$).

b) Collision theory

For a chemical reaction to occur, the reactant molecules must collide with enough energy. The minimum kinetic energy required for a reaction to occur is called the activation energy (E_A).

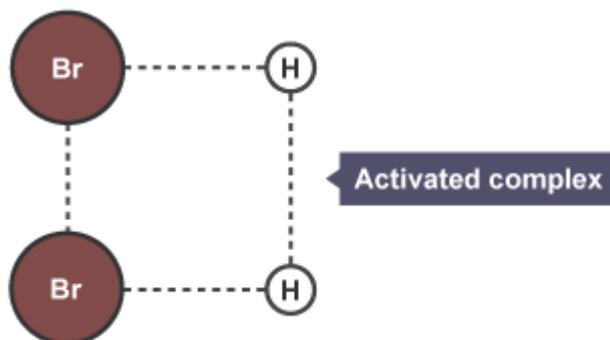
This example shows the stages of reaction between hydrogen and bromine.

Reactant molecules collide



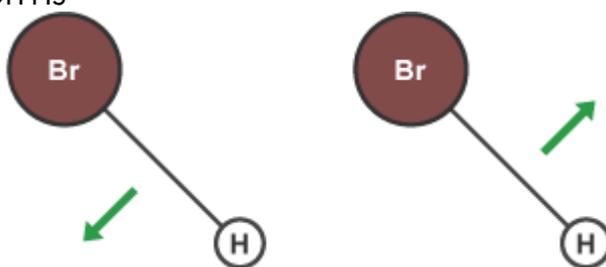
As the reactant molecules collide they must have enough energy to overcome the repulsive forces (caused by outer electrons) and start to break the bonds between the atoms.

Activated complex



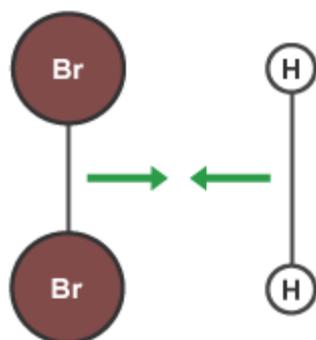
An intermediate stage is reached in which a high energy, unstable arrangement of atoms is formed called the activated complex.

Product molecule forms



Energy is given out as new bonds are formed and the atoms are rearranged into the product molecule(s).

For a successful collision to take place, the collision geometry must be right (the reactant molecules have to be facing the right way!) so that the activated complex can be formed. Looking at the reaction between hydrogen and bromine:



Potentially successful collision geometry



Unsuccessful collision as no activated complex can be formed

c) Altering factors

The rate that reactant molecules collide can be controlled by altering any of the four factors:

- temperature
- concentration
- particle size
- use of a catalyst

Only some of the collisions that take place cause a chemical change to happen. These are called 'successful' collisions. Greater the number of 'successful' collisions increases reaction rate.

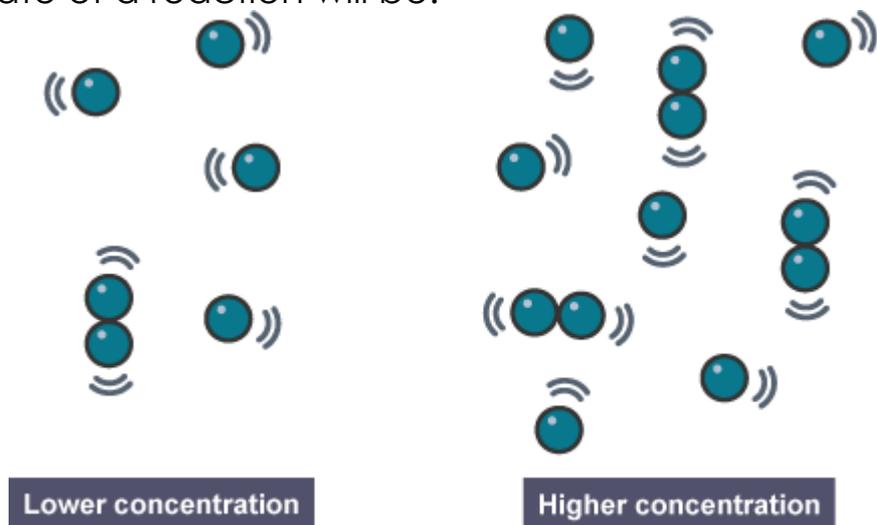
Temperature

If the temperature is increased, the particles have more energy and so move more quickly. Increasing the temperature increases the rate of reaction because the particles collide more often.



Concentration

If the concentration of reactants is increased, there are more reactant particles moving together. There will be more collisions and so the reaction rate is increased. The higher the concentration of reactants, the faster the rate of a reaction will be.

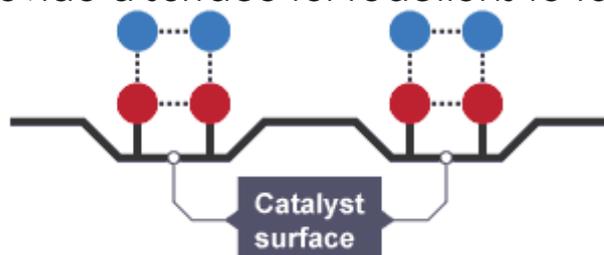


Particle size

By decreasing the particle size of a reactant, we are increasing its surface area. A smaller particle size of reactants provides a greater surface area that collisions can take place on. The greater the surface area increases rate of reaction.

Use of a catalyst

A catalyst can provide a surface for reactions to take place on.



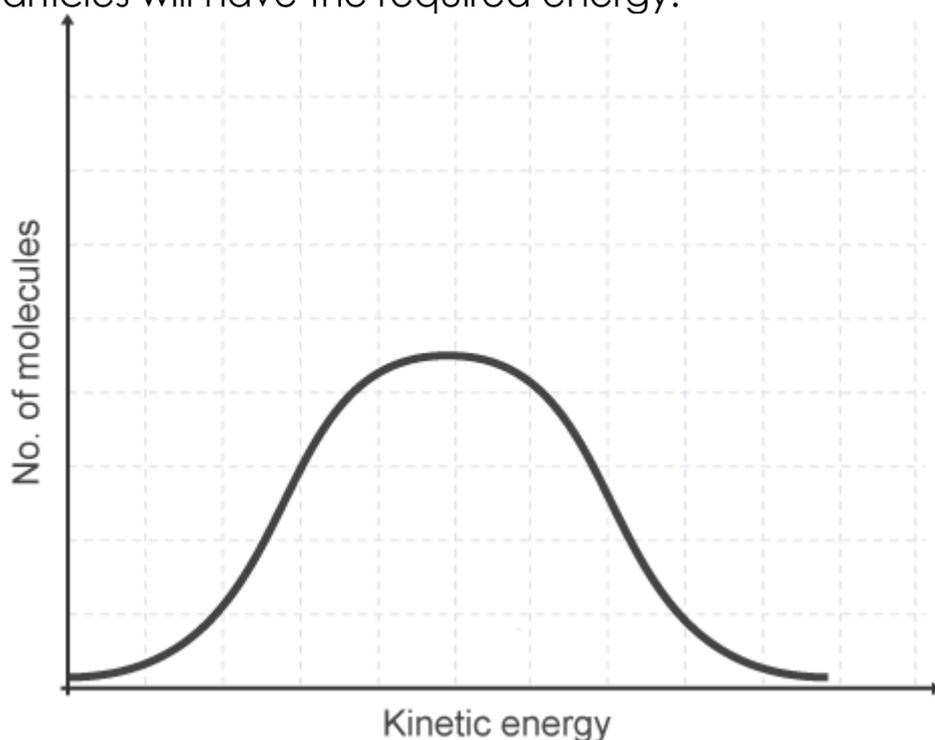
Reactant molecules are held at a favourable angle for collisions to occur, increasing the likelihood of successful collisions.

d) Activation energy

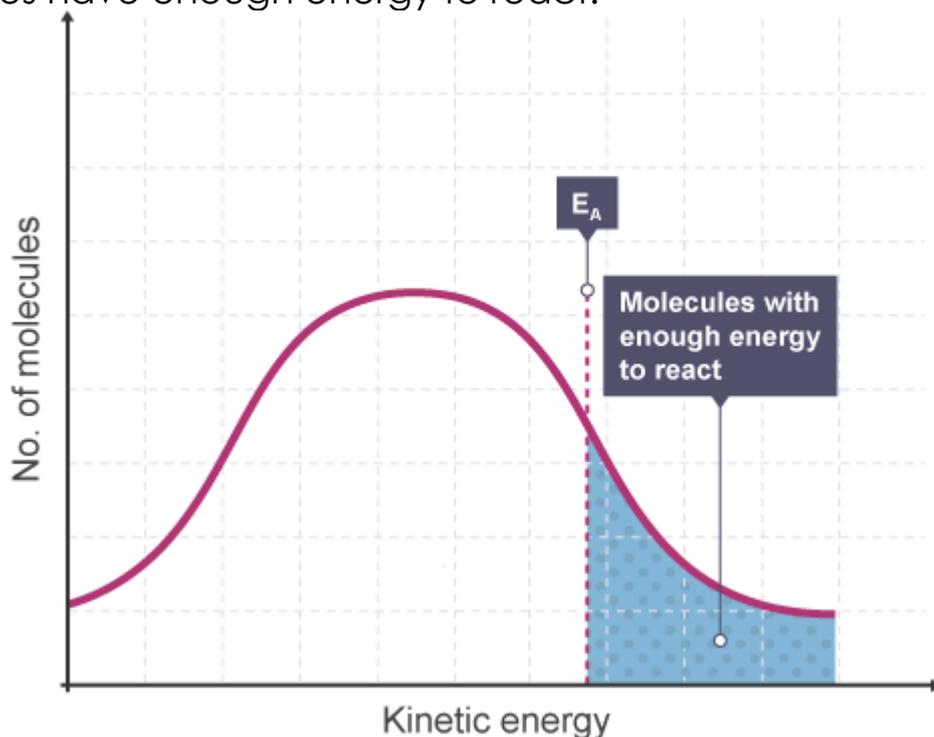
The activation energy is the minimum energy required for a reaction to occur. This means that the reactant molecules have enough kinetic energy to collide successfully and overcome the repulsion caused by outer electrons.

If the activation energy is high for a reaction, then only a few particles will have enough energy to collide so the reaction will be slow.

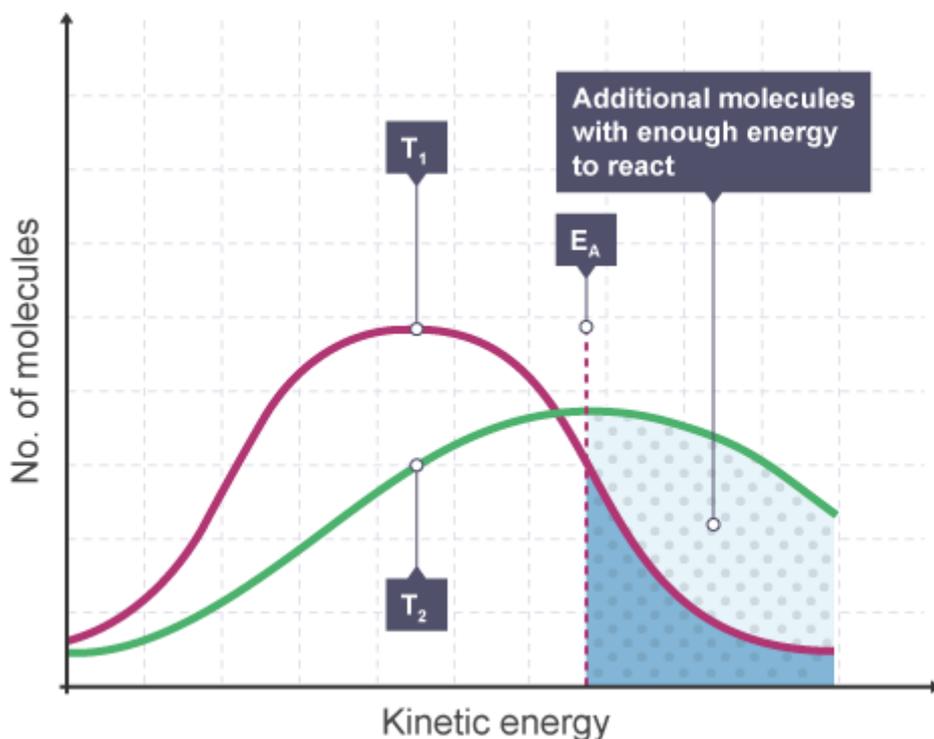
If a reaction has a low activation energy then the reaction will be fast as a lot of particles will have the required energy.



By showing the activation energy on a graph, we can see how many molecules have enough energy to react.



The effect of temperature, on a reaction, can be shown using these graphs.



Line T_2 shows a slight increase of temperature so causes a large increase in the number of molecules with kinetic energy (E_K) greater than the activation energy ($E_K > E_A$)

There is a significant increase in the rate of reaction. In fact, a 10°C rise in temperature results in the rate of reaction doubling.

Catalysts

A catalyst alters the rate of a reaction, allowing it to be done at a lower temperature. Catalysts are therefore used in the chemical industry to make manufacturing processes more economical.

Some examples of catalysts used in industry are:

- Iron – used to make ammonia by the Haber Process
- Platinum – used in manufacture of nitric acid (Ostwald Process)
- Rhodium and Platinum - in catalytic converters
- Nickel – to make margarine by hardening vegetable oil
- Vanadium (V) Oxide – in the contact process, to make sulphuric acid

e) Potential energy diagrams

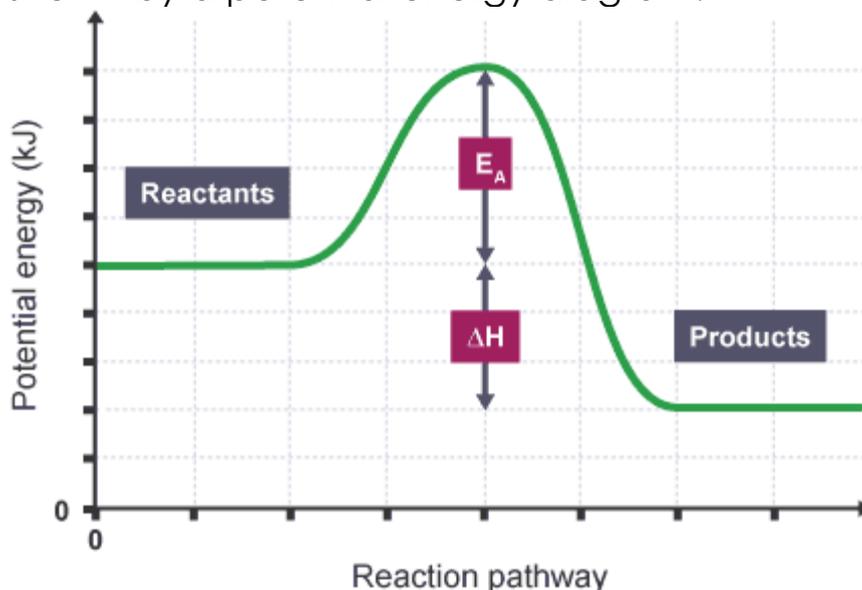
Chemical reactions involve a change in energy, usually a loss or gain of heat energy. The heat stored by a substance is called its enthalpy (H).

ΔH is the overall enthalpy change for a reaction. Potential energy diagrams can be used to calculate both the enthalpy change and the activation energy for a reaction.

Exothermic reactions

An exothermic reaction is one in which heat energy is given out. The products must have less energy than the reactants because energy has been released.

This can be shown by a potential energy diagram:



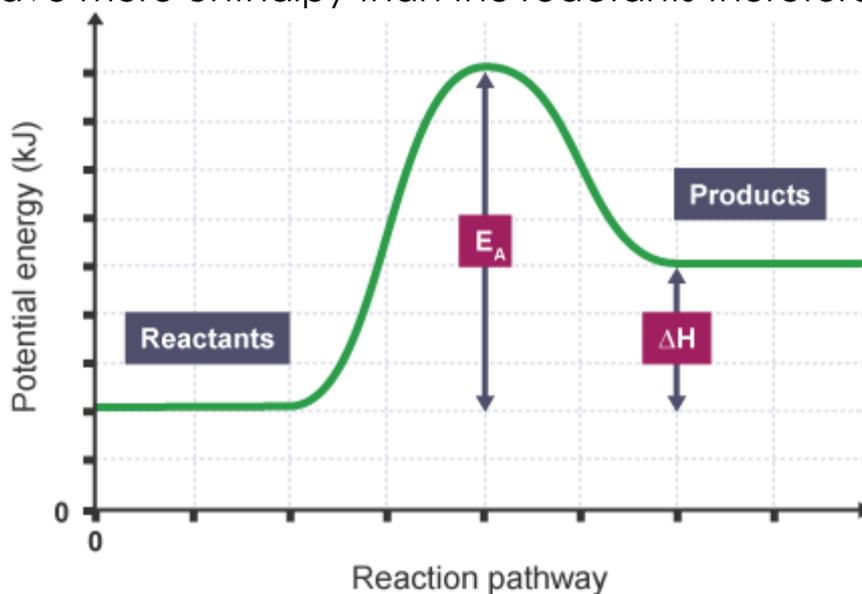
E_A is the activation energy (energy required to start the reaction)

ΔH is the quantity of energy given out (ie the enthalpy change)

For exothermic reactions ΔH will always be negative.

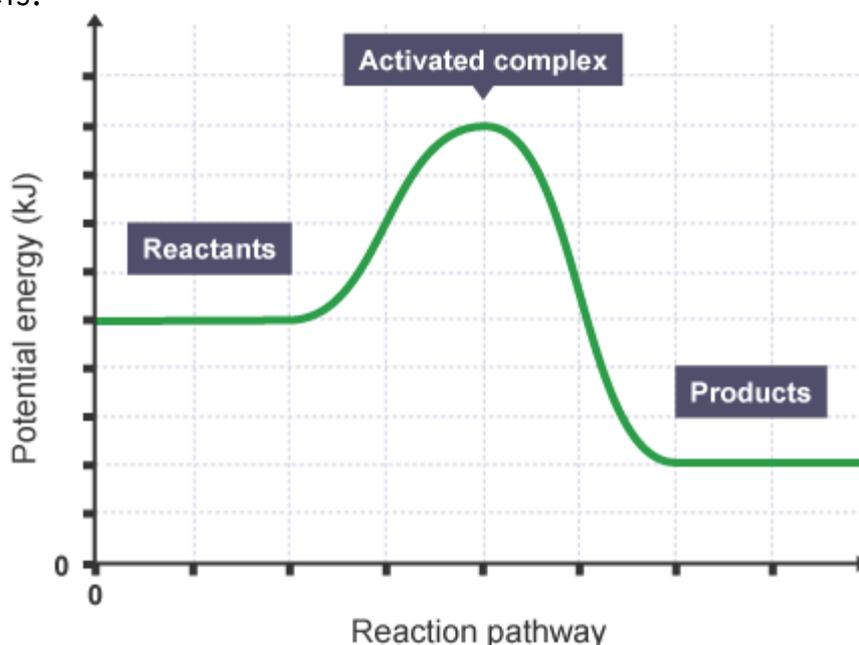
Endothermic reactions

An endothermic reaction is one in which heat energy is absorbed. The products have more enthalpy than the reactants therefore ΔH is positive.



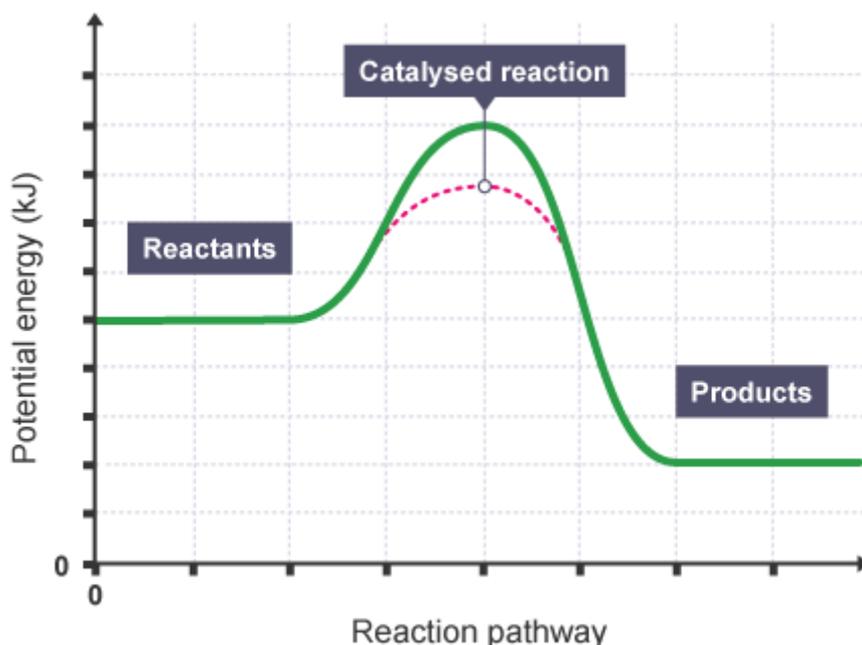
Activated complex

The activated complex (high energy intermediate state where bonds are breaking and forming) can be shown on potential energy diagrams. It is the 'energy barrier' that must be overcome when changing reactants into products.



Catalysts

A catalyst provides an alternative reaction pathway which involves less energy and so the catalyst lowers the activation energy.



The use of a catalyst does not affect the reactants or products, so ΔH stays the same.

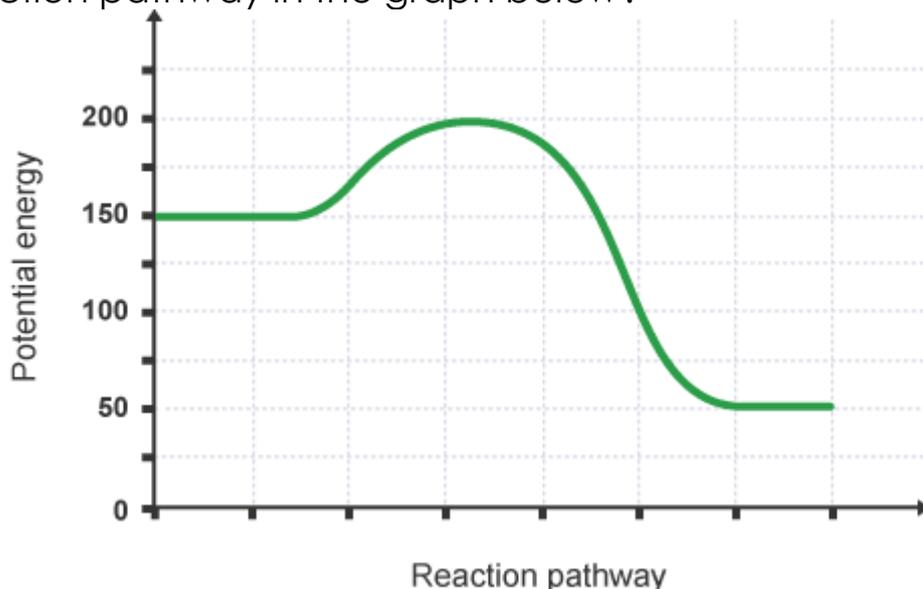
Concentration of solutions

Solutions are formed when solutes dissolve in solvents. If the number of moles of solute and the volume of solvent used is known, the concentration of the solution can be calculated.

The concentration of a solution is measured in moles per litre (mol l^{-1}) and can be calculated using this formula triangle:

Controlling The Rate Minitest

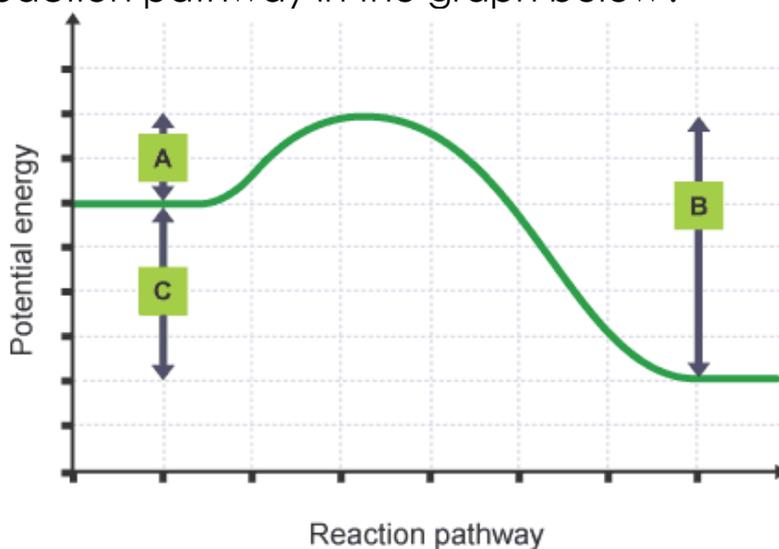
- Which of the following is the energy threshold that must be overcome in order for collisions to be successful?
 - Activated complex
 - Activation energy
 - Enthalpy change
- What is the enthalpy change for the forward reaction shown by the reaction pathway in the graph below?



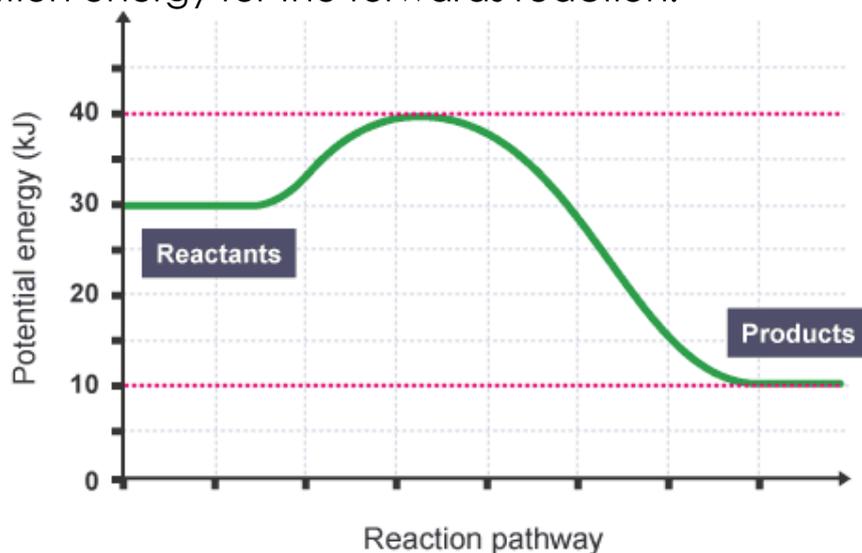
- 100 kJ mol⁻¹
 - +100 kJ mol⁻¹
 - +50 kJ mol⁻¹
- What effect would the use of a catalyst have on a chemical reaction?
 - Activation energy remains unchanged; enthalpy increases
 - Activation energy increases; enthalpy unchanged
 - Activation energy decreases; enthalpy unchanged
 - Which of the following factors would not increase the number of collisions between reactant molecules?
 - Increasing the particle size of the reaction
 - Increasing the temperature of the reaction
 - Increasing the concentration of the reaction

- 5 What is the main reason that a small increase in the temperature of a reaction mixture results in a large increase in the rate of the reaction?
- The activation energy is lowered
 - The enthalpy change is decreased
 - The kinetic energy of the particles has increased

- 6 What is the activation energy for the reverse reaction represented by the reaction pathway in the graph below?



- A
 - B
 - C
- 7 Using the reaction pathway shown in the graph below, calculate the activation energy for the forwards reaction.

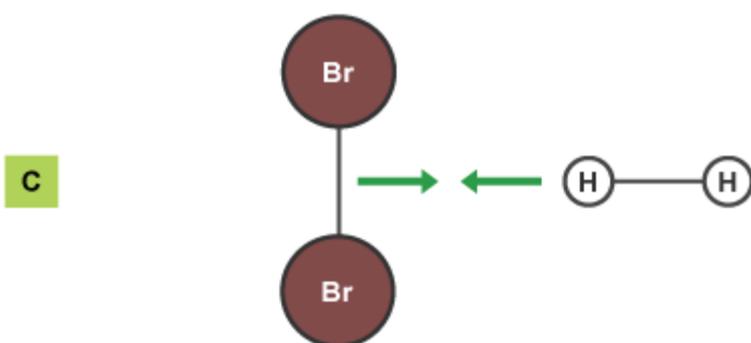
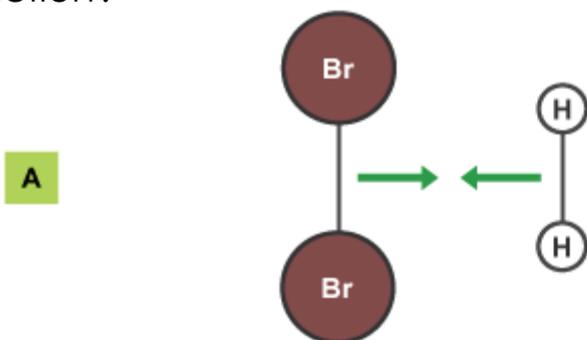


- 10 kJ mol⁻¹
- 20 kJ mol⁻¹
- 30 kJ mol⁻¹

8 What is the name given to a high energy intermediate state that is established when bonds inside the reactant molecules are breaking and new bonds are being formed?

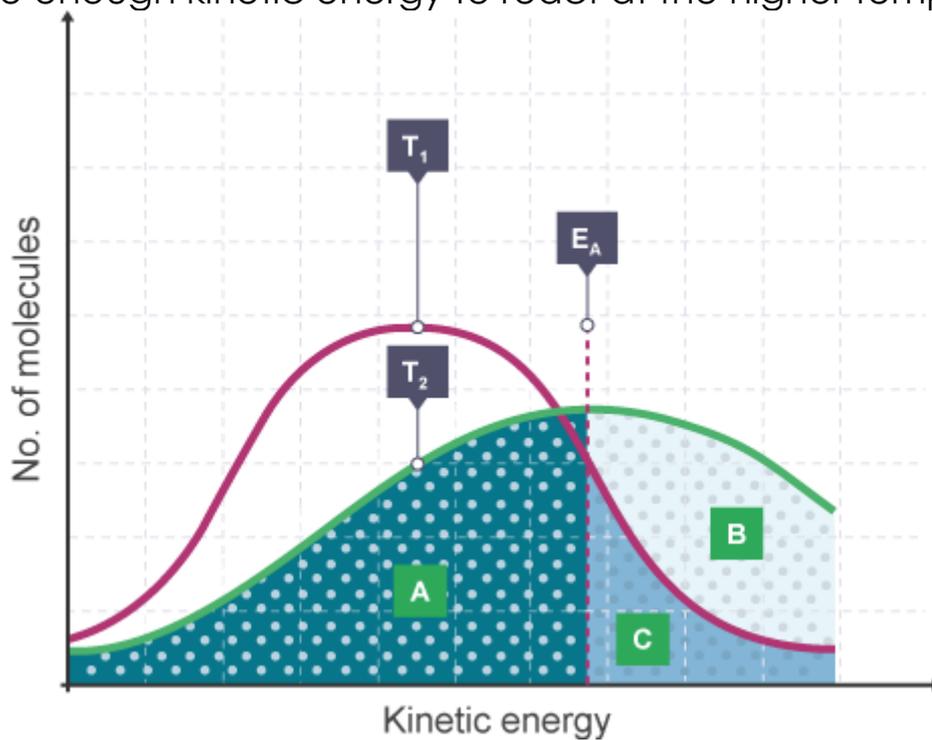
- Enthalpy change
- Exothermic reaction
- Activated complex

9 Which of the following collisions is most likely to result in a successful reaction?



- A
- B
- C

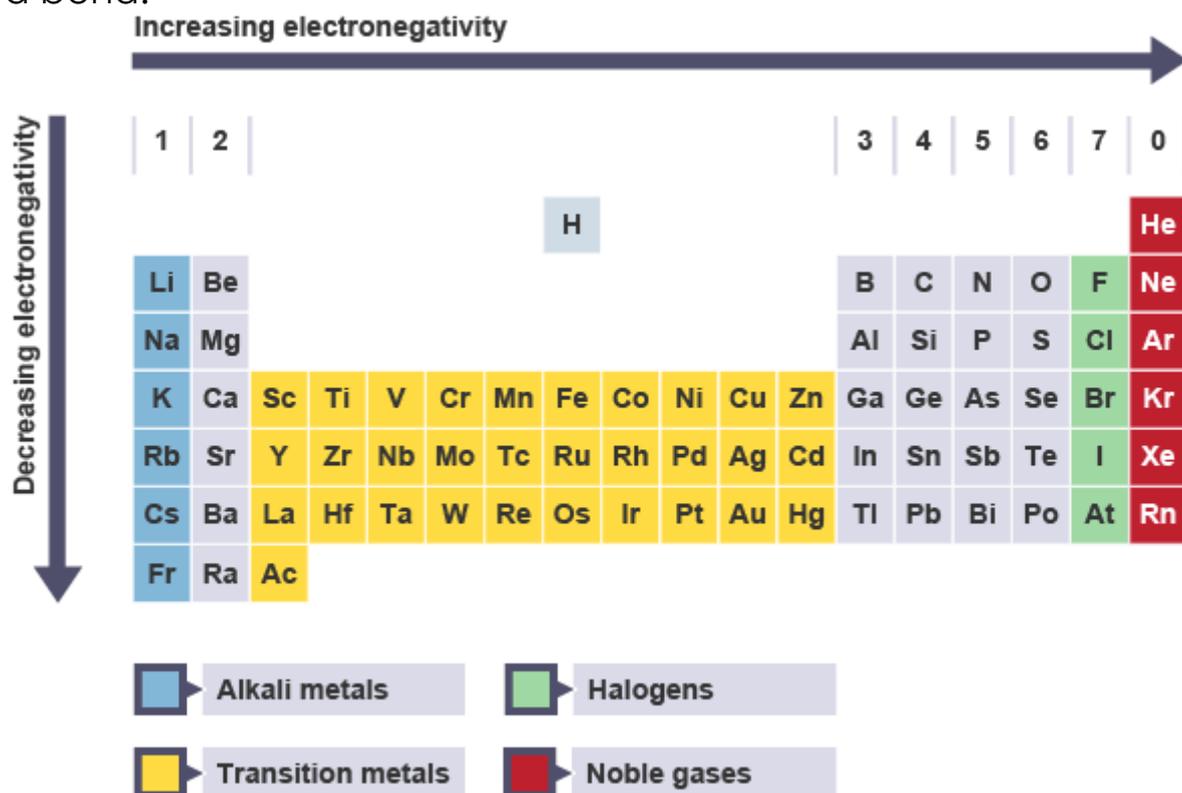
- 10 Which area(s) of the following graph represent the molecules that have enough kinetic energy to react at the higher temperature (T_2)?



- A, B + C
- B + C
- C

c) Electronegativity

Electronegativity is a measure of an atom's attraction for the electrons in a bond.



Across a period from left to right the electronegativity of atoms increases. As you move from left to right across the periodic table, atoms have a greater charge in their nucleus and a smaller covalent radius. This allows the nucleus to attract the bonding electrons more strongly.

Going down a group electronegativity decreases.

As you move down a group in the periodic table, atoms increase in size, with a greater number of energy levels.

The extra energy levels and increased covalent radius keep the bonding electrons further away from the nucleus.

This screening effect means that atoms further down groups have less attraction for the bonding electrons.

Both of these trends show that fluorine is highly electronegative (it pulls a shared pair of bonding electrons towards itself).

3 – Bonding in Elements

a) Metallic bonding

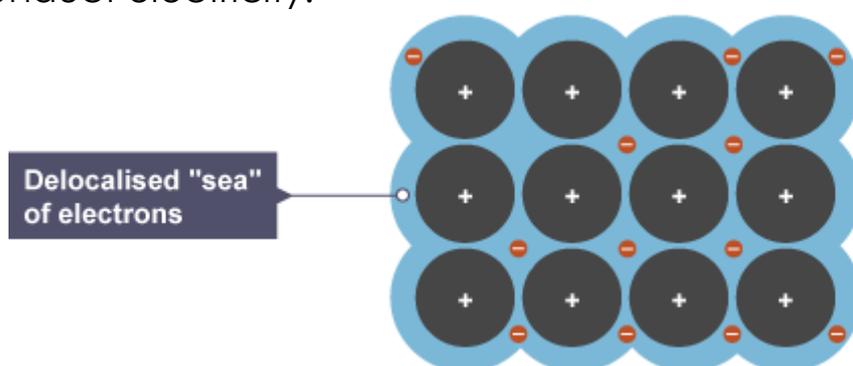
All the chemical elements are arranged in the periodic table in horizontal rows (periods) in order of increasing atomic number and also in vertical columns (groups). Elements in the same group have similar reactivities. This allows chemists to make predictions about the reactivity or type of bonding that elements have. Within the first 20 elements there are various different types of bonding displayed.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

Metallic bonding occurs between the atoms of metal elements. The outer electrons are delocalised (free to move).

This produces an electrostatic force of attraction between the positive metal ions and the negative delocalised electrons.

This delocalised 'sea of electrons' is responsible for metal elements being able to conduct electricity.



b) Covalent molecules

Discrete covalent molecules are small groups of atoms held together by strong covalent bonds inside the molecule and weak intermolecular forces between the molecules.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

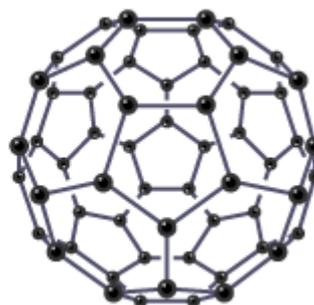
The covalent bond itself is a shared pair of electrons electrostatically attracted to the positive nuclei of two non-metal atoms. The atoms achieve a stable outer electron arrangement (a noble gas arrangement) by sharing electrons.

Most of the discrete covalent molecules are diatomic elements:

- Hydrogen (H_2)
- Nitrogen (N_2)
- Oxygen (O_2)
- Fluorine (F_2)
- Chlorine (Cl_2)

There are also some larger covalent molecular elements:

- Phosphorous (P_4)
- Sulfur (S_8)
- Fullerenes (C_{60})

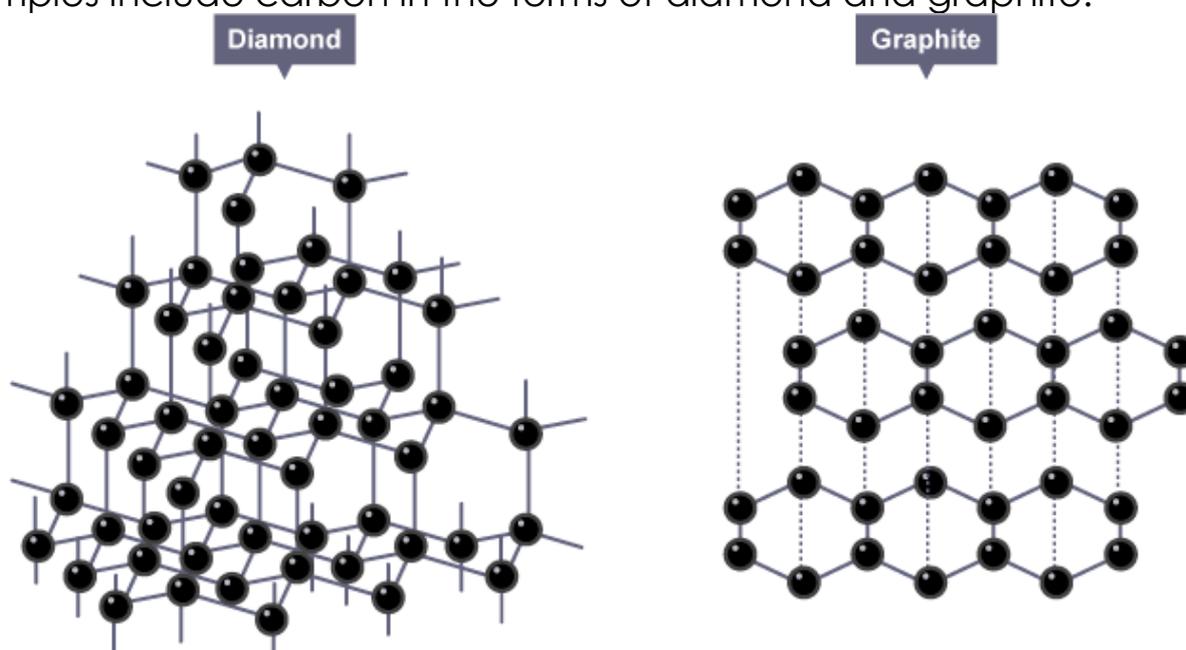


c) Covalent network

Covalent networks are large, rigid three-dimensional arrangements of atoms held together by strong covalent bonds.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

They have high melting points because they only contain strong bonds. Examples include carbon in the forms of diamond and graphite.



d) Monatomic elements

Group 0 elements (the noble gases) exist as single, unattached particles. They are stable atoms. They have fuller outer energy levels so they do not usually form molecules with other atoms.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

They have low melting and boiling points as they are easily separated by overcoming the weak forces of attraction between the atoms.

e) Summary of bonding

Structure	Density	Melting Point	Conduction
Monatomic (noble gas)	Low	Low	Non-conductor
Covalent molecular (N ₂ , P ₄ , S ₈)	Low	Low	Non-conductor
Covalent network	Very high	Very high	Non-conductor (except graphite)
Metallic lattice (all metals)	High	High	Conductor

Periodicity & Bonding in Elements Minitest

- Which type of bonding holds aluminium atoms together in a sample of the element?
 - Covalent molecular
 - Metallic
 - Covalent network
- Which of the following elements has a monatomic structure?
 - Nitrogen
 - Neon
 - Carbon
- Which of these elements has a covalent molecular structure?
 - Helium
 - Sodium
 - Phosphorus
- Which of the following is not a property of all covalent networks?
 - Strong, rigid molecules
 - Very high melting and boiling points
 - Conduct electricity
- Which of the following is the correct equation for the first ionisation energy of chlorine?
 - $Cl(g) \rightarrow Cl^+(g) + e^-$
 - $Cl(g) + e^- \rightarrow Cl^-(g)$
 - $Cl(aq) + e^- \rightarrow Cl^-(aq)$
- Which of the following is the correct equation for the second ionisation energy of magnesium?
 - $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$
 - $Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$
 - $Mg(g) \rightarrow Mg^+(g) + e^-$
- Which of the following describes a property of an atom with a high electronegativity?
 - It has a weak attraction for the electrons in a covalent bond
 - It has a strong attraction for the electrons in a covalent bond
 - It has no attraction for the electrons in a covalent bond
- Why does fluorine have a lower covalent radius than lithium?
 - Fluorine has more protons inside its nucleus
 - Lithium has a greater number of energy levels
 - Fluorine has a greater electronegativity
- Why is there such a large increase between the second and third ionisation energies of calcium?
 - The third electron must be removed from a full energy level
 - Calcium atoms have a large nuclear charge
 - There is a shielding effect
- Going down a group in the periodic table, what trend is observed in covalent radius?
 - The covalent radius decreases
 - The covalent radius increases
 - There is no pattern or trend

4 – Bonding and Structure

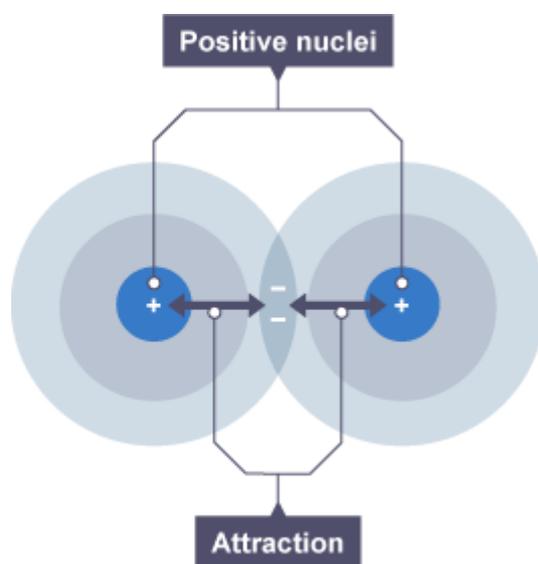
a) Intramolecular bonding

There are several types of bonding inside molecules (intramolecular bonds) and between molecules (intermolecular bonds).

The intramolecular bonding types have different properties, but all can be arranged into a bonding continuum, where the bonding present inside molecules has varying degrees of ionic character.

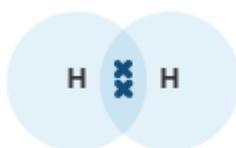
Pure covalent bonds

A covalent bond is a shared pair of electrons, electrostatically attracted to the positive nuclei of two atoms.



Atoms can share electrons in order to achieve a stable outer electron arrangement (a noble gas arrangement).

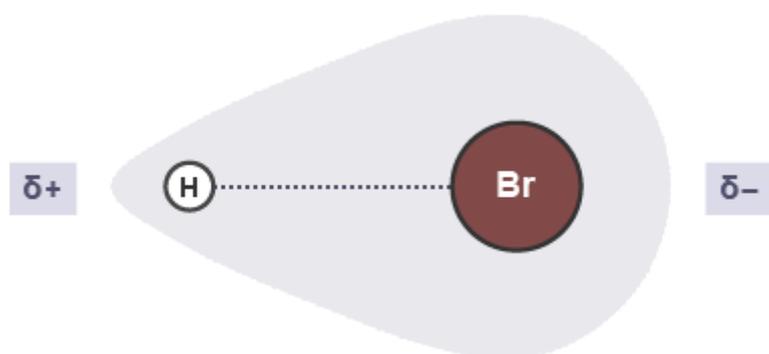
Pure covalent bonds exist between two atoms with the same electronegativities. A pure covalent bond has no ionic character at all.



Diatomic elements are good examples of pure covalent bonds where the electrons are evenly shared between both nuclei.

Polar covalent bonds

It is unusual for pure covalent bonds to exist between atoms of different elements. Usually, one of the atoms involved in the covalent bond will be more electronegative and will have a greater attraction for the bonding pair of electrons. This gives rise to polar covalent bonding.



A polar covalent bond is a bond formed when a shared pair of electrons is not shared equally. This is due to one of the elements having a higher electronegativity than the other.

The shared pair of electrons between an atom of hydrogen and an atom of bromine is not shared equally. Since bromine has a greater electronegativity than hydrogen, it will pull the bonding electrons towards itself.

This makes bromine slightly negative ($\delta-$) and hydrogen slightly positive ($\delta+$). This is known as a dipole.

The shaded area shows where the bonding electrons are likely to be found.

Ionic bonds

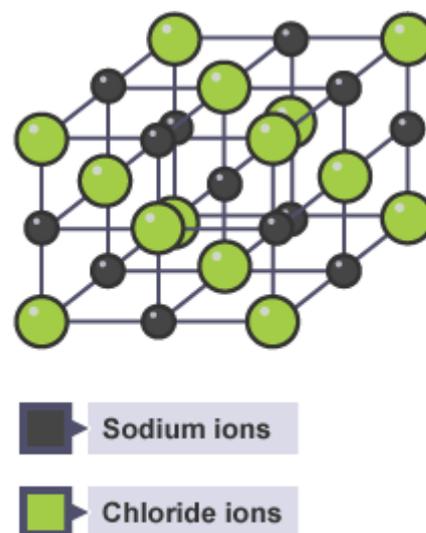
While there is an uneven sharing of bonding electrons in polar covalent bonds, an even larger difference in electronegativity gives rise to ionic bonds.

Ionic bonds are usually (but not always) formed between a metal and non-metal with a large difference in electronegativity, eg Sodium Chloride.

Chlorine has a far larger electronegativity and so pulls the bonding electrons towards itself completely, thus gaining an electron and forming a negative ion.

Sodium, due to its low electronegativity, will lose its bonding electron to chlorine and form a positive ion.

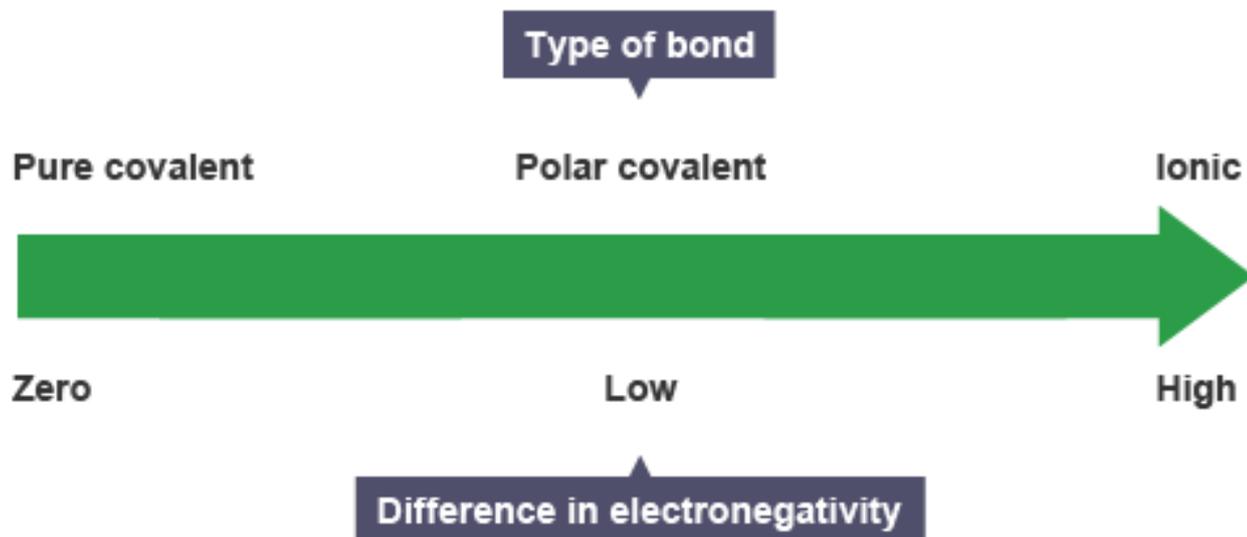
The ionic bond is the electrostatic force of attraction between a metal ion and a non-metal ion. Ions are arranged into a 3-dimensional ionic lattice of positive and negative ions.



Bonding continuum

Pure covalent bonds, polar covalent bonds and ionic bonds all exist as part of the same bonding continuum.

The type of bonding present in a compound is determined by the differences in electronegativity between the elements involved.



b) Intermolecular bonds

Intermolecular bonds are found between molecules. They are also known as Van der Waals forces, and there are several types to consider.

London dispersion forces

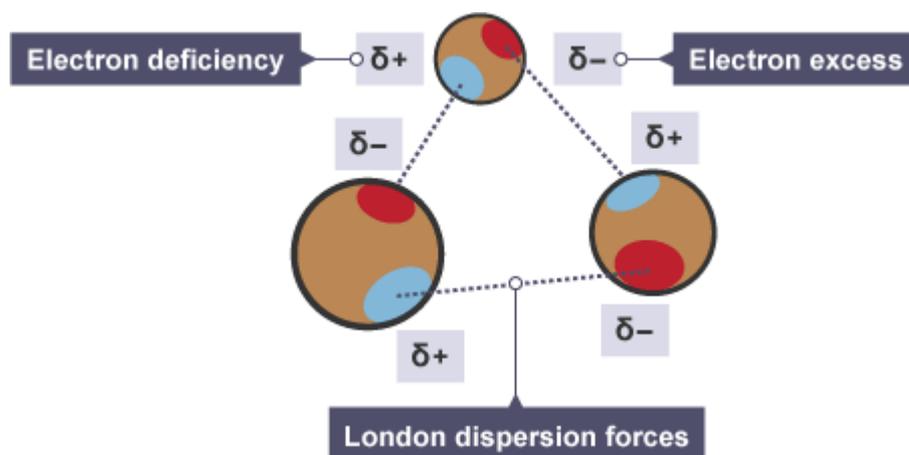
London dispersion forces (or London forces) are the weakest type of intermolecular bond. They exist between all atoms and molecules.

Molecular elements (oxygen, nitrogen etc) and monatomic elements (the noble gases) will condense (move closer together) forming solids if cooled to sufficiently low temperatures.

This shows that there must be an attraction between the individual molecules (or atoms in the case of monatomic substances) that is being overcome.

London forces are caused by an uneven distribution of electrons within an atom. This results in a slightly negative charge (δ^-) and slightly positive charge (δ^+) on either side of the atom. A temporary dipole has been established.

The opposite charges then attract each other.



The strength of London dispersion forces depends on the size of the molecule or atom.

Larger atoms and molecules have more electrons. This leads to larger dipoles being established.

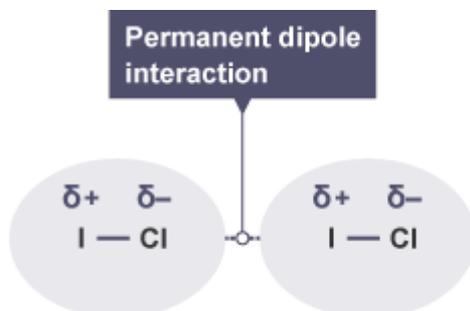
London dispersion forces increase the larger the atomic size.

Permanent dipole interactions

Molecules with a permanent dipole are polar.

Polar molecules display attractions between the oppositely charged ends of the molecules.

This type of intermolecular bond is stronger than London dispersion forces (which can be called temporary dipoles).



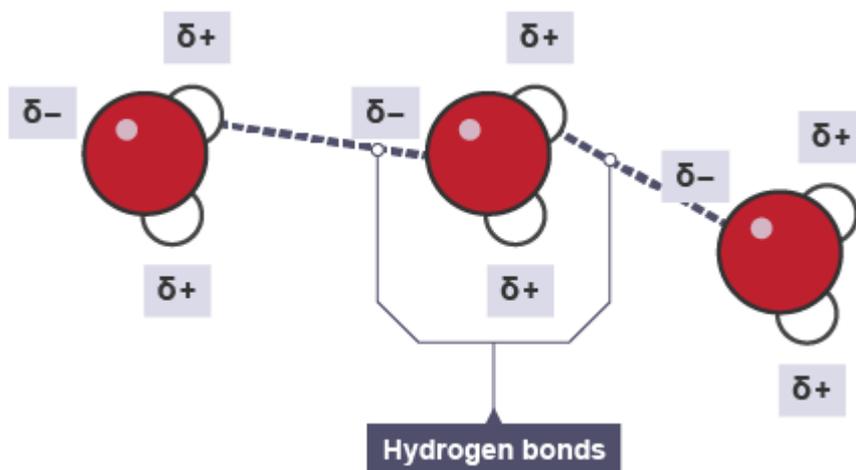
Hydrogen bonding

Hydrogen bonding is the strongest type of intermolecular bond.

It is a specific type of permanent dipole to permanent dipole attraction that occurs when a hydrogen atom is covalently bonded to a highly electronegative element such as nitrogen, oxygen or fluorine.

As with permanent dipole to permanent dipole attractions, the oppositely charged ends of molecules attract.

In the diagram below, the hydrogen bonds are shown as the $\delta+$ hydrogen atoms of one molecule are attracted to the $\delta-$ oxygen atoms of another.



Water, ammonia, alcohols and alkanolic acids all contain hydrogen bonding.

Bonding strength

The relative strengths of bonds are:

Covalent bonds > Hydrogen bonds > Permanent dipole interactions > London dispersion forces.

This can be shown by comparing two molecules with the same numbers of electrons.

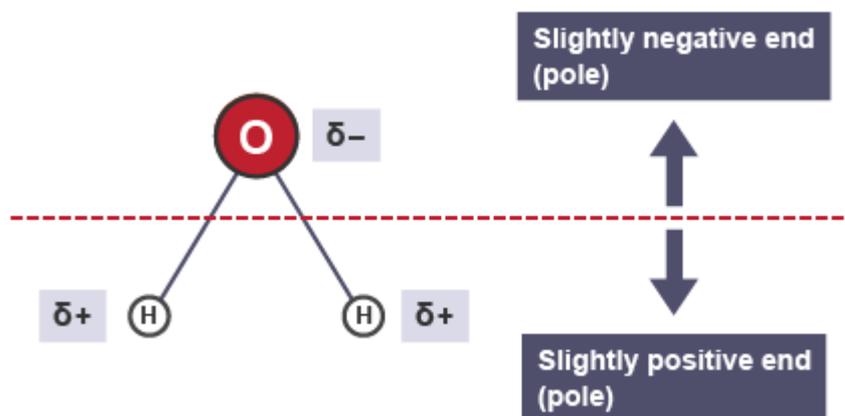
Bromine atoms contain 35 electrons. Bromine exists as a diatomic molecule (Br_2) and has 70 electrons in total.

Bromine molecules contain pure covalent bonds so are held together by London dispersion forces (temporary dipoles caused by momentary uneven sharing of electrons. Bromine is a liquid at room temperature (melting point -7°C).

Iodine monochloride (I-Cl) is also a diatomic molecule, but contains a polar covalent bond. Since this molecule is a dipole it displays permanent dipole to permanent dipole interactions. It exists as a solid at room temperature (melting point 27°C).

Polar and non-polar molecules

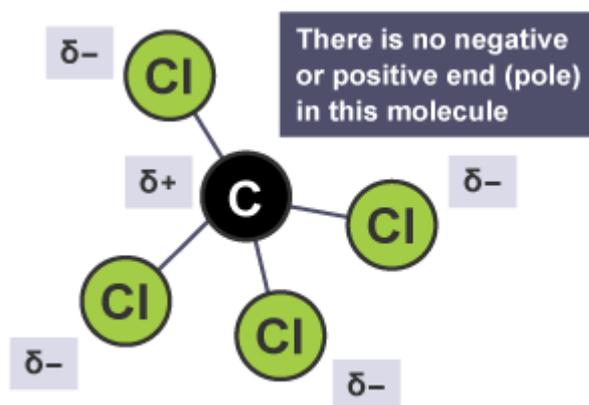
A substance that contains polar covalent bonds may not be overall polar. This is due to the shape of the molecule.



Water molecules are polar molecules. Both of the bonds inside the molecule are polar bonds.

Due to the non-symmetrical shape of the molecule (bent), the molecule itself is polar (it has a δ^- side and δ^+ side).

Carbon tetrachloride has four polar covalent bonds. It is a tetrahedral molecule, so the charges are symmetrical.



The $\delta+$ part of the molecule (the carbon atom) is locked away at the centre and no matter what direction you approach the molecule from you see a $\delta-$ charge.

This means that the molecule itself is non-polar due to symmetry.

c) Properties relating to bonding

The properties of elements and compounds can be explained by considering the different types of bonding present inside the molecules.

Ionic lattice

All ionic compounds have a high melting point and boiling point. They conduct when molten or in solution as the ions are free to move.

They can be broken down by electrolysis.

Covalent network

All covalent network structures have very high melting points and boiling points. They are all hard and do not conduct electricity.

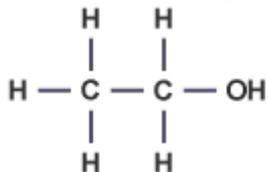
Covalent molecular

They have low melting points and boiling points. They do not conduct electricity. Some covalent molecular compounds have higher melting points than expected.

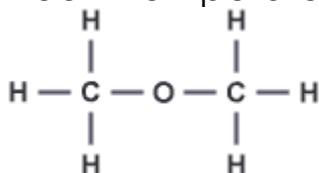
Sometimes two molecules that seem similar can have very different properties because of the intermolecular bonds present. Consider ethanol and ether.

Both molecules have the same molecular formula (C_2H_6O) yet there is a large difference in boiling points.

Ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) is a liquid at room temperature (boiling point 79°C).



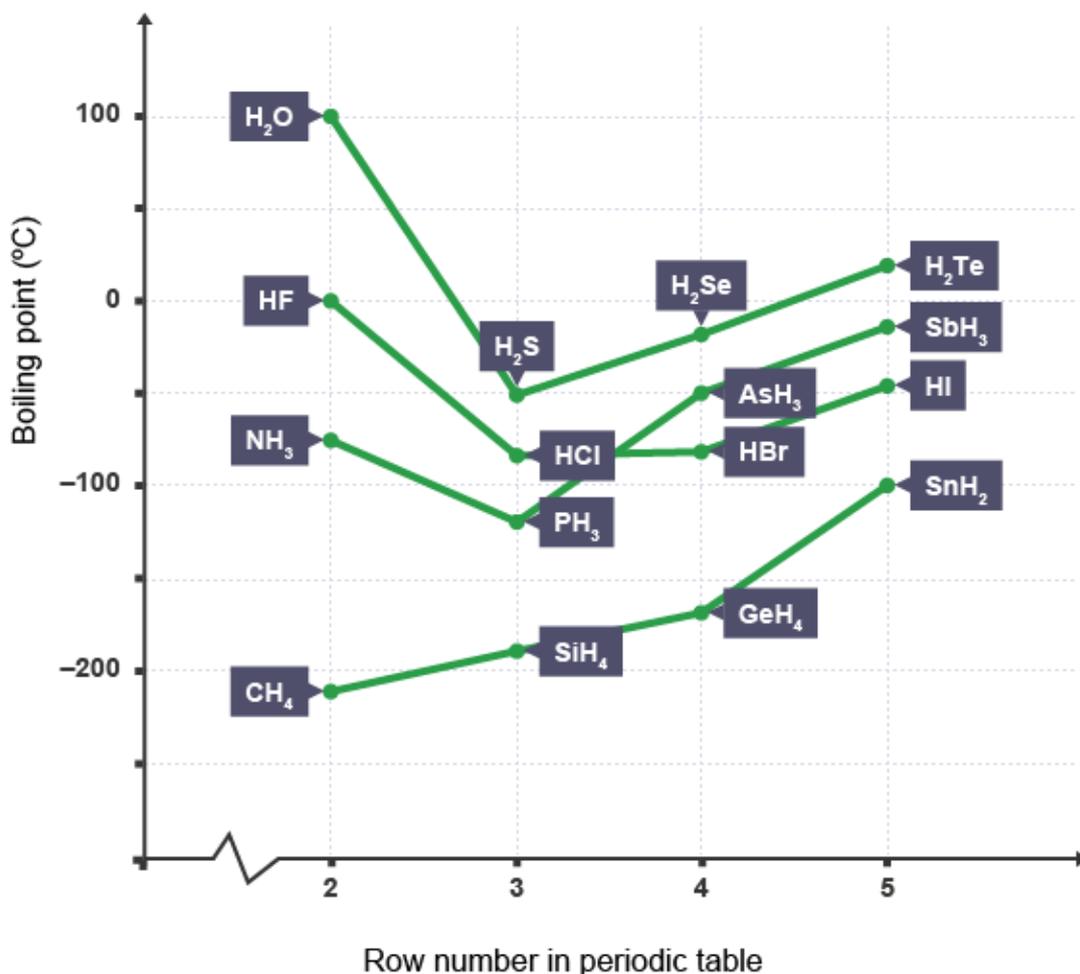
Ether (CH_3OCH_3) is a gas at room temperature (boiling point -23°C).



Both ethanol and ether have the same molecular mass, however, their melting points and boiling points are different due to the ethanol containing hydrogen bonding.

Ether is a symmetrical molecule, so even though it contains polar bonds, the molecule itself is non-polar and the only intermolecular force present will be weak London dispersion forces.

Ammonia (NH_3) and hydrogen fluoride (HF) both also have higher boiling points than might be predicted due to presence of hydrogen bonding between the molecules.



The properties of water

Water has some unusual properties due to the hydrogen bonding between its molecules.

Properties:

- The density of ice is less than water. This is due to the water expanding as it is frozen because of the hydrogen forming an open type lattice.
- high surface tension
- relatively high viscosity



Summary of Solubilities

All ionic lattices and polar covalent molecules are:

- soluble in water and other polar solvents
- insoluble in non-polar solutions (eg hexane)

All non-polar covalent substances are:

- soluble in non-polar solvents
- insoluble in water and other polar solvents

As a general rule "like dissolves like" so polar molecules are soluble in polar solvents (such as water) and vice-versa.



Structure and bonding Minitest

- Which type of bonding results from the electrostatic force of attraction between positively and negatively charged ions?
 - Metallic bonding
 - Polar covalent bonding
 - Ionic bonding
- Which of these bonding types has the least ionic character in the bonding continuum?
 - Pure covalent
 - Polar covalent
 - Ionic bonding
- Which of the following is most likely to contain ionic bonding?
 - Solid at room temperature, low melting point, does not conduct electricity
 - Solid at room temperature, high melting point, does not conduct electricity
 - Solid at room temperature, high melting point, conducts electricity when in solution
- Which of the following is the strongest type of intermolecular bond?
 - London dispersion forces
 - Permanent dipole to dipole attractions
 - Hydrogen bonding
- Which intermolecular bond is caused by the temporary uneven distribution of electrons?
 - London dispersion forces
 - Permanent dipole to dipole attractions
 - Hydrogen bonding
- Which of the following molecules are polar?
 - Carbon tetrachloride (CCl_4)
 - Carbon dioxide (CO_2)
 - Hydrogen fluoride (HF)
- Which of the following solvents is most likely to dissolve an ionic lattice of sodium bromide?
 - Hexane
 - Cyclohexane
 - Water
- Which of the following properties of water is not related to hydrogen bonding?
 - Liquid water is a poor conductor of electricity
 - Solid water (ice) is less dense than liquid water
 - Liquid water has a high surface tension
- Why is methanol a liquid at room temperature while methane (CH_4) is a gas?
 - Methane is a smaller, less viscous molecule
 - Methanol contains hydrogen bonding
 - Methanol contains only pure covalent bonding

Term	Meaning
Activated complex	the activated complex is a very unstable arrangement of atoms formed at the maximum of the potential energy barrier, during a chemical reaction
Activation energy	is the minimum kinetic energy required by colliding particles before reaction will occur, since a high energy activated complex must be formed
Adsorption	occurs when molecules become bonded to the surface of a catalyst
Allotrope	one of two or more existing forms of an element. For example, diamond, fullerene and graphite are allotropes of carbon
Bonding electrons	are shared pairs of electrons from both atoms forming the covalent bond
Chemical bonding	is the term used to describe the mechanism by which atoms are held together
Chemical structure	describes the way in which atoms, ions or molecules are arranged
Collision theory	of reactions suggests that, for a chemical reaction to occur, particles must collide
Covalent bond	formed when two atoms share electrons in their outer shell to complete the filling of that shell
Covalent radius	half the distance between the nuclei of two bonded atoms of an element
Delocalised	electrons, in metallic bonding, are free from attachment to any one metal ion and are shared amongst the entire structure
Desorption	occurs when the bonds between the molecules and the surface break and the molecules leave the surface of the catalyst
Diatomic	molecules with only two atoms are described as diatomic (e.g. oxygen, O ₂ , and carbon monoxide, CO.)
Dipole	an atom or molecule in which a concentration of positive charges is separated from a concentration of negative charge
Electronegativity	a measure of the attraction that an atom involved in a bond has for the electrons of the bond
Enthalpy change	for a reaction is defined as the change in heat energy when 1 mole of reactant is converted to product(s) at <i>constant pressure</i> , and has the symbol ΔH and units of kJ mol ⁻¹

Fullerenes	are molecules of pure carbon constructed from 5- and 6-membered rings combined into hollow structures. The most stable contains 60 carbon atoms in a shape resembling a football
Hydrogen bonds	are electrostatic forces of attraction between molecules containing a hydrogen atom bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen, and a highly electronegative atom on a neighbouring molecule
Intermolecular forces	are those which attract molecules together. They are weaker than chemical bonds
Intramolecular forces	are forces of attraction which exist within a molecule
Ionisation energy	the energy required to remove one mole of electrons from one mole of atoms in the gaseous state
Isoelectronic	means having the same arrangement of electrons. For example, the noble gas neon, a sodium ion (Na ⁺) and a magnesium ion (Mg ²⁺) are isoelectronic
Lattice	a lattice is a regular 3D arrangement of particles in space. The term is applied to metal ions in a solid, and to positive and negative ions in an ionic solid
London dispersion forces	the forces of attraction which result from the electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules
Lone pairs	are pairs of electrons in the outer shell of an atom which take no part in bonding
Miscible	fluids are fluids which mix with or dissolve in each other in all proportions
Periodicity	is the regular recurrence of similar properties when the elements are arranged in order of increasing atomic number
Polar covalent bond	a covalent bond between atoms of different electronegativity, which results in an uneven distribution of electrons and a partial charge along the bond
Potential energy diagram	shows the enthalpy of reactants and products, and the enthalpy change during a chemical reaction
Properties	of a substance are their physical and chemical characteristics. These are often a reflection of the chemical bonding and structure of the material.
Thermochemical equation	states the enthalpy change for the reaction defined, with reactants and products in the states shown
Viscosity	is the resistance to flow that is exhibited by all liquids