Chapter 1: Atomic structure

Knowledge organiser

Development of the model of the atom

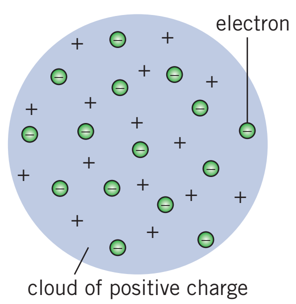
Dalton's model

John Dalton thought of the **atom** as a solid sphere that could not be divided into smaller parts. His model did not include **protons**, **neutrons**, or **electrons**.

The plum pudding model

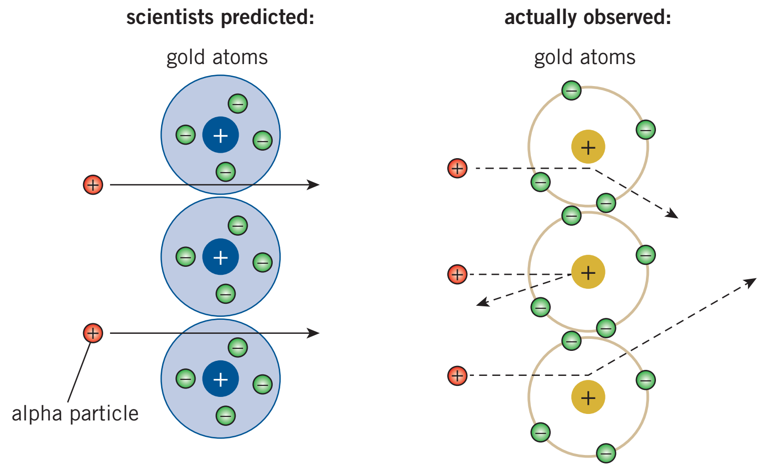
Scientists' experiments resulted in the discovery of sub-atomic charged particles. The first to be discovered were electrons – tiny, negatively charged particles.

The discovery of electrons led to the plum pudding model of the atom – a cloud of positive charge, with negative electrons embedded in it. Protons and neutrons had not yet been discovered.



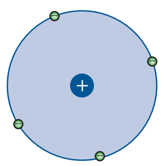
Alpha scattering experiment

1. Scientists fired small, positively charged particles (called alpha particles) at a piece of gold foil only a few atoms thick.
2. They expected the alpha particles to travel straight through the gold.
3. They were surprised that some of the alpha particles bounced back and many were deflected (alpha scattering).
4. To explain why the alpha particles were repelled the scientists suggested that the positive charge and mass of an atom must be concentrated in a small space at its centre. They called this space the **nucleus**.



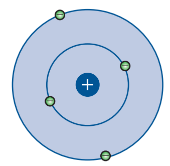
Nuclear model

Scientists replaced the plum pudding model with the nuclear model and suggested that the electrons **orbit** the nucleus, but not at set distances.



Electron shell (Bohr) model

Niels Bohr calculated that electrons must orbit the nucleus at fixed distances. These orbits are called **shells** or **energy levels**.



The proton

Further experiments provided evidence that the nucleus contained smaller particles called protons. A proton has an opposite charge to an electron.

The neutron

James Chadwick carried out experiments that gave evidence for a particle with no charge. Scientists called this the neutron and concluded that the protons and neutrons are in the nucleus, and the electrons orbit the nucleus in shells.

Relative mass

One property of protons, neutrons, and electrons is **relative mass** – their masses compared to each other. Protons and neutrons have the same mass, so are given a relative mass of 1. It takes almost 2000 electrons to equal the mass of a single proton – their relative mass is so small that we can consider it as 0.

Size

The atom has a radius of 1×10^{-10} m. Nuclei (plural of nucleus) are around 10 000 times smaller than atoms and have a radius of around 1×10^{-14} m.

Elements and compounds

Elements are substances made of one type of atom. Each atom of an element will have the same number of protons.

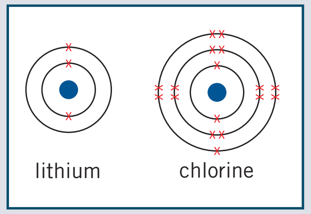
Compounds are made of different types of atoms chemically bonded together. The atoms in a compound have different numbers of protons.

Drawing atoms

Electrons in an atom are placed in fixed shells. You can put

- up to two electrons in the first shell
- eight electrons each in the second and third shells.

You must fill up a shell before moving on to the next one.



Mixtures

- A mixture consists of two or more elements or compounds that are not chemically combined together.
- The substances in a mixture can be separated using physical processes.
- These processes do not use chemical reactions.

Separating mixtures

- filtration – insoluble solids and a liquid
- crystallisation – soluble solid from a solution
- simple distillation – solvent from a solution
- fractional distillation – two liquids with similar boiling points
- paper chromatography – identify substances from a mixture in solution

Atoms and particles

	Relative charge	Relative mass	
Proton	+1	1	= atomic number
Neutron	0	1	= mass number – atomic number
Electron	–1	0 (very small)	= same as the number of protons

All atoms have equal numbers of protons and electrons, meaning they have no overall charge:

total negative charge from electrons = total positive charge from protons

Isotopes

Atoms of the same element can have a different number of neutrons, giving them a different overall mass number. Atoms of the same element with different numbers of neutrons are called **isotopes**.

The **relative atomic mass** is the average mass of all the atoms of an element:

$$\text{relative atomic mass} = \frac{(\text{abundance of isotope 1} \times \text{mass of isotope 1}) + (\text{abundance of isotope 2} \times \text{mass of isotope 2})...}{100}$$

Key terms

Make sure you can write a definition for these key terms.

abundance, atom, atomic number, aqueous, compound, electron, element, energy level, isotope, neutron, nucleus, orbit, product, proton, reactant, relative atomic mass, relative charge, relative mass, shell

Chapter 1: Atomic structure

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C1 questions	Answers
1	What is an atom?	smallest part of an element that can exist
2	What is Dalton's model of the atom?	atoms as solid spheres that could not be divided into smaller parts
3	What is the plum pudding model of the atom?	sphere of positive charge with negative electrons embedded in it
4	What did scientists discover in the alpha scattering experiment?	some alpha particles were deflected by the gold foil – this showed that an atom's mass and positive charge must be concentrated in one small space (the nucleus)
5	Describe the nuclear model of the atom.	dense nucleus with electrons orbiting it
6	What did Niels Bohr discover?	electrons orbit in fixed energy levels (shells)
7	What did James Chadwick discover?	uncharged particle called the neutron
8	Where are protons and neutrons?	in the nucleus
9	What is the relative mass of each sub-atomic particle?	proton: 1, neutron: 1, electron: 0 (very small)
10	What is the relative charge of each sub-atomic particle?	proton: +1, neutron: 0, electron: −1
11	How can you find out the number of protons in an atom?	the atomic number on the Periodic Table
12	How can you calculate the number of neutrons in an atom?	mass number – atomic number
13	Why do atoms have no overall charge?	equal numbers of positive protons and negative electrons
14	How many electrons would you place in the first, second, and third shells?	up to 2 in the first shell and up to 8 in the second and third shells
15	What is an element?	substance made of one type of atom
16	What is a compound?	substance made of more than one type of atom chemically joined together
17	What is a mixture?	two or more substances not chemically combined
18	What are isotopes?	atoms of the same element (same number of protons) with different numbers of neutrons
19	What are the four physical processes that can be used to separate mixtures?	filtration, crystallisation, distillation, fractional distillation, chromatography
20	What is relative mass?	the average mass of all the atoms of an element

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Chapter 2: The Periodic Table

Knowledge organiser

Development of the Periodic Table

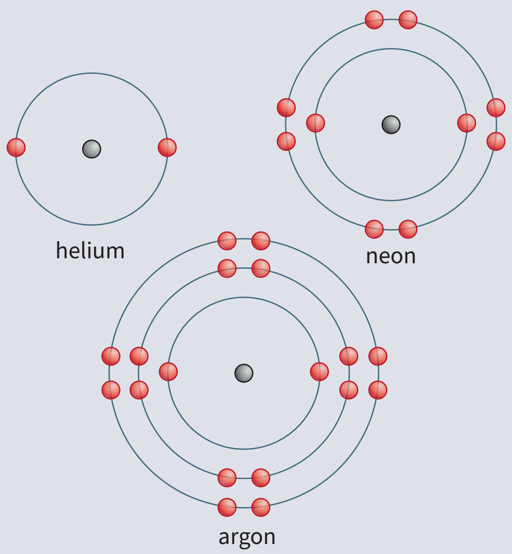
The Periodic Table has changed over time as scientists have organised it differently. Mendeleev was able to accurately predict the properties of undiscovered elements based on the gaps in the table.

	First lists of elements	Mendeleev's Periodic Table	Modern Periodic Table
How are elements ordered?	by atomic mass	normally by atomic mass but some elements were swapped around	by atomic number
Are there gaps?	no gaps	gaps left for undiscovered elements	no gaps – all elements up to a certain atomic number have been discovered
How are elements grouped?	not grouped	grouped by chemical properties	grouped by the number of electrons in the outer shells
Metals and non-metals	no clear distinction	no clear distinction	metals to the left, non-metals to the right
Problems	some elements grouped inappropriately	incomplete, with no explanation for why some elements had to be swapped to fit in the appropriate groups	—

Group 0

Elements in **Group 0** are called the **noble gases**. Noble gases have the following properties:

- full outer shells with eight electrons, so do not need to lose or gain electrons
- are very unreactive (**inert**) so exist as single atoms as they do not bond to form molecules
- boiling points that increase down the group.



Key terms Make sure you can write a definition for these key terms.

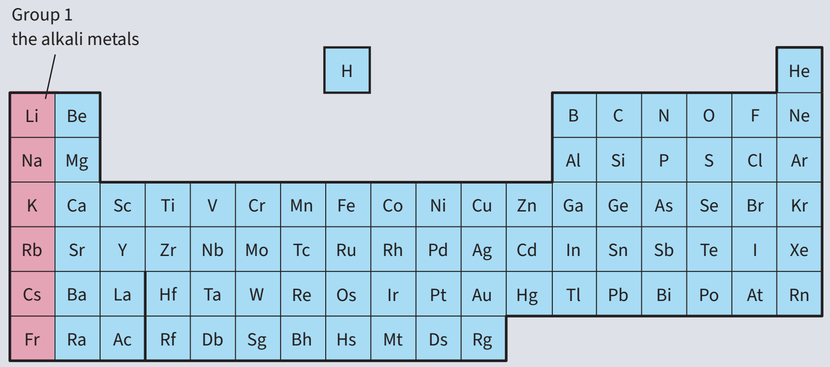
alkali metals chemical properties displacement groups halogens inert isotopes noble gas organised Periodic Table reactivity undiscovered unreactive

Group 1 elements

Group 1 elements react with oxygen, chlorine, and water, for example:

lithium + oxygen ➞ lithium oxide
lithium + chlorine ➞ lithium chloride
lithium + water ➞ lithium hydroxide + hydrogen

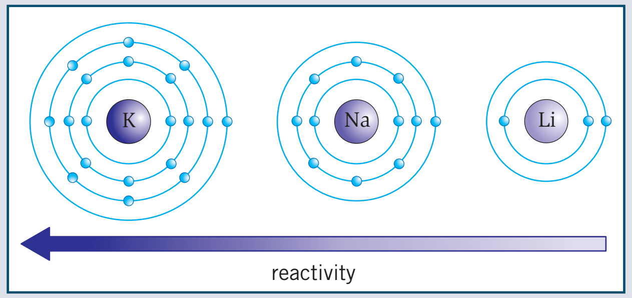
Group 1 elements are called **alkali metals** because they react with water to form an alkali (a solution of their metal hydroxide).



Group 1 properties

Group 1 elements all have one electron in their outer shell. **Reactivity** increases down Group 1 because as you move down the group:

- the atoms increase in size
- the outer electron is further away from the nucleus, and there are more shells shielding the outer electron from the nucleus
- the electrostatic attraction between the nucleus and the outer electron is weaker so it is easier to lose the one outer electron
- the melting point and boiling point decreases down Group 1.



Group 7 elements

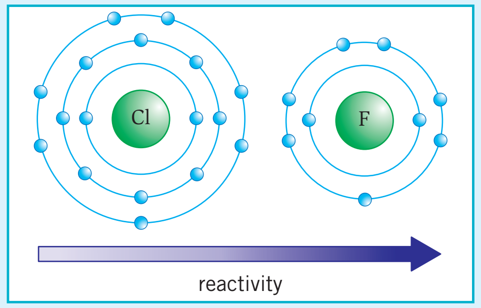
Group 7 elements are called the **halogens**. They are non-metals that exist as molecules made up of pairs of atoms.

Name	Formula	State at room temperature	Melting point and boiling point	Reactivity
fluorine	F_2	gas	increases down the group	decreases down the group
chlorine	Cl_2	gas		
bromine	Br_2	liquid		
iodine	I_2	solid		

Group 7 reactivity

Reactivity decreases down Group 7 because as you move down the group:

- the atoms increase in size
- the outer shell is further away from the nucleus, and there are more shells between the nucleus and the outer shell
- the electrostatic attraction from the nucleus to the outer shell is weaker so it is harder to gain one electron to fill the outer shell.



Group 7 displacement

More reactive Group 7 elements can take the place of less reactive ones in a compound. This is called **displacement**.

For example, fluorine displaces chlorine as it is more reactive:

fluorine + potassium chloride ➞ potassium fluoride + chlorine

Chapter 2: The Periodic Table

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C2 questions	Answers
1	How is the modern Periodic Table ordered?	by atomic number
2	How were the early lists of elements ordered?	by atomic mass
3	Why did Mendeleev swap the order of some elements?	to group them by their chemical properties
4	Why did Mendeleev leave gaps in his Periodic Table?	leave room for elements that had not yet been discovered
5	Why do elements in a group have similar chemical properties?	have the same number of electrons in their outer shell
6	Where are metals and non-metals located on the Periodic Table?	metals to the left, non-metals to the right
7	What name is given to the Group 1 elements?	alkali metals
8	Why are the alkali metals named this?	they are metals that react with water to form an alkali
9	Give the general equations for the reactions of alkali metals with oxygen, chlorine, and water.	metal + oxygen → metal oxide metal + chlorine → metal chloride metal + water → metal hydroxide + hydrogen
10	How does the reactivity of the alkali metals change down the group?	increases (more reactive)
11	Why does the reactivity of the alkali metals increase down the group?	they are larger atoms, so the outermost electron is further from the nucleus, meaning there are weaker electrostatic forces of attraction and more shielding between the nucleus and outer electron, and it is easier to lose the electron
12	What name is given to the Group 7 elements?	halogens
13	Give the formulae of the first four halogens.	F_2, Cl_2, Br_2, I_2
14	How do the melting points of the halogens change down the group?	increase (higher melting point)
15	How does the reactivity of the halogens change down the group?	decrease (less reactive)
16	Why does the reactivity of the halogens decrease down the group?	they are larger atoms, so the outermost shell is further from the nucleus, meaning there are weaker electrostatic forces of attraction and more shielding between the nucleus and outer shell, and it is harder to gain an electron
17	What is a displacement reaction?	when a more reactive element takes the place of a less reactive one in a compound
18	What name is given to the Group 0 elements?	noble gases
19	Why are the noble gases inert?	they have full outer shells so do not need to lose or gain electrons
20	How do the melting points of the noble gases change down the group?	increase (higher melting point)

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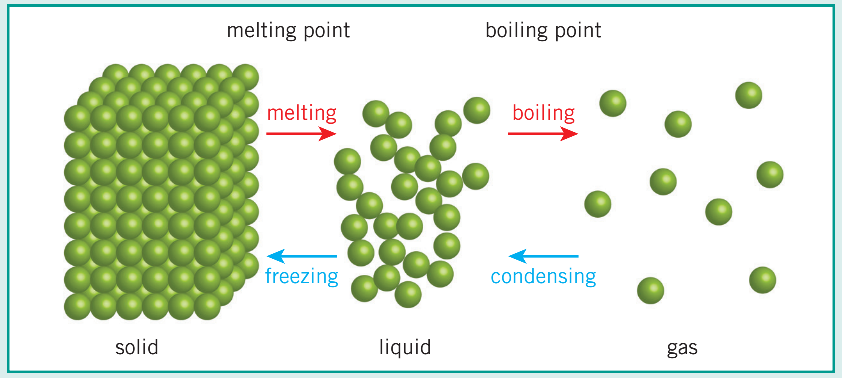
Chapter 3: Bonding 1

Knowledge organiser

Particle model

The three states of matter can be represented in the particle model.



(HT only) This model assumes that:

- there are no forces between the particles
- that all particles in a substance are spherical
- that the spheres are solid.

The amount of energy needed to change the state of a substance depends on the forces between the particles. The stronger the forces between the particles, the higher the melting or boiling point of the substance.

Covalent bonding

Atoms can share or transfer electrons to form strong chemical bonds.

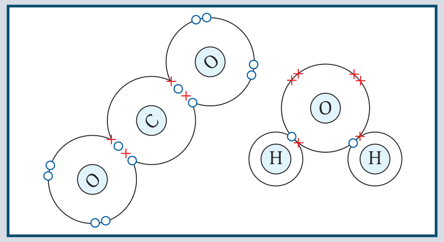
A **covalent bond** is when electrons are *shared* between **non-metal** atoms.

The number of electrons shared depends on how many extra electrons an atom needs to make a full outer shell.

If you include electrons that are shared between atoms, each atom has a full outer shell.

Single bond = each atom shares one pair of electrons.

Double bond = each atom shares two pairs of electrons.



Covalent structures

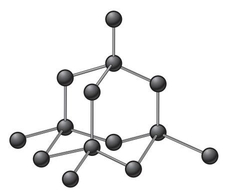
There are three main types of covalent structure:

Structure and bonding

Giant covalent

Many billions of atoms, each one with a strong covalent bond to a number of others.

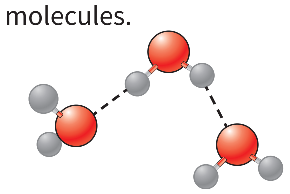
An example of a giant covalent structure is diamond.



Small molecules

Each molecule contains only a few atoms with strong covalent bonds between these atoms. Different molecules are held together by weak **intermolecular forces**.

For example, water is made of small molecules.



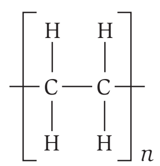
Large molecules

Many repeating units joined by covalent bonds to form a chain.

The small section is bonded to many identical sections to the left and right. The 'n' represents a large number.

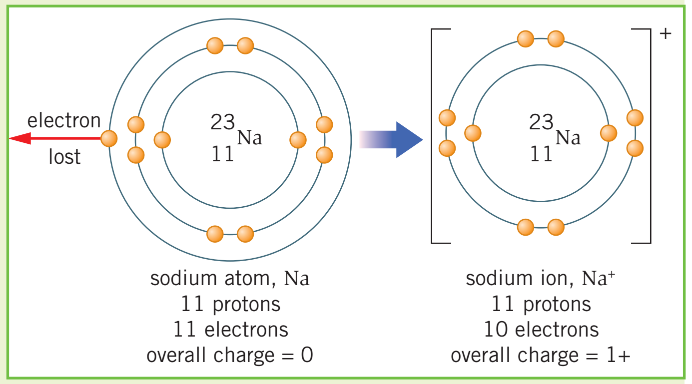
Separate chains are held together by intermolecular forces that are stronger than in small molecules.

Polymers are examples of long molecules.



Ions

Atoms can gain or lose electrons to give them a full outer shell. The number of protons is then different from the number of electrons. The resulting particle has a charge and is called an **ion**.



Conductivity

Solid ionic substances do not conduct electricity because the ions are fixed in position and not free to carry charge.

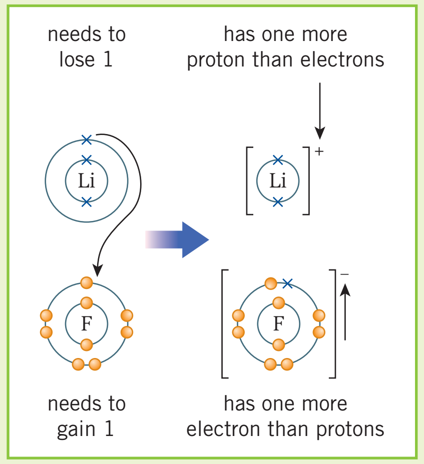
When melted or dissolved in water, ionic substances do conduct electricity because the ions are free to move and carry charge.

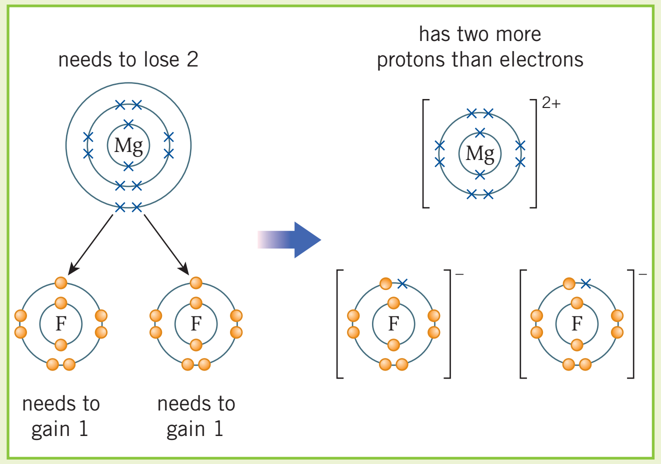
Melting points

Ionic substances have high melting points because the electrostatic force of attraction between oppositely charged ions is strong and so requires lots of energy to break.

Ionic bonding

When metal atoms react with non-metal atoms they **transfer** electrons to the non-metal atom.

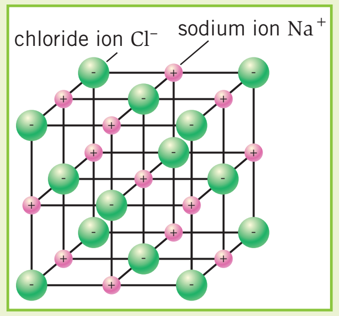




Metal atoms lose electrons to become positive ions. Non-metal atoms gain electrons to become negative ions.

Giant ionic lattice

When metal atoms transfer electrons to non-metal atoms you end up with positive and negative ions. These are attracted to each other by the strong **electrostatic force of attraction**. This is called ionic bonding.

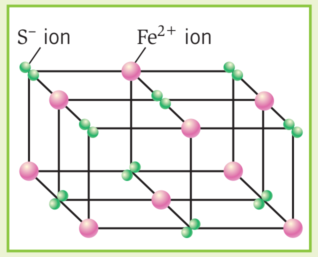


The electrostatic force of attraction works in all directions, so many billions of ions can be bonded together in a 3D structure.

Formulae

The formula of an ionic substance can be worked out

1. from its bonding diagram: for every one magnesium ion there are two fluoride ions – so the formula for magnesium fluoride is MgF_2
2. from a lattice diagram: there are nine Fe^{2+} ions and 18 S^- ions – simplifying this ratio gives a formula of FeS_2



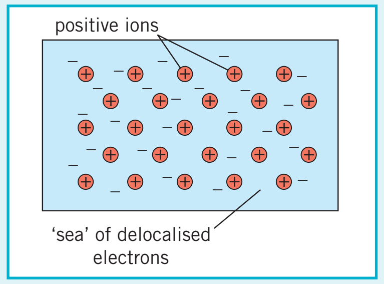
Metals: structure and properties

The atoms that make up metals form layers. The electrons in the outer shells of the atoms are **delocalised** – this means they are free to move through the whole structure.

The positive metal ions are then attracted to these delocalised electrons by the electrostatic force of attraction.

Some important properties of metals are:

- pure metals are **malleable** because the layers can slide over each other
- they are good **conductors** of electricity and of thermal energy because delocalised electrons are free to move through the whole structure
- they have high melting and boiling points because the electrostatic force of attraction between metal ions and delocalised electrons is strong so lots of energy is needed to break it.



Chapter 3: Bonding 2

Knowledge organiser

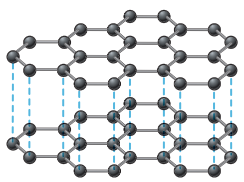
Properties			
Properties	High melting and boiling points because the strong covalent bonds between the atoms must be broken to melt or boil the substances. This requires a lot of energy. Solid at room temperature.	Low melting and boiling points because only the intermolecular forces need to be overcome to melt or boil the substances, not the bonds between the atoms. This does not require a lot of energy as the intermolecular forces are weak. Normally gaseous or liquid at room temperature.	Melting and boiling points are low compared to giant covalent substances but higher than for small molecules. Large molecules have stronger intermolecular forces than small molecules, which require more energy to overcome. Normally solid at room temperature.

Most covalent structures do not conduct electricity because they do not have **delocalised electrons** or ions that are free to move to carry charge.

Graphite

Graphite is a giant covalent structure, but is different to other giant covalent substances.

Structure



Made only of carbon – each carbon atom bonds to three others, and forms hexagonal rings in layers. Each carbon atom has one spare electron, which is delocalised and therefore free to move around the structure.

Hardness

The layers can slide over each other because they are not covalently bonded. Graphite is therefore softer than diamond, even though both are made only of carbon, as each atom in diamond has four strong covalent bonds.

Conductivity

The delocalised electrons are free to move through graphite, so can carry charges and allow an electrical current to flow. Graphite is therefore a conductor of electricity.

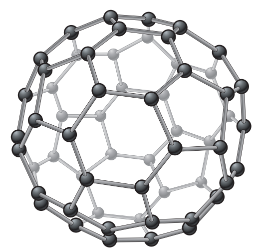
Graphene

Graphene consists of only a single layer of graphite. Its strong covalent bonds make it a strong material that can also conduct electricity. It could be used in composites and high-tech electronics.

Fullerenes

- hollow cages of carbon atoms bonded together in one molecule
- can be arranged as a sphere or a tube (called a **nanotube**)
- molecules held together by weak intermolecular forces, so can slide over each other
- conduct electricity

Spheres

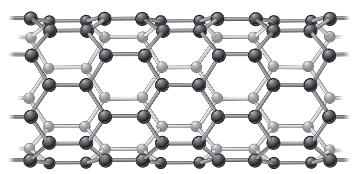


Buckminsterfullerene was the first fullerene to be discovered, and has 60 carbon atoms.

Other fullerenes exist with different numbers of carbon atoms arranged in rings that form hollow shapes.

Fullerenes like this can be used as lubricants and in drug delivery.

Nanotubes

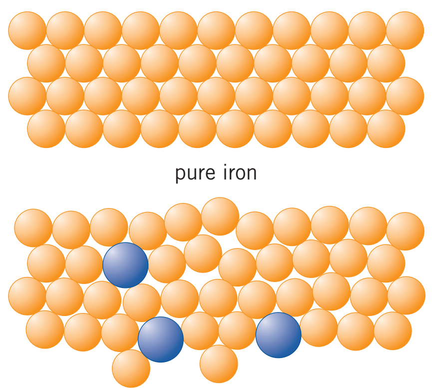


The carbon atoms in nanotubes are arranged in cylindrical tubes.
Their high **tensile strength** (they are difficult to break when pulled) makes them useful in electronics.

Alloys

Pure metals are often too soft to use as they are. Adding atoms of a different element can make the resulting mixture harder because the new atoms will be a different size to the pure metal's atoms. This will disturb the regular arrangement of the layers, preventing them from sliding over each other.

The harder mixture is called an **alloy**.



Key terms Make sure you can write a definition for these key terms.

conductivity *conductor* *delocalised electron* *electrostatic force of attraction* *ion* *lattice* *layer* *malleable* *surface area to volume ratio* *transfer*

Chapter 3: Bonding

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C3 questions	Answers
1	How are covalent bonds formed?	by atoms sharing electrons
2	Which type of atoms form covalent bonds between them?	non-metals
3	Describe the structure and bonding of a giant covalent substance.	billions of atoms bonded together by strong covalent bonds
4	Describe the structure and bonding of small molecules.	small numbers of atoms group together into molecules with strong covalent bonds between the atoms and weak intermolecular forces between the molecules
5	Describe the structure and bonding of polymers.	many identical molecules joined together by strong covalent bonds in a long chain, with weak intermolecular forces between the chains
6	Why do giant covalent substances have high melting points?	it takes a lot of energy to break the strong covalent bonds between the atoms
7	Why do small molecules have low melting points?	only a small amount of energy is needed to break the weak intermolecular forces
8	Why do large molecules have higher melting and boiling points than small molecules?	the intermolecular forces are stronger in large molecules
9	Why do most covalent substances not conduct electricity?	do not have delocalised electrons or ions
10	Describe the structure and bonding in graphite.	each carbon atom is bonded to three others in hexagonal rings arranged in layers – it has delocalised electrons and weak forces between the layers
11	Why can graphite conduct electricity?	the delocalised electrons can move through the graphite
12	Explain why graphite is soft.	layers are not bonded so can slide over each other
13	What is graphene?	one layer of graphite
14	Give two properties of graphene.	strong, conducts electricity
15	What is a fullerene?	hollow cage of carbon atoms arranged as a sphere or a tube
16	What is a nanotube?	hollow cylinder of carbon atoms
17	Give two properties of nanotubes.	high tensile strength, conduct electricity
18	Give three uses of fullerenes.	lubricants, drug delivery (spheres), high-tech electronics
19	What is an ion?	atom that has lost or gained electrons
20	Which kinds of elements form ionic bonds?	metals and non-metals
21	What charges do ions from Groups 1 and 2 form?	Group 1 forms 1+, Group 2 forms 2+
22	What charges do ions from Groups 6 and 7 form?	Group 6 forms 2–, Group 7 forms 1–
23	Name the force that holds oppositely charged ions together.	electrostatic force of attraction
24	Describe the structure of a giant ionic lattice.	regular structure of alternating positive and negative ions, held together by the electrostatic force of attraction
25	Why do ionic substances have high melting points?	electrostatic force of attraction between positive and negative ions is strong and requires lots of energy to break
26	Why don't ionic substances conduct electricity when solid?	ions are fixed in position so cannot move, and there are no delocalised electrons
27	When can ionic substances conduct electricity?	when melted or dissolved
28	Why do ionic substances conduct electricity when melted or dissolved?	ions are free to move and carry charge
29	Describe the structure of a pure metal.	layers of positive metal ions surrounded by delocalised electrons
30	Describe the bonding in a pure metal.	strong electrostatic forces of attraction between metal ions and delocalised electrons
31	What are four properties of pure metals?	malleable, high melting/boiling points, good conductors of electricity, good conductors of thermal energy
32	Explain why pure metals are malleable.	layers can slide over each other easily
33	Explain why metals have high melting and boiling points.	electrostatic force of attraction between positive metal ions and delocalised electrons is strong and requires a lot of energy to break
34	Why are metals good conductors of electricity and of thermal energy?	delocalised electrons are free to move through the metal
35	What is an alloy?	mixture of a metal with atoms of another element
36	Explain why alloys are harder than pure metals.	different sized atoms disturb the layers, preventing them from sliding over each other

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Chapter 4: Calculations

Knowledge organiser

Formula mass

Every substance has a **formula mass**, M_r.

formula mass M_r = sum (relative atomic mass of all the atoms in the formula)

Avogadro's constant (HT only)

One mole of a substance contains 6.02×10^{23} atoms, ions, or molecules. This is **Avogadro's constant**.

One mole of a substance has the same mass as the M_r of the substance. For example, the M_r (H_2O) = 18, so 18 g of water molecules contains 6.02×10^{23} molecules, and is called one mole of water.

You can write this as: moles = $\frac{\text{mass}}{M_r}$

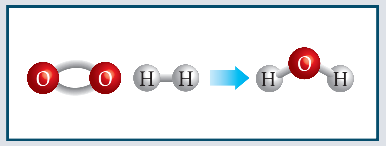
Balancing symbol equations

When writing symbol equations you need to ensure that the number of each atom on each side is equal.

$H_2 + O_2 \rightarrow H_2O$

unbalanced

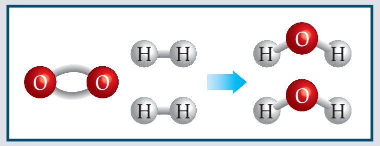
there are 2 hydrogen atoms on each side, but 2 oxygen atoms in the reactants and 1 in the product



$2H_2 + O_2 \rightarrow 2H_2O$

balanced

there are 4 hydrogen atoms on each side, and 2 oxygen atoms on each side



State symbols

A balanced symbol equation should also include state symbols.

State	Symbol
solid	(s)
liquid	(l)
gas	(g)
aqueous or dissolved in water	(aq)

Ratios

Look back at the reaction. In the reaction between hydrogen and oxygen, the ratio of hydrogen to oxygen molecules is 2:1. This means that for every *one* molecule of oxygen, you would need *two* molecules of hydrogen, for example:

- if you had 10 molecules of oxygen you would need 20 molecules of hydrogen
- if you had 1 mole of oxygen you would need 2 moles of hydrogen
- if you had 1.75 moles of oxygen you would need 3.5 moles of hydrogen.

A balanced symbol equation shows the ratios of the reactants and products in a chemical reaction.

mol is a the unit of moles

Concentration

Concentration is the amount of solute in a volume of solvent.

The unit of concentration is g/dm³. Concentration can be calculated using:

$$\textit{concentration (g/dm}^3\textit{)} = \frac{\textit{mass (g)}}{\textit{volume (dm}^3\textit{)}}$$

Sometimes volume is measured in cm³:

$$\textit{volume (dm}^3\textit{)} = \frac{\textit{volume (cm}^3\textit{)}}{1000}$$

- lots of solute in little solution = high concentration
- little solute in lots of solution = low concentration

Concentration in mol/dm³

Concentration can also be measured in mol/dm³.

$$\textit{concentration of solution (mol/dm}^3\textit{)} = \frac{\textit{number of moles of solute}}{\textit{volume of solution (dm}^3\textit{)}}$$

You can use this formula and mass = moles × M_r to calculate the mass of solute dissolved in a solution.

- The greater the mass of solute in solution, the greater the number of moles of solute, and therefore the greater the concentration.
- If the same number moles of solute is dissolved in a smaller volume of solution, the concentration will be greater.

Using balanced equations (HT only)

In a balanced symbol equation the sum of the M_r of the reactants equals the sum of the M_r of the products.

If you are asked what mass of a product will be formed from a given mass of a specific reactant, you can use the steps below to calculate the result.

1. balance the symbol equation
2. calculate moles of the substance with a known mass using moles = $\frac{\text{mass}}{M_r}$
3. using the balanced symbol equation, work out the number of moles of the unknown substance
4. calculate the mass of the unknown substance using mass = moles × M_r

If you are asked to balance an equation, you can use the steps below to work out the answer.

1. work out M_r of all the substances
2. calculate the number of moles of each substance in the reaction using moles = $\frac{\text{mass}}{M_r}$
3. convert to a whole number ratio
4. balance the symbol equation

Excess and limiting reactants (HT only)

reactants, often one of the reactants will run out before the others. You then have some of the other reactants left over. The reactant that is left over is in **excess**. The reactant that runs out is the **limiting reactant**.

To work out which reactants are in excess and which is the limiting reactant, you need to:

1. write the balanced symbol equation for the reaction
2. pick one of the reactants and its quantity as given in the question
3. use the ratio of the reactants in the balanced equation to see how much of the other reactant you need
4. compare this value to the quantity given in the question
5. determine which reactant is in excess and which is limiting.

Key terms

Make sure you can write a definition for these key terms.

Avogadro's constant · concentration · excess · formula · mass · limiting reactant · mole

Chapter 4: Calculations

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C4 questions	Answers
1	What is a mole?	mass of a substance that contains 6.02×10^{23} particles
2	Give the value for Avogadro's constant.	6.02×10^{23}
3	Which formula is used to calculate the number of moles from mass and M_r?	moles = $\frac{\text{mass}}{M_r}$
4	Which formula is used to calculate the mass of a substance from number of moles and M_r?	mass = moles × M_r
5	What is a limiting reactant?	the reactant that is completely used up in a chemical reaction
6	What is a unit for concentration?	g/dm^3 or mol/dm^3
7	Which formula is used to calculate concentration from mass and volume?	concentration (g/dm^3) = $\frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$
8	Which formula is used to calculate volume from concentration and mass?	volume (dm^3) = $\frac{\text{mass (g)}}{\text{concentration (g/dm}^3\text{)}}$
9	Which formula is used to calculate mass from concentration in g/dm^3 and volume?	mass (g) = concentration (g/dm^3) × volume (dm^3)
10	How can you convert a volume reading in cm^3 to dm^3?	divide by 1000
11	If the amount of solute in a solution is increased, what happens to its concentration?	increases
12	If the volume of water in a solution is increased, what happens to its concentration?	decreases
13	How can concentration in mol/dm^3 be calculated?	$\frac{\text{moles of solute}}{\text{volume (dm}^3\text{)}}$

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Chapter 5: Chemical changes 1

Knowledge organiser

Reactions of metals

The **reactivity** of a metal is how chemically reactive it is. When added to water, some metals react very vigorously – these metals have *high* reactivity. Other metals will barely react with water or acid, or won't react at all – these metals have *low* reactivity.

Reactivity series

The reactivity series places metals in order of their reactivity.

Sometimes, for example in the table below, hydrogen and carbon are included in the series, even though they are non-metals.

Reaction with water	Reaction with acid	Reactivity series: Metal	Reactivity series: Reactivity	Extraction method
fizzes, gives off hydrogen gas	explodes	potassium	high reactivity	electrolysis
fizzes, gives off hydrogen gas	explodes	sodium		electrolysis
fizzes, gives off hydrogen gas	explodes	lithium		electrolysis
fizzes, gives off hydrogen gas	fizzes, gives off hydrogen gas	calcium		electrolysis
reacts very slowly	fizzes, gives off hydrogen gas	magnesium		electrolysis
reacts very slowly	fizzes, gives off hydrogen gas	aluminium	Decreasing reactivity	electrolysis
		(carbon)		
reacts very slowly	fizzes, gives off hydrogen gas	zinc		reduction with carbon
reacts very slowly	fizzes, gives off hydrogen gas	iron		reduction with carbon
no reaction	reacts slowly with warm acid	tin		reduction with carbon
no reaction	reacts slowly with warm acid	lead		reduction with carbon
		(hydrogen)		
no reaction	no reaction	copper		reduction with carbon
no reaction	no reaction	silver		mined from the Earth's crust
no reaction	no reaction	gold	low reactivity	mined from the Earth's crust

Metal extraction

Some metals, like gold, are so unreactive that they are found as pure metals in the Earth's crust and can be mined.

Most metals exist as compounds in rock and have to be extracted from the rock. If there is enough metal compound in the rock to be worth extracting it is called an **ore**.

Metals that are less reactive than carbon can be extracted by reduction with carbon. For example:

iron oxide + carbon → iron + carbon dioxide

Metals that are more reactive than carbon can be extracted using a process called **electrolysis**.

Reduction and oxidation

If a substance gains oxygen in a reaction, it has been **oxidised**.

If a substance loses oxygen in a reaction, it has been **reduced**.

For example:

iron + oxygen → iron oxide

iron has been oxidised

iron oxide + carbon → iron + carbon dioxide

iron oxide has been reduced

Salts

When acids react with metals or metal compounds, they form salts. A salt is a compound where the hydrogen from an acid has been replaced by a metal. For example nitric acid, HNO_3, reacts with sodium to form $NaNO_3$. The H in nitric acid is replaced with Na.

The table shows how to name salts.

Acid	hydrochloric acid	sulfuric acid	nitric acid
Formula	HCl	H_2SO_4	HNO_3
Ions formed in solution	H^+ and Cl^-	$2H^+$ and SO_4^{2-}	H^+ and NO_3^-
Type of salt formed	metal chloride	metal sulfate	metal nitrate
Sodium salt example	sodium chloride, $NaCl$	sodium sulfate, Na_2SO_4	sodium nitrate, $NaNO_3$

Displacement reactions

In a **displacement** reaction a *more* reactive element takes the place of a *less* reactive element in a compound.

For example:

copper sulfate + iron → iron sulfate + copper

$$CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$$

Iron is more reactive than copper, so iron displaces the copper in copper sulfate.

Reactivity and ions

A metal's reactivity depends on how readily it forms an **ion** by losing electrons.

In the displacement reaction of copper sulfate and iron, iron forms an ion more easily than copper.

At the end of the reaction you are left with iron ions, not copper ions.

Ionic equations (HT only)

When an ionic compound is dissolved in a solution, we can write the compound as its separate ions. For example, $CuSO_4(aq)$ can be written as $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$.

The displacement reaction of copper sulfate and iron can be written as:

$$Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

The SO_4^{2-} is unchanged in the reaction – it is a **spectator ion**. Spectator ions are removed from the equation to give an **ionic equation**:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

Metals, covalent substances, and solid ionic substances do not split into ions in the ionic equation.

Steps for writing an ionic equation (HT only)

1. check symbol equation is balanced
2. identify all aqueous ionic compounds
3. write those compounds out as ions
4. remove spectator ions.

Reduction and oxidation: electrons (HT only)

Oxidation and reduction (**redox** reactions) can be defined in terms of oxygen, but can also be defined as the loss or gain of electrons.

Oxidation is the *loss* of electrons, and reduction is the *gain* of electrons.

In the example displacement reaction:

- iron atoms have been oxidised
- copper ions have been reduced.

Half equations (HT only)

In the displacement reaction, an iron atom loses two electrons to form a iron ion:

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

A copper ion gains two electrons to form a copper atom:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

These two equations are called **half equations** – they each show half of the ionic equation.

Acids and alkalis

Acids are compounds that, when dissolved in water, release H^+ ions. There are three main acids: sulfuric acid H_2SO_4, nitric acid HNO_3, and hydrochloric acid HCl.

Alkalis are compounds that, when dissolved in water, release OH^- ions.

The **pH** scale is a measure of acidity and alkalinity. It runs from 1 to 14.

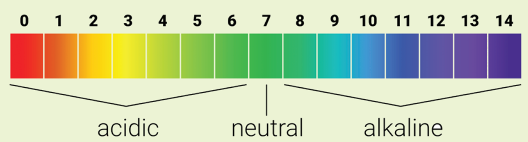
- Aqueous solutions with pH < 7 are acidic.
- Aqueous solutions with pH > 7 are alkaline.
- Aqueous solutions with pH = 7 are neutral.

Indicators

Indicators can show if something is an acid or an alkali.

- **Universal indicator** can also tell us the approximate pH of a solution.
- Electronic pH probes can give us the exact pH of a solution.

The pH scale



Chapter 5: Chemical changes 2

Knowledge organiser

Reactions of acids

Reactions of acids with metals

Acids react with some metals to form salts and hydrogen gas.

magnesium + hydrochloric acid → sodium chloride + hydrogen

Neutralisation reactions

Reactions of acids with metal hydroxides

Acids react with metal hydroxides to form salts and water.

hydrochloric acid + sodium hydroxide → sodium chloride + water

The ionic equation for this reaction is always:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Reactions of acids with metal oxides

Acids react with metal oxides to form salts and water.

hydrochloric acid + sodium oxide → sodium chloride + water

Reactions of acids with metal carbonates

Acids react with metal carbonates to form a salt, water, and carbon dioxide.

hydrochloric acid + sodium carbonate → sodium chloride + water + carbon dioxide

Alkalis and bases

Bases neutralise acids to form water in **neutralisation** reactions. Some metal hydroxides dissolve in water to form alkaline solutions, called alkalis.

Some metal oxides and metal hydroxide do not dissolve in water. They are **bases**, but are not alkalis.

Strong and weak acids

Sulfuric acid, nitric acid, and hydrochloric acid, are all **strong acids**. This means that, when dissolved in water, every molecule splits up into ions – they are completely ionised:

- $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
- $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
- $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

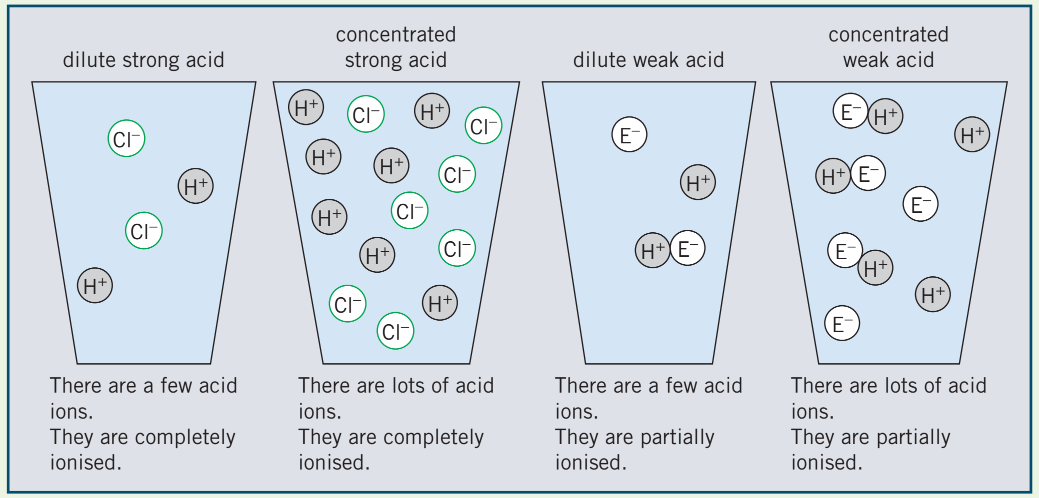
Ethanoic acid, citric acid, and carbonic acid are **weak acids**. This means that only a percentage of their molecules split up into ions when dissolved in water – they are partially ionised.

For a given concentration, the *stronger* the acid, the *lower* the pH.

Concentrated and dilute acids

Concentration tells us how much of a substance there is dissolved in water:

- more concentrated acids have lots of acid in a small volume of water
- less concentrated acids (dilute acids) have little acid in a large volume of water.



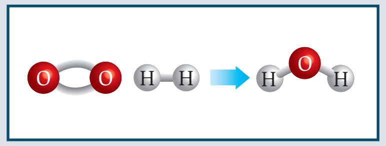
Balancing symbol equations

When writing symbol equations you need to ensure that the number of each atom on each side is equal.

$$H_2 + O_2 \rightarrow H_2O$$

unbalanced

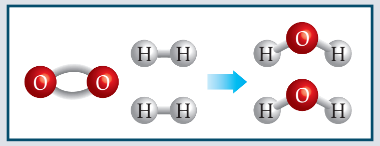
there are 2 hydrogen atoms on each side, but 2 oxygen atoms in the reactants and 1 in the product



$$2H_2 + O_2 \rightarrow 2H_2O$$

balanced

there are 4 hydrogen atoms on each side, and 2 oxygen atoms on each side



State symbols

A balanced symbol equation should also include state symbols.

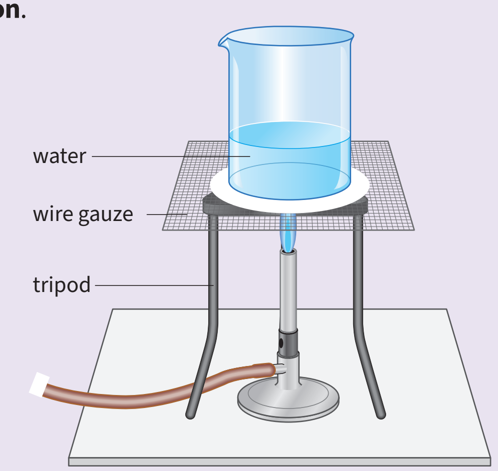
State	Symbol
solid	(s)
liquid	(l)
gas	(g)
aqueous or dissolved in water	(aq)

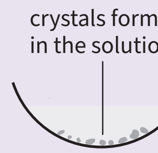
Crystallisation

You can produce a solid salt from an insoluble base by **crystallisation**.

The experimental method is:

1. Choose the correct acid and base to produce the salt.
2. Put some of the dilute acid into a flask. Heat gently with a Bunsen burner.
3. Add a small amount of the base and stir.
4. Keep adding the base until no more reacts – the base is now in excess.
5. Filter to remove the unreacted base.
6. Add the remaining solution to an evaporating dish.
7. Use a water bath or electric heater to evaporate the water. The salt crystals will be left behind.





Key terms

Make sure you can write a definition for these key terms.

displacement, electrolysis, extraction, half equation, ion, ionic equation, metal, ore, oxidation, reactivity, reactivity series, redox, reduction, spectator ion, state symbols

Chapter 5: Chemical changes

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C5 questions	Answers
1	What does reactivity mean?	how vigorously a substance chemically reacts
2	How can metals be ordered by their reactivity?	by comparing their reactions with water, acid, or oxygen
3	What name is given to a list of metals ordered by their reactivity?	reactivity series
4	In terms of electrons, what makes some metals more reactive than others?	they lose their outer shell electron(s) more easily
5	Why are gold and silver found naturally as elements in the Earth's crust?	they are very unreactive
6	What is an ore?	rock containing enough of a metal compound to be economically worth extracting
7	How are metals less reactive than carbon extracted from their ores?	reduction with carbon
8	In terms of oxygen, what is oxidation?	addition of oxygen
9	In terms of oxygen, what is reduction?	removal of oxygen
10	Why can metals like potassium and aluminium not be extracted by reduction with carbon?	they are more reactive than carbon
11	How are metals more reactive than carbon extracted from their ores?	electrolysis
12	What is a displacement reaction?	a more reactive substance takes the place of a less reactive substance in a compound
13	What is an ionic equation?	equation which gives some substances as ions and has spectator ions removed
14	What type of substance is given as ions in an ionic equation?	ionic compounds in solution (or liquid)
15	What is a spectator ion?	ion that is unchanged in a reaction
16	What is a half equation?	equation that shows whether a substance is losing or gaining electrons
17	In terms of electrons, what is oxidation?	loss of electrons
18	In terms of electrons, what is reduction?	gain of electrons
19	In terms of pH, what is an acid?	a solution with a pH of less than 7
20	In terms of pH, what is a neutral solution?	a solution with a pH of 7
21	In terms of H^+ ions, what is an acid?	a substance that releases H^+ ions when dissolved in water
22	How is the amount of H^+ ions in a solution related to its pH?	the more H^+ ions, the lower the pH
23	What are the names and formulae of three main acids?	hydrochloric acid, HCl; sulfuric acid, H_2SO_4; nitric acid, HNO_3
24	How do you measure the pH of a substance?	universal indicator or pH probe
25	What is a strong acid?	an acid where the molecules or ions completely ionise in water
26	What is a weak acid?	an acid where the molecules or ions partially ionise in water
27	What is a salt?	compound formed when a metal ion takes the place of a hydrogen ion in an acid
28	Which type of salts do sulfuric acid, hydrochloric acid, and nitric acid form?	sulfates, chlorides, nitrates
29	What are the products of a reaction between a metal and an acid?	salt + hydrogen
30	What are the products of a reaction between a metal hydroxide and an acid?	salt + water
31	What are the products of a reaction between a metal oxide and an acid?	salt + water
32	What are the products of a reaction between a metal carbonate and an acid?	salt + water + carbon dioxide
33	What is a base?	substance that reacts with acids in neutralisation reactions
34	What is an alkali?	substance that dissolves in water to form a solution above pH 7
35	What is a neutralisation reaction?	a reaction between an acid and a base to produce water
36	What is the ionic equation for a reaction between an acid and an alkali?	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
37	How can you obtain a solid salt from a solution?	crystallisation
38	When an acid reacts with a metal, which species is oxidised?	the metal
39	When an acid reacts with a metal, which species is reduced?	hydrogen
40	What are the four state symbols and what do they stand for?	(s) solid, (l) liquid, (g) gas, (aq) aqueous or dissolved in water

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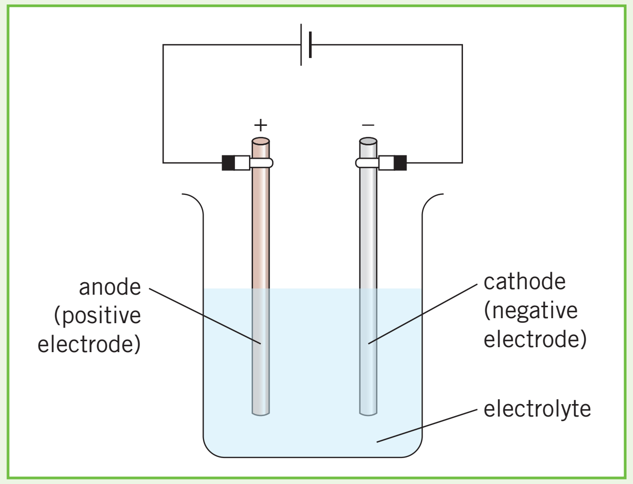
Chapter 6: Electrolysis

Knowledge organiser

Electrolysis

In the process of **electrolysis**, an electric current is passed through an **electrolyte**. An electrolyte is a liquid or solution that contains ions and so can conduct electricity. This causes the ions to move to the **electrodes**, where they form pure elements.



Electrolysis of molten compounds

Solid ionic compounds do not conduct electricity as the ions cannot move. To undergo electrolysis they must be molten or dissolved, so the ions are free to move.

When an ionic compound is molten:

- The positive metal ions are *attracted* to the **cathode**, where they will *gain* electrons to form the pure metal
- The negative non-metal ions are *attracted* to the **anode**, where they will *lose* electrons and become the pure non-metal.

For example, molten sodium chloride, NaCl, can undergo electrolysis to form sodium at the cathode and chlorine at the anode.

Half equations (HT only)

sodium chloride → sodium + chlorine

$2NaCl(l) \rightarrow 2Na(s) + Cl_2(g)$

- at the cathode:

$2Na^+(l) + 2e^- \rightarrow 2Na(s)$

- at the anode:

$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Electrolysis of aqueous solutions

Solid ionic compounds can also undergo electrolysis when dissolved in water.

- It requires less energy to dissolve ionic compounds in water than it does to melt them.
- However, in the electrolysis of solutions, the pure elements are not always produced. This is because the water can also undergo ionisation:

$$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$$

Products at the anode

In In the electrolysis of a solution, if the non-metal contains oxygen then oxygen gas is formed at the anode:

- The $OH^-(aq)$ ions formed from the ionisation of water are attracted to the anode.
- The $OH^-(aq)$ ions lose electrons to the anode and form oxygen gas.
- $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

If the non-metal ion is a halogen, then the halogen gas is formed at the anode.

- $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

most reactive

potassium
sodium
calcium
magnesium
aluminium
(carbon)
zinc
iron
tin
lead
(hydrogen)
copper
silver
gold
platinum

least reactive

Products at the cathode

In the electrolysis of a solution, if the metal is *more* **reactive** than hydrogen then hydrogen gas is formed at the cathode:

- The $H^+(aq)$ ions from the ionisation of water are attracted to the cathode and react with it.
- The $H^+(aq)$ ions gain electrons from the cathode and form hydrogen gas.
- $2H^+(aq) + 2e^- \rightarrow H_2(g)$
- The metal ions remain in solution.

Electrolysis of aluminium oxide

Electrolysis can be used to extract metals from their ionic compounds.

Electrolysis is used if the metal is more reactive than carbon.

Aluminium is extracted from aluminium oxide by electrolysis.

1 The aluminium oxide is mixed with a substance called **cryolite**, which lowers the melting point.
2 The mixture is then heated until it is molten.
3 The resulting molten mixture undergoes electrolysis.

aluminium oxide → aluminium + oxygen

$2Al_2O_3(l) \rightarrow 4Al(l) + 3O_2(g)$

cathode: pure aluminium is formed

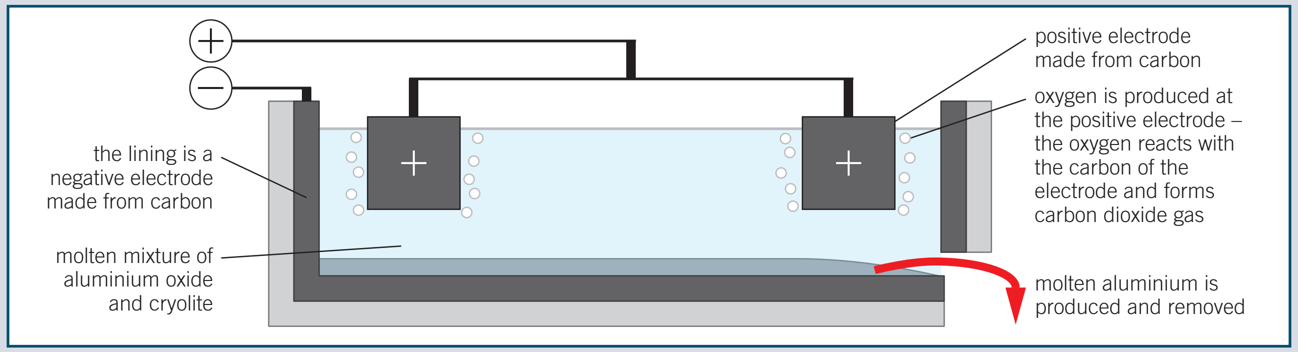
$Al^{3+}(l) + 3e^- \rightarrow Al(l)$

anode: oxygen is formed

$2O^{2-}(l) \rightarrow O_2(g) + 4e^-$

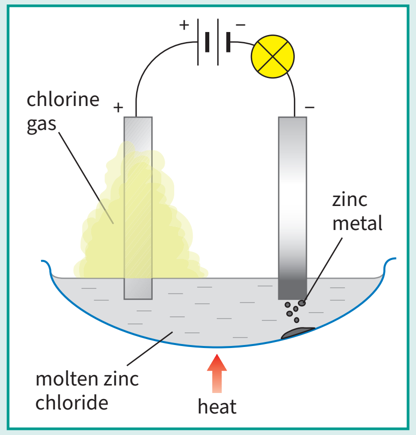
In the electrolysis of aluminium, the anode is made of graphite.

The graphite reacts with the oxygen to form carbon dioxide and so slowly wears away. It therefore needs to be replaced frequently.



Electrolysis of zinc chloride

Molten zinc chloride is broken down by electrolysis. This means zinc metal is collected at the cathode and a pale green chlorine gas is collected at the anode. Free ions from the molten zinc chloride are able to move around and carry electric currents, hence why the bulb lights up.



Key terms Make sure you can write a definition for these key terms.

anode *cathode* *cryolite* *electrode* *electrolysis* *electrolyte* *reactivity*

Chapter 6: Electrolysis

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C6 questions	Answers
1	What is electrolysis?	process of using electricity to extract elements from a compound
2	What is the name of the positive electrode?	anode
3	What is the name of the negative electrode?	cathode
4	What is an electrolyte?	liquid or solution that contains ions and so can conduct electricity
5	Where are metals formed?	cathode
6	Where are non-metals formed?	anode
7	How can ionic substances be electrolysed?	by melting or dissolving them, and then passing a direct current through them
8	Why can solid ionic substances not be electrolysed?	they do not conduct electricity, or the ions cannot move
9	In the electrolysis of solutions, when is the metal *not* produced at the cathode?	when the metal is more reactive than hydrogen
10	In the electrolysis of a metal halide solution, what is produced at the anode?	halogen
11	In the electrolysis of a metal sulfate solution, what is produced at the anode?	oxygen
12	What is the half equation for the ionisation of water?	$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$
13	What metals are extracted from ionic compounds by using electrolysis?	metals that are more reactive than carbon
14	In the electrolysis of aluminium oxide, why is the aluminium oxide mixed with cryolite?	to lower the melting point
15	In the electrolysis of aluminium oxide, what are the anodes made of?	graphite
16	In the electrolysis of aluminium oxide, why do the anodes need to be replaced?	they react with the oxygen being formed

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Chapter 7: Energy changes

Knowledge organiser

Energy changes

During a chemical reaction, energy transfers occur.

Energy can be transferred:

- to the surroundings – **exothermic**
- from the surroundings – **endothermic**

This energy transfer can cause a temperature change.

Energy is always conserved in chemical reactions.

This means that there is the same amount of energy in the Universe at the start of a chemical reaction as at the end of the chemical reaction.

The surroundings

When chemists say energy is transferred from or to "the *surroundings*" they mean "everything that isn't the reaction".

For example, imagine you have a reaction mixture in a test tube. If you measure the temperature in the test tube using a thermometer, the thermometer is then part of the surroundings.

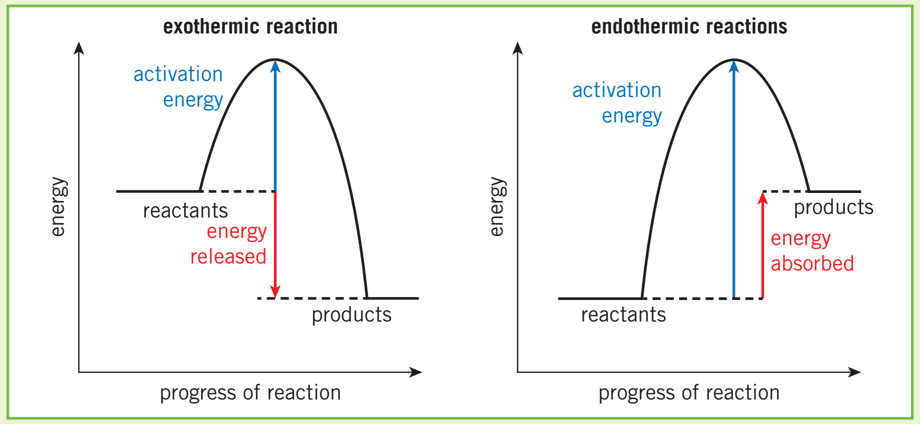
- If the thermometer records an increase in temperature, the reaction in the test tube is exothermic.
- If the thermometer records a decrease in temperature, the reaction in the test tube is endothermic.

Reaction	Energy transfer	Temperature change	Example	Everyday use	Bonds
exothermic	to the surroundings	temperature of the surroundings increases	• **oxidation** • **combustion** • **neutralisation**	• self-heating cans • hand warmers	more energy released when making bonds than required to break bonds
endothermic	from the surroundings	temperature of the surroundings decreases	• **thermal decomposition** • citric acid and sodium hydrogen carbonate	• sports injury packs	less energy released when making bonds than required to break bonds

Reaction profiles

A **reaction profile** shows whether a reaction is exothermic or endothermic.

The **activation energy** is the minimum amount of energy that particles must have to react when they collide.



Bonds (HT only)

Atoms are held together by strong chemical bonds. In a reaction, those bonds are broken and new ones are made between different atoms.

- Breaking a bond requires energy so is endothermic.
- Making a bond releases energy so is exothermic.

Breaking bonds

If a lot of energy is released when making the bonds and only a little energy is required to break them, then overall energy is released and the reaction as a whole is exothermic.

Making bonds

If a little energy is released when making the bonds and a lot is required to break them, then overall energy is taken in and the reaction as a whole is endothermic.

Bond calculations

Different bonds require different amounts of energy to be broken (their **bond energies**). To work out the overall energy change of a reaction, you need to:

1. work out how much energy is required to break all the bonds in the reactants
2. work out how much energy is released when making all the bonds in the products.

overall energy transferred = energy required to break bonds – energy required to make bonds

- A positive number means an endothermic reaction.
- A negative number means an exothermic number.

Key terms

Make sure you can write a definition for these key terms.

activation energy   battery   bond energy   combustion   endothermic   exothermic   neutralisation   oxidation   reaction profile   thermal decomposition

Chapter 7: Energy changes

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C7 questions	Answers
1	What is an exothermic energy transfer?	transfer to the surroundings
2	What is an endothermic energy transfer?	transfer from the surroundings
3	What is a reaction profile?	diagram showing how the energy changes in a reaction
4	What is the activation energy?	minimum amount of energy required before a collision will result in a reaction
5	What is bond energy?	the energy required to break a bond or the energy released when a bond is formed
6	In terms of bond breaking and making, what is an exothermic reaction?	less energy is required to break the bonds than is released when making the bonds
7	In terms of bond breaking and making, what is an endothermic reaction?	more energy is required to break the bonds than is released when making the bonds

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Chapter 8: Rates and equilibrium 1

Knowledge organiser

Rates of reaction

The **rate of a reaction** is how quickly the reactants turn into the products.

To calculate the rate of a reaction, you can measure:

- how quickly a reactant is used up

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

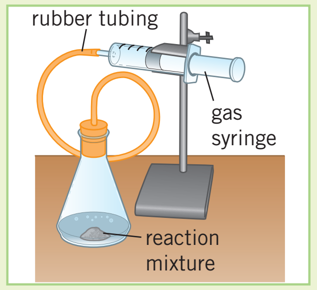
- how quickly a product is produced.

$$\text{mean rate of reaction} = \frac{\text{quantity of product formed}}{\text{time taken}}$$

For reactions that involve a gas, this can be done by measuring how the mass of the reaction changes or the volume of gas given off by the reaction.

Volume of gas produced

The reaction mixture is connected to a gas syringe or an upside down measuring cylinder. As the reaction proceeds the gas is collected.



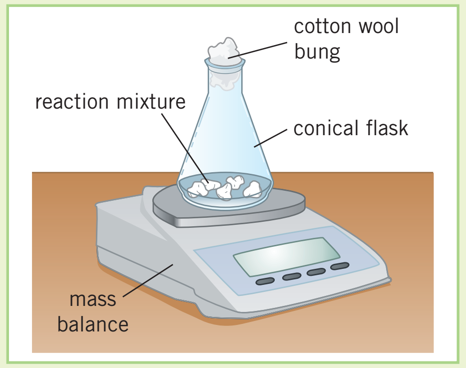
The rate for the reaction is then:

$$\text{rate} = \frac{\text{volume of gas produced}}{\text{time taken}}$$

Volume is measured in cm^3 and time in seconds, so the unit for rate is cm^3/s.

Change in mass

The reaction mixture is placed on a mass balance. As the reaction proceeds and the gaseous product is given off, the mass of the flask will decrease.



The rate for the reaction is then:

$$\text{rate} = \frac{\text{change in the mass}}{\text{time taken}}$$

The mass is measured in grams and time is measured in seconds. Therefore, the unit of rate is g/s.

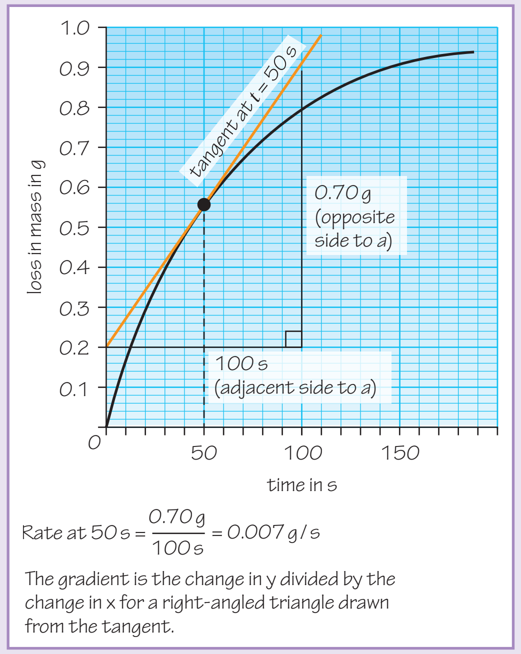
Calculating rate from graphs (HT only)

The results from an experiment can be plotted on a graph.

- A steep gradient means a high rate of reaction – the reaction happens quickly.
- A shallow gradient means a low rate of reaction – the reaction happens slowly.

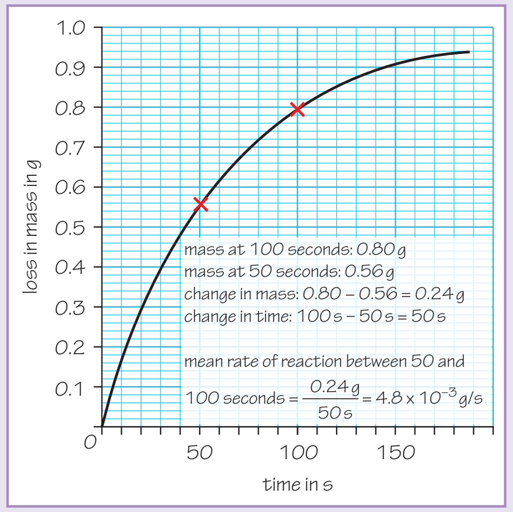
Mean rate at specific time

To obtain the rate at a specific time draw a **tangent** to the graph and calculate its **gradient**.



Mean rate between two points in time

To get the mean rate of reaction between two points in time:



Collision theory

For a reaction to occur, the reactant particles need to collide. When the particles collide, they need to have enough energy to react or they will just bounce apart. This amount of energy is called the **activation energy**.

You can increase the rate of a reaction by:

- increasing the **frequency of collisions**
- increasing the energy of the particles when they collide.

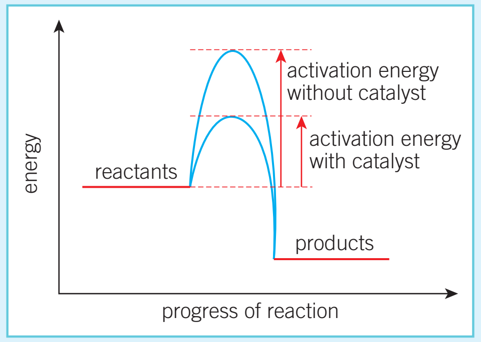
Factors affecting rate of reaction

Condition that increases rate	How is this condition caused?	Why it has that effect
increasing the temperature	Heat the container in which the reaction is taking place.	1 particles move faster, leading to more frequent collisions 2 particles have more energy, so more collisions result in a reaction note that these are two *separate* effects
increasing the concentration of solutions	Use a solution with more solute in the same volume of solvent.	there are more reactant particles in the reaction mixture, so collisions become more frequent
increasing the pressure of gases	Increase the number of gas particles you have in the container or make the container smaller.	less space between particles means more frequent collisions
increasing the surface area of solids	Cut the solid into smaller pieces, or grind it to create a powder, increasing the surface area. Larger pieces decrease the surface area.	only reactant particles on the surface of a solid are able to collide and react; the greater the surface area the more reactant particles are exposed, leading to more frequent collisions

Catalysts

Some reactions have specific substances called **catalysts** that can be added to increase the rate. These substances are not used up in the reaction.

A catalyst provides a different reaction pathway that has a lower activation energy. As such, more particles will collide with enough energy to react, so more collisions result in a reaction.



Chapter 8: Rates and equilibrium 2

Knowledge organiser

Reaction conditions

The conditions of a reaction refer to the external environment of the reaction. When the reaction occurs in a closed system, you can change the conditions by:

- changing the concentration of one of the substances
- changing the temperature of the entire reaction vessel
- changing the pressure inside the vessel.

Le Châtelier's principle (HT only)

At equilibrium, the amount of reactants and products is constant. In order to change the amounts of reactant and product at equilibrium the *conditions* of the reaction must be changed. The closed system will then counteract the change by favouring either the forward reaction or the reverse reaction. This is known as **Le Châtelier's principle**.

For example, lowering the concentration of the product in the system causes the forward reaction to be **favoured** to increase the concentration of the product.

Changing concentrations (HT only)

Change	Effect	Explanation
decrease concentration of product	favours the forward reaction	opposes the change by making *less* reactant and *more* product
increase concentration of product	favours the reverse reaction	opposes the change by making *more* reactant and *less* product
decrease concentration of reactant	favours the reverse reaction	opposes the change by making *more* reactant and *less* product
increase concentration of reactant	favours the forward reaction	opposes the change by making *less* reactant and *more* product

Changing temperature (HT only)

Change	Effect	Explanation
increase temperature of surroundings	favours the endothermic reaction	opposes the change by decreasing the temperature of the surroundings
decrease temperature of surroundings	favours the exothermic reaction	opposes the change by increasing the temperature of the surroundings

Changing pressure (HT only)

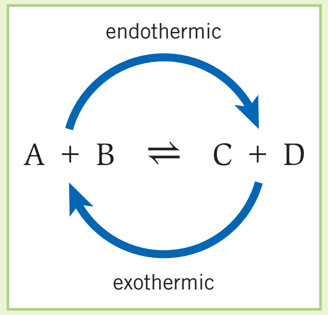
Change	Effect	Explanation
increase the pressure	favours the reaction that results in fewer molecules	decreasing the number of molecules within the vessel opposes the change because it decrease pressure
decrease the pressure	favours the direction that results in more molecules	increasing the number of molecules within the vessel opposes the change because it increase pressure

Key terms

Make sure you can write a definition for these key terms.

activation energy · *catalyst* · *collision* · *collision theory* · *closed system* · *conditions* · *dynamic equilibrium* · *frequency of collision* · *gradient* · *Le Châtelier's principle* · *rate of reaction* · *reversible reaction* · *tangent*

Reversible reactions



In some reactions, the products can react to produce the original reactants. This is called a **reversible reaction**. When writing chemical equations for reversible reactions, use the ⇌ symbol.

In this reaction:

- A and B can react to form C and D – the forward reaction
- C and D can react to form A and B – the reverse reaction.

The different directions of the reaction have opposite energy changes.

If the forward reaction is *endothermic*, the reverse reaction will be *exothermic*.

The same amount of energy is transferred in each direction.

Equilibrium

In a **closed system** no reactants or products can escape. If a reversible reaction is carried out in a closed system, it will eventually reach **dynamic equilibrium** – a point in time when the forward and reverse reactions have the same rate.

At dynamic equilibrium:

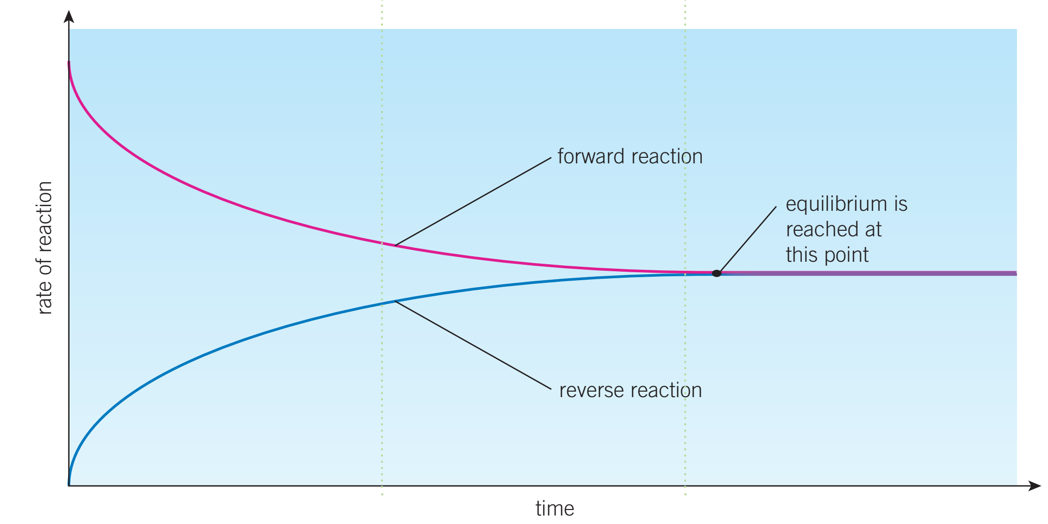
- the reactants are still turning into the products
- the products are still turning back into the reactants
- *the rates of* these two processes are *equal*, so overall the amount of reactants and products are constant.

Dynamic equilibrium

At dynamic equilibrium the amount of reactant and product are constant, but not necessarily equal.

You could have a mixture of reactants and products in a 50:50 ratio, in a 75:25 ratio, or in any ratio at all. The **conditions** of the reaction are what change that ratio.

How dynamic equilibrium is reached

Progress of reaction	start of reaction	middle of reaction	at dynamic equilibrium
Amount of A + B	high	decreasing	constant
Frequency of collisions A + B	high	decreasing	constant
Rate of forward reaction	high	decreasing	same as rate of reverse reaction
			
Amount of C + D	zero	increasing	constant
Frequency of collisions C + D	no collisions	increasing	constant
Rate of reverse reaction	zero	increasing	same as rate of forward reaction

Chapter 8: Rates and equilibrium

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C8 questions	Answers
1	What is the rate of a reaction?	how quickly reactants are used up or products are produced
2	What is the equation for calculating the mean rate of reaction?	$\text{mean rate} = \frac{\text{change in quantity of product or reactant}}{\text{time taken}}$
3	What is the unit for rate of reaction in a reaction involving a change in mass?	g/s
4	What is the unit for rate of reaction in a reaction involving a change in volume?	cm^3/s
5	What is the activation energy?	the minimum amount of energy colliding particles have to have before a reaction will take place
6	What effect does increasing concentration have on the rate of reaction?	increases
7	Why does increasing concentration have this effect?	more reactant particles in the same volume lead to more frequent collisions
8	What effect does increasing pressure have on the rate of reaction?	increases
9	Why does increasing pressure have this effect?	less space between particles means more frequent collisions
10	What effect does increasing surface area have on the rate of reaction?	increases
11	Why does increasing surface area have this effect?	more reactant particles are exposed and able to collide, leading to more frequent collisions
12	What effect does increasing temperature have on the rate of reaction?	increases
13	Why does increasing temperature have this effect?	particles move faster, leading to more frequent collisions – particles have the same activation energy, so more collisions result in a reaction
14	What is a catalyst?	a substance that increases the rate of a reaction but is not used up in the reaction
15	How do catalysts increase the rate of a reaction?	lower the activation energy of the reaction, so more collisions result in a reaction
16	What is a reversible reaction?	the reactants turn into products and the products turn into reactants
17	Which symbol shows a reversible reaction?	$\rightleftharpoons$
18	What is dynamic equilibrium?	the point in a reversible reaction when the rate of the forward and reverse reactions are the same
19	What are the three reaction conditions that can be changed?	concentration, temperature, pressure
20	What is Le Châtelier's principle?	the position of equilibrium will shift to oppose external changes
21	What is the effect of increasing the concentration of reactants on a reaction at dynamic equilibrium?	favours the forward reaction
22	What is the effect of increasing the concentration of reactants on a reaction at dynamic equilibrium?	favours the forward reaction
23	What is the effect of decreasing the concentration of products on a reaction at dynamic equilibrium?	favours the forward reaction
24	What is the effect of increasing pressure on a reaction at dynamic equilibrium?	favours the reaction that leads to the fewest molecules
25	What is the effect of decreasing pressure on a reaction at dynamic equilibrium?	favours the reaction that leads to the most molecules
26	What is the effect of increasing temperature on a reaction at dynamic equilibrium?	favours the endothermic reaction
27	What is the effect of decreasing temperature on a reaction at dynamic equilibrium?	favours the exothermic reaction

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Chapter 9: Crude oils and fuels

Knowledge organiser

Crude oil

Crude oil is incredibly important to our society and economy. It is formed from the remains of ancient biomass – living organisms (mostly plankton) that died many millions of years ago.

Raw crude oil is a thick black liquid made of a large number of different compounds mixed together. Most of the compounds are **hydrocarbons** of various sizes. Hydrocarbons are molecules made of carbon and hydrogen only.

Combustion

Hydrocarbons are used as **fuels**. This is because when they react with oxygen they release a lot of energy. This reaction is called **combustion**. Complete combustion is a type of combustion where the only products are carbon dioxide and water.

Properties

Whether or not a particular hydrocarbon is useful as a fuel depends on its properties:

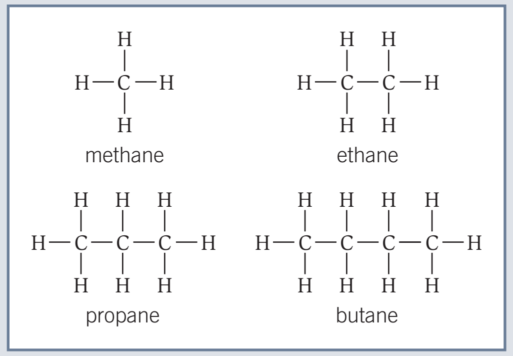
- **flammability** – how easily it burns
- **boiling point** – the temperature at which it boils
- **viscosity** – how thick it is

Its properties in turn depend on the length of the molecule.

Chain length	Flammability	Boiling point	Viscosity
long chain	low	high	high (very thick)
short chain	high	low	low (very runny)

Alkanes

One family of hydrocarbon molecules are called **alkanes**. Alkane molecules only have single bonds in them. The first four alkanes are:



The different alkanes have different numbers of carbon atoms and hydrogen atoms. You can always work the molecular formula of an alkane by using C_nH_{2n+2}.

Key terms

Make sure you can write a definition for these key terms.

alkanes · alkenes · boiling point · combustion · cracking · crude oil · feedstock · flammability · fractional distillation · fuel · hydrocarbon · viscosity

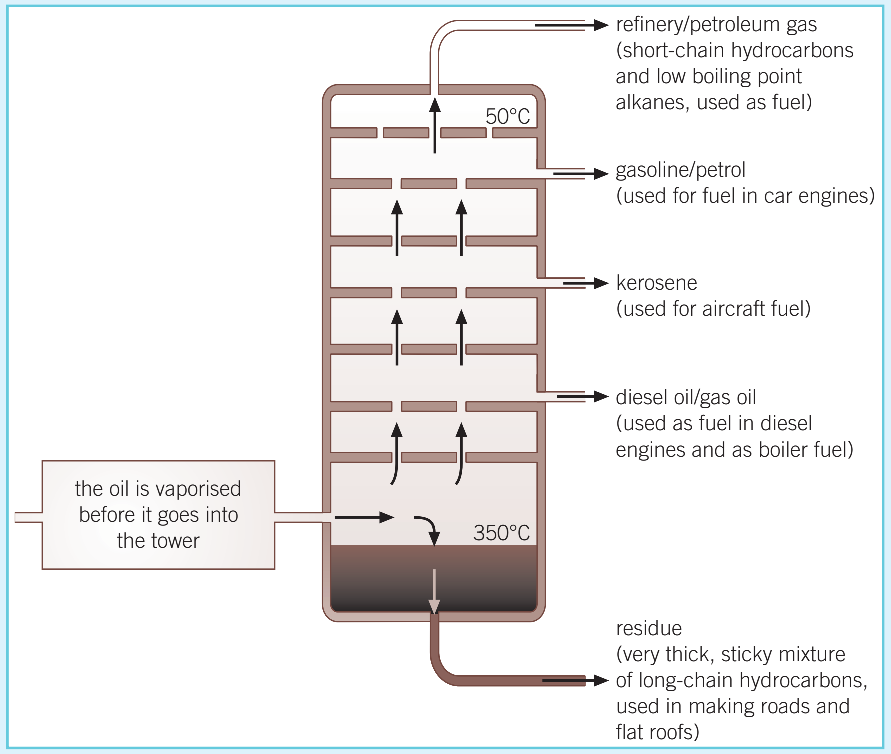
Fractional distillation

The different hydrocarbons in crude oil are separated into fractions based on their boiling points in a process called **fractional distillation**. All the molecules in a fraction have a similar number of carbon atoms, and so a similar boiling point.

The process takes place in a fractionating column, which is hot at the bottom and cooler at the top.

The process works like this:

1. crude oil is vapourised (turned into a gas by heating)
2. the hydrocarbon gases enter the column
3. the hydrocarbon gases rise up the column
4. as hydrocarbon gases rise up the column they cool down
5. when the different hydrocarbons reach their boiling point in the column they condense
6. the hydrocarbon fraction is collected.



Products from fractional distillation

Many useful products come from the separation of crude oil by fractional distillation.

Fuels	Feedstock	Useful materials produced
petrol, diesel oil, kerosene, heavy fuel oil, and liquefied petroleum gases	fractions form the raw material for other processes and the production of other substances	solvents, lubricants, polymers, and detergents

Cracking

Not all hydrocarbons are as useful as each other. Longer molecules tend to be less useful than shorter ones. As such, there is a higher demand for shorter-chain hydrocarbons than longer-chain hydrocarbons.

A process called **cracking** is used to break up longer hydrocarbons and turn them into shorter ones.

Cracking produces shorter alkanes and **alkenes**.

Two methods of cracking are:

- catalytic cracking – vaporise the hydrocarbons, then pass them over a hot catalyst
- steam cracking – mix the hydrocarbons with steam at a very high temperature

Alkenes

Alkenes are a family of hydrocarbons that contain double bonds between carbon atoms.

Alkenes are also used as fuels, and to produce polymers and many other materials.

They are much more reactive than alkanes. When mixed with bromine water, the bromine water turns from orange to colourless. This can be used to tell the difference between alkanes and alkenes.

Chapter 9: Crude oil and fuels

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C9 questions	Answers
1	What is a hydrocarbon?	compound containing carbon and hydrogen only
2	How is crude oil formed?	over millions of years from the remains of ancient biomass
3	What are the alkanes?	hydrocarbons that only have single bonds
4	What are the first four alkanes?	methane, ethane, propane, butane
5	What is the general formula for the alkanes?	C_nH_{2n+2}
6	How does boiling point depend on the chain length?	longer the chain, higher the boiling point
7	How does viscosity depend on chain length?	longer the chain, higher the viscosity
8	How does flammability depend on chain length?	longer the chain, lower the flammability
9	How can the different alkanes in crude oil be separated?	fractional distillation
10	What is a fraction?	a group of hydrocarbons with similar chain lengths
11	Name five useful fuels produced from fractional distillation.	petrol, diesel oil, kerosene, heavy fuel oil, and liquefied petroleum gases
12	Name four useful materials produced from crude oil fractions.	solvents, lubricants, polymers, detergents
13	What is cracking?	breaking down a hydrocarbon with a long chain into smaller molecules
14	Name two methods to carry out cracking.	steam cracking and catalytic cracking
15	What are the products of cracking?	short chain alkanes and alkenes
16	What are alkenes?	hydrocarbons with a double bond
17	What are alkenes used for?	formation of polymers
18	Describe the reactivity of alkenes compared to alkanes.	alkenes are much more reactive
19	How can you test for alkenes?	alkenes turn orange bromine water colourless

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Chapter 10: Chemical analysis

Knowledge organiser

Pure and impure

In chemistry, a **pure** substance contains a single element or compound that is not mixed with any other substance.

Pure substances melt and boil at specific temperatures.

An **impure** substance contains more than one type of element of compound in a **mixture**.

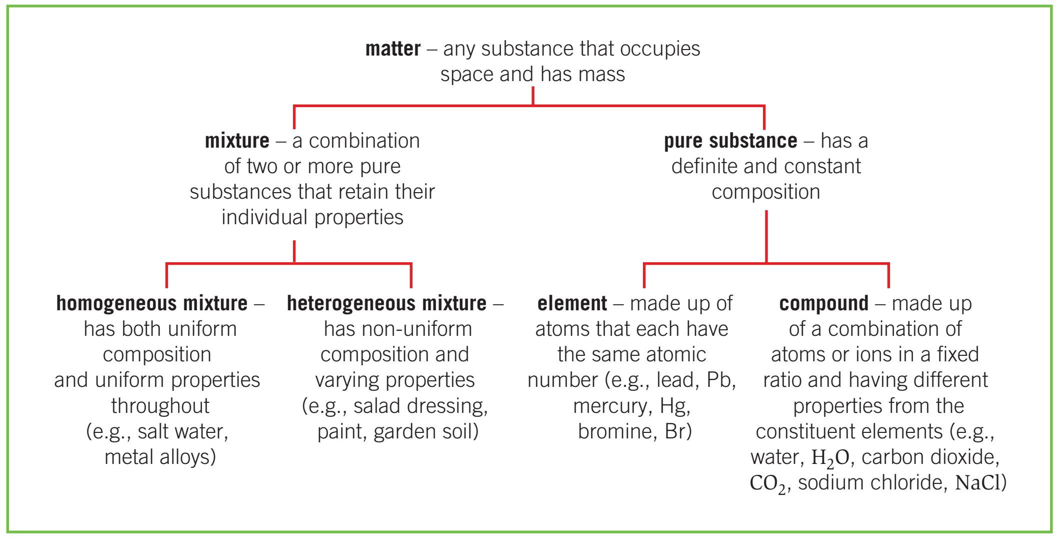
Impure substances melt and boil at a range of temperatures.

Formulations

Formulations are examples of mixtures. They have many different components (substances that make them up) in very specific proportions (amounts compared to each other).

Scientists spend a lot of time trying to get the right components in the right proportions to make the most useful product.

Formulations include fuels, cleaning agents, paints, alloys, fertilisers, and foods.



Testing gases

Common gases can be identified using the follow tests:

Gas	What you do	What you observe if gas is present
hydrogen	hold a lighted splint near the gas	hear a squeaky pop
oxygen	hold a glowing splint near the gas	splint re-lights
carbon dioxide	bubble the gas through limewater	the limewater turns milky (cloudy white)
chlorine	hold a piece of damp litmus near the gas	bleaches the litmus white

Chromatography

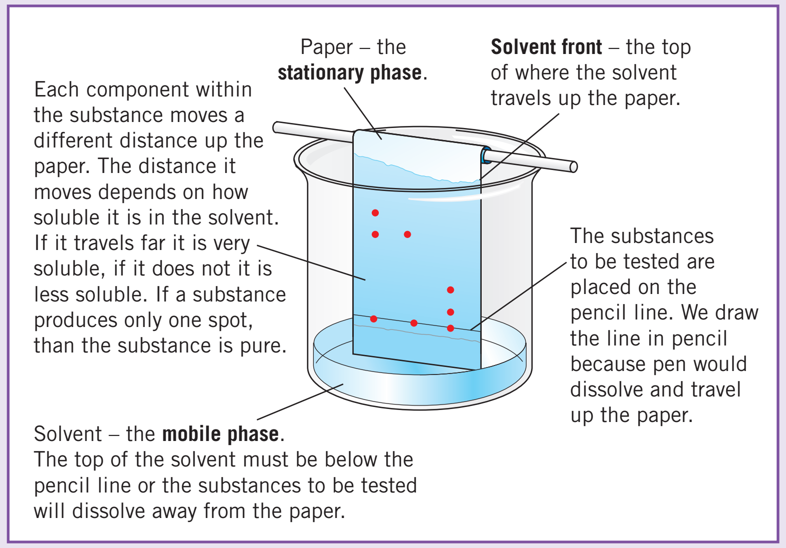
Chromatography is a method to separate different components in a mixture. It is set up as shown here, with a piece of paper in a beaker containing a small amount of solvent.

The **R_f value** is a ratio of how far up the paper a certain spot moves compared to how far the **solvent** has travelled.

$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

It will always be a number between 0 and 1.

The R_f value depends on the solvent and the temperature, and different substances will have different R_f values. The R_f values for particular solvents can be used to identify a substance.



Separating mixtures

Mixtures can be separated by

- filtration – separates insoluble solids from a liquid
- crystallisation – evaporates a solvent (liquid) leaving the solute (solid)
- simple distillation – separates solvent from a solution as long as the solvent has a lower boiling point than the solute
- fractional distillation – separates two or more substances from a liquid mixture.

Key terms

Make sure you can write a definition for these key terms.

chromatography formulation impure mobile phase precipitate pure R_f value solvent solvent front stationary phase

Chapter 10: Chemical analysis

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C10 questions	Answers
1	In chemistry, what is a pure substance?	something made of only one type of substance
2	What is the difference between the melting and boiling points of a pure and impure substance?	pure – sharp/one specific temperature impure – broad/occurs across a range of temperatures
3	What is a formulation?	a mixture designed for a specific purpose
4	What are some examples of formulations?	fuels, cleaning agents, paints, medicines, alloys, fertilisers, and foods
5	What is chromatography?	a process for separating coloured mixtures
6	How is R_f calculated?	$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$
7	What is the test for hydrogen?	a lit splint gives squeaky pop
8	What is the test for oxygen?	re-lights a glowing splint
9	What is the test for carbon dioxide?	turns limewater milky if bubbled through it
10	What is the test for chlorine?	bleaches damp litmus paper

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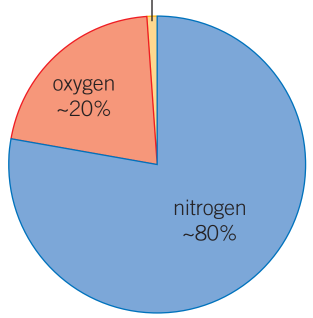
Chapter 11: The Earth's atmosphere

Knowledge organiser

The Earth's changing atmosphere

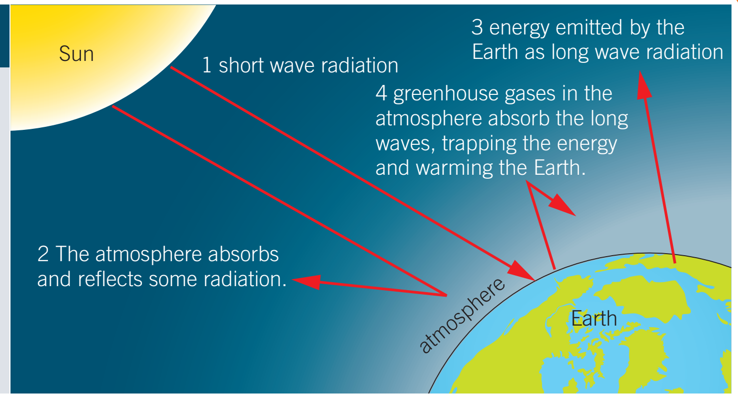
Period	Proportions of gases	Evidence
about 4.6 billion years to about 2.7 billion years ago	• **carbon dioxide, CO_2** Released by volcanoes. Biggest component of the **atmosphere**. • **oxygen, O_2** Very little oxygen present. • **nitrogen, N_2** Released by volcanoes. • **water vapour, H_2O** Released by volcanoes. Existed as vapour as Earth was too hot for it to condense. • **other gases** Ammonia, NH_3, and methane, CH_4, may also have been present.	Because it was billions of years ago there is very little evidence to draw upon.
about 2.7 billion years ago to about 200 million years ago	• **carbon dioxide, CO_2** Amount in atmosphere begins to reduce because: ◦ water condenses to form the oceans, in which CO_2 then dissolves ◦ algae (and later plants) start to photosynthesise carbon dioxide + water $\xrightarrow{\text{light}}$ glucose + oxygen $6CO_2 + 6H_2O \longrightarrow C_6H_{12}O_6 + 6O_2$ ◦ CO_2 precipitates in the oceans as solid carbonates (sediments) that form rocks ◦ CO_2 taken in by plants and animals. When they die, the carbon in them is locked up as fossil fuels • **oxygen, O_2** Starts to increase as a product of photosynthesis. • **nitrogen, N_2** Continues to increase. Nitrogen is a very stable molecule so any process that produces it causes the overall amount to build up over time. • **water vapour, H_2O** Starts to decrease. As the Earth cools, the vapour condenses and forms the oceans.	Still limited as billions of years ago, but can look at processes that happen today (like photosynthesis) and make theories about the past.
about 200 million years ago until the present	• **carbon dioxide, CO_2** about 0.04% • **oxygen, O_2** about 20% • **nitrogen, N_2** about 80% • **water vapour, H_2O** Very little overall. Collects in large clouds as part of the water cycle. • **other gases** Small proportions of other gases such as the noble gases.  	Ice core evidence for millions of years ago and lots of global measurements taken recently.

Greenhouse gases

Greenhouse gases, such as carbon dioxide, methane, and water vapour, absorb radiation and maintain temperatures on the Earth to support life.

However, in the last 150 years, more greenhouse gases have been released due to human activities.

- carbon dioxide – combustion of fossil fuels, deforestation
- methane – planting rice fields, cattle farming



Global warming

Scientists have gathered peer-reviewed evidence to demonstrate that increasing the amount of greenhouse gases in the atmosphere will increase the overall average temperature of the Earth. This is called **global warming**.

However, it is difficult to make predictions about the atmosphere as it is so big and complex. This leads some people to doubt what scientists say.

Global climate change

Global warming leads to another process called **global climate change** – how the overall weather patterns over many years and across the entire planet will change.

There are many different effects of climate change, including:

- sea levels rising
- extreme weather events
- changes in the amount and time of rainfall
- changes to ecosystems and habitats
- polar ice caps melting.

Carbon footprints

Increasing the amount of greenhouse gases in the atmosphere increases the global average temperature of the Earth, which results in global climate change.

As such, it is important to reduce the release of greenhouse gases into the atmosphere. The amount of carbon dioxide and methane that is released into the atmosphere by a product, person, or process is called its **carbon footprint**.

Other pollutants released in combustion of fuels

Pollutant	Origin	Effect
carbon monoxide	incomplete combustion of fuels	colourless and odourless toxic gas
particulates (soot and unburnt hydrocarbons)	incomplete combustion of fuels especially in diesel engines	**global dimming**, respiratory problems, potential to cause cancer
sulfur dioxide	sulfur impurities in the fuel reacting with oxygen from the air	**acid rain** and respiratory problems
oxides of nitrogen	nitrogen from the air being heated near an engine and reacting with oxygen	acid rain and respiratory problems

Key terms

Make sure you can write a definition for these key terms.

acid rain | atmosphere | carbon footprint | global climate change | carbon monoxide | global dimming | global warming | greenhouse gas | particulate | pollutant

Chapter 11: The Earth's atmosphere

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C11 questions	Answers
1	What is the atmosphere?	a layer of gas surrounding the Earth
2	What was the early atmosphere composed of?	mostly carbon dioxide
3	How did the oceans form?	water vapour condensing as the Earth cooled
4	How did the amount of carbon dioxide in the atmosphere decrease to today's levels?	dissolved in the oceans, photosynthesis, converted to fossil fuels, precipitated as insoluble metal carbonates
5	When did life start to appear, and what was the impact of this on oxygen in the atmosphere?	about 2.7 billion years ago; amount of atmospheric oxygen increased as it was released in photosynthesis
6	How has the amount of nitrogen in the atmosphere changed over time?	increased slowly as it is a very stable molecule
7	Why can scientists not be sure about the composition of the Earth's early atmosphere?	it was billions of years ago and evidence is limited
8	What is the current composition of the atmosphere?	approximately 80% nitrogen, 20% oxygen, and trace amounts of other gases such as carbon dioxide, water vapour, and noble gases
9	What is a greenhouse gas?	a gas that traps radiation from the Sun
10	What type of radiation do greenhouse gases absorb?	longer wavelength infrared radiation
11	Name three greenhouse gases.	methane, carbon dioxide, water vapour
12	Give two ways recent human activities have increased the amount of atmospheric carbon dioxide.	burning fossil fuels, deforestation
13	Give two ways recent human activities have increased the amount of atmospheric methane.	rice farming, cattle farming
14	What is global warming?	an increase in the overall global average temperature
15	What is global climate change?	the change in long-term weather patterns across the planet
16	What are some possible effects of climate change?	sea levels rising, extreme weather events, changes in the amount and time of rainfall, changes to ecosystems and habitats, polar ice caps melting
17	What is a carbon footprint?	the amount of carbon a product, process, or person releases into the atmosphere over its lifetime
18	How is carbon monoxide formed, and what is the danger associated with it?	incomplete combustion; colourless and odourless toxic gas
19	How are particulates formed, and what are the dangers associated with them?	incomplete combustion; global dimming, respiratory problems, potential to cause cancer
20	How is sulfur dioxide formed, and what are the dangers associated with it?	sulfur impurities in fossil fuels react with oxygen during combustion; acid rain, respiratory problems
21	How are oxides of nitrogen formed, and what are the dangers associated with them?	atmospheric oxygen and nitrogen react in the heat of a combustion engine; acid rain, respiratory problems

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Chapter 12: The Earth's resources 1

Knowledge organiser

Natural and synthetic resources

We use the Earth's resources to provide us with warmth, fuel, shelter, food, and transport.

- Natural resources are used for food, timber, clothing, and fuels.
- Synthetic resources are made by scientists. They can replace or supplement natural resources.

When choosing and synthesising resources, it is important to consider **sustainable development**. This is development that meets the needs of current generations without compromising the ability of future generations to meet their own needs.

Finite and renewable resources

Some resources are **finite**. This means that they will eventually run out.

Fossil fuels are an example of a finite resource. They take so long to form that we use them faster than they are naturally formed.

Resources that will not run out are called **renewable** resources.

Wood is an example of a renewable resource. Trees can be grown to replace any that are cut down for wood.

Potable water

Water is a vital resource for life. **Potable** water is water that is safe to drink. However, most water on Earth is not potable.

Type of water	What it has in it
pure water	just water molecules and nothing else
potable water	water molecules, low levels of salts, safe levels of harmful microbes
salty water (sea water)	water molecules, dangerously high levels of salt, can have high levels of harmful microbes
fresh water (from rivers, lakes, or underground)	water molecules, low levels of salt, often has harmful microbes at high levels

Fresh water

In the UK, potable water is produced from rain water that collects in lakes and rivers. To produce potable water:

1. Choose an appropriate source of fresh water.
2. Pass the water through filters to remove large objects.
3. **Sterilise** the water to kill any microbes using ozone, chlorine, or UV light.

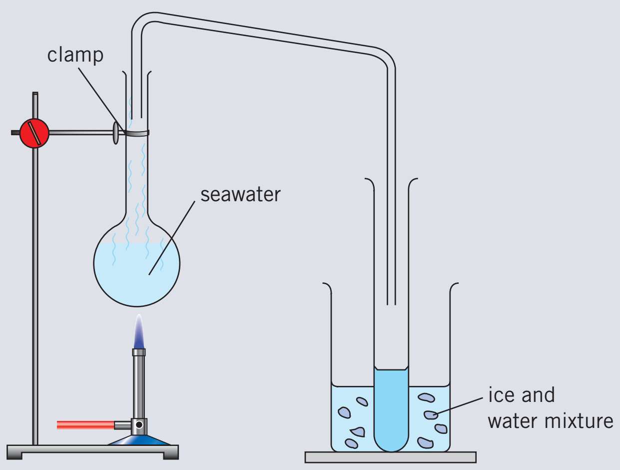
Salty water

Some countries do not have lots of fresh water available. **Desalination** is the process to turn saltwater into potable water. This requires a lot of energy and can be done by:

- distillation
- **reverse osmosis**

Reverse osmosis involves using membranes to separate the salts dissolved in the water. The water needs to be pressurised and the salty water corrodes the pumps. As such, it is an expensive process.

Distillation



Waste water

Human activities produce lots of waste water as sewage, agricultural waste, and industrial waste.

- **Sewage** and agricultural waste contain organic matter and harmful microbes.
- Industrial waste contains organic matter and harmful chemicals.

These need to be removed before the water can be put back into the environment.

Treating sewage water

screening and grit removal
The sewage passes through a metal grid that filters out large objects.

sedimentation
The sewage is left so that solid sediments settle out of the water. The sediments sink to the bottom of the tank. The liquid sits above the sediment.

Treating sludge

sewage sludge
This sediment is called **sludge**. Sludge contains organic matter, water, dissolved compounds, and small solid particles.

anaerobic treatment
Bacteria are added to digest the organic matter. These bacteria break down the matter anaerobically – with a limited supply of oxygen.

biogas
The anaerobic digestion of sludge produces biogas. Biogas is a mixture of methane, carbon dioxide and hydrogen sulfide. It can be used as fuel.

remaining sludge used as fuel
The remaining sludge can be dried out and can also be burnt as a fuel.

Treating effluent

effluent
The remaining liquid is called **effluent**. This effluent has no solid matter visible, but still contains some matter and harmful microorganisms.

aerobic treatment
Bacteria are added to the effluent. These bacteria feed on organic matter and the harmful microorganisms in the effluent. The bacteria break down the matter by aerobic respiration – oxygen needs to be present.

bacteria removed
The bacteria are allowed to settle out of the water.

discharged back to rivers
The water is now safe enough to be released back into the environment.

Knowledge organiser

Metal extraction (HT only)

Metals are used for many different things. Some metals can be extracted from their ores by reduction or electrolysis.

However, metal ores are a finite resource and these processes require lots of energy.

Scientists are looking for new ways to extract metals that are more sustainable.

Phytomining and **bioleaching** are two alternative processes used to extract copper from low grade ores (ores with only a little copper in them).

Phytomining

1. Grow plants near the metal ore.
2. Harvest and burn the plants.
3. The ash contains the metal compound.
4. Process the ash by electrolysis or displacement with scrap metal.

Bioleaching

1. Grow bacteria near the metal ore.
2. Bacteria produce leachate solutions that contain metal compound.
3. Process the leachate by electrolysis or displacement with scrap metal.

Both of these methods avoid the digging, moving, and disposing of large amounts of rock associated with traditional mining techniques.

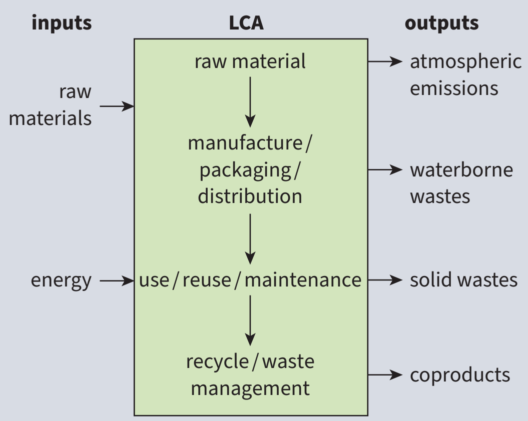
Life cycle assessment

A **life cycle assessment (LCA)** is a way of looking at the whole life of a product and assessing its impact on the environment and sustainability. It is broken down into four categories:

- extracting and processing raw materials
- manufacturing and packaging
- use and operation during its lifetime
- disposal at the end of its useful life, including transport and distribution at each stage

Some parts of an LCA are objective, such as the amount of water used or waste produced in the production of a product.

However, other parts of an LCA require judgements, such as the polluting effect, and so LCAs are not a completely objective process.



Key terms

Make sure you can write a definition for these key terms.

aerobic *anaerobic* *biodegrade* *bioleaching* *distillation* *effluent* *finite resources* *life cycle assessment* *phytomining* *potable water* *recycling* *renewable resources* *reverse osmosis* *screening* *sedimentation* *sewage* *sludge* *sterilisation* *sustainable development*

Disposal of products

When someone finishes with a product, it can be

- added to a landfill
  This can cause habitat loss and other problems in the local ecosystem. Some items persist in landfills as they do not **biodegrade** and could be there for hundreds of years.
- incinerated
  Some products can be incinerated to produce useful energy. However, the combustion can often be incomplete and result in harmful pollutants being released to the atmosphere.
- reused
  This is when an item is used again for a similar purpose.
- **recycled**
  Recycling requires energy, but conserves the limited resources and often requires less energy than needed to make brand new materials.

The table shows information about the extraction, processing, and disposal of some common materials. This information is used when making a LCA.

Material	Extraction/processing	Disposal
metal	• quarrying and mining cause habitat loss • machinery involved in mining release greenhouse gases • extraction from metal ores require lots of energy	• metals can normally be recycled by melting them down and then casting them into new shapes • metals in landfill can persist for a long time
plastic	normally come from fossil fuels that are non-renewable	• many plastic products can be reused and recycled • plastics often end up in landfills where they persist as they are not biodegradable • incinerating plastics releases lots of harmful pollutants like carbon monoxide and particulates
paper	produced from trees that require land and lots of water to grow lots of water also used in the production process	• many paper products can be recycled • paper products can also be incinerated or they can decay naturally in a landfill • incineration and decay release greenhouse gases
glass	produced by heating sand, which requires a lot of energy	• many glass products can be reused, or crushed and recycled • if glass is added to landfills it persists as it is not biodegradable
ceramics	• come from clay and rocks • generally require quarrying, which requires energy, releases pollutants from heavy machinery, and causes habitat loss	• most ceramics are not commonly recycled in the UK, and once broken cannot be reused • ceramics tend to persist in landfills

Chapter 12: The Earth's resources

Retrieval questions

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many as you can. Check and repeat.

	C12 questions	Answers
1	What do we use the Earth's resources for?	warmth, shelter, food, fuel, transport
2	What are some examples of natural resources?	cotton, wool, timber
3	What are some examples of synthetic resources?	plastic, polyester, acrylic
4	What is a finite resource?	a resource that will eventually run out
5	What is sustainable development?	development that meets the needs of current generations without compromising the ability of future generations to meet their own needs
6	What are the four main types of water?	pure water, salt water, fresh water, potable water
7	What is potable water?	water that is safe to drink
8	In the UK, how is potable water extracted from fresh water?	filtration and sterilisation
9	What is sterilisation?	killing microbes
10	What are three examples of sterilising agents?	chlorine gas, UV light, and ozone
11	How can potable water be produced from salt water?	desalination
12	How can desalination be carried out?	distillation or reverse osmosis
13	What are the three main types of waste water?	sewage, agricultural waste, industrial waste
14	What can waste water contain?	organic matter, harmful microbes, harmful chemicals
15	What is the first step in processing waste water?	screening and grit removal
16	What is sedimentation?	separating the waste water into sludge and effluent
17	How is sludge treated?	anaerobic respiration
18	How is effluent treated?	aerobic respiration
19	What is phytomining?	using plants to extract copper
20	What is bioleaching?	using bacteria to extract copper
21	What is a life cycle assessment?	a way of assessing the energy costs and environmental effect of a product across its lifetime
22	What are the four stages of a life cycle assessment?	• extracting and processing raw materials • manufacturing and packaging • use and operation during its lifetime • disposal at the end of its useful life
23	How can we reduce the amount of new materials manufactured?	by reducing, reusing, or recycling products
24	In what ways can materials that are not recycled be disposed?	landfill or incineration

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