Course Companion

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Atomic structure and the periodic table

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Atomic structure and the periodic table: evidence and predictions

Why it matters. In 1871 the Russian chemist Dmitri **Mendeleev** 门捷列夫 published a periodic table with gaps in it. Under aluminium he left a space for an element nobody had found, which he called "eka-aluminium". He predicted that it would be a metal with a relative atomic mass of about 68, a density of about 5.9 g/cm^3 and a low melting point. In 1875 the French chemist Paul-Émile Lecoq de Boisbaudran discovered a new metal, gallium, and measured its density as 4.7 g/cm^3. Mendeleev wrote to tell him that his sample must be impure. Lecoq de Boisbaudran purified it, measured again, and found 5.9. A table that could predict a new element before anyone had seen it was clearly more than a list. This topic explains what atoms are made of, how we found out, and why the periodic table works.

In this unit you will:

- tell elements, compounds and mixtures apart, write formulae and balance equations;
- choose a method to separate a mixture;
- explain how evidence changed the model of the atom;
- count the protons, neutrons and electrons in atoms, ions and isotopes, and calculate relative atomic mass;
- link electron arrangement to position in the periodic table, and explain the trends in Groups 0, 1 and 7;
- (chemistry only) compare the transition metals with Group 1.

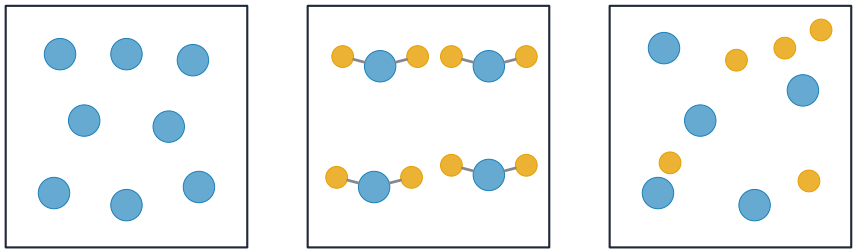
AQA GCSE Chemistry (8462) topic 4.1, also in Combined Science: Trilogy (8464) except section 4.1.3 on transition metals. It is tested on Paper 1. Higher tier (HT) content is labelled.

Elements, compounds and equations

An **atom** 原子 is the smallest part of an **element** 元素 that can exist. All the atoms of one element have the same number of protons. An element can be made of single atoms, like helium, or of molecules, like oxygen, O_2: two atoms, but still only one element.

A **compound** 化合物 contains two or more elements chemically combined in fixed proportions. Water, H_2O, always has two hydrogen atoms for every oxygen atom. A compound has different properties from its elements: sodium is a soft metal that reacts violently with water, and chlorine is a poisonous green gas, but sodium chloride is the salt you put on food. A compound can be separated into its elements only by

a **chemical reaction** 化学反应, which makes new substances.



element
only one type of atom

compound
different atoms bonded in a fixed ratio

mixture
not bonded, easy to separate

An element has one type of atom; a compound has different atoms bonded in a fixed ratio; a mixture is not bonded

Learn the names and symbols of the first twenty elements, and write symbols with the right capital letters: Co is cobalt, but CO is carbon monoxide.

Atomic number	Name and symbol	Atomic number	Name and symbol
1	hydrogen H	11	sodium Na
2	helium He	12	magnesium Mg
3	lithium Li	13	aluminium Al
4	beryllium Be	14	silicon Si
5	boron B	15	phosphorus P
6	carbon C	16	sulfur S
7	nitrogen N	17	chlorine Cl
8	oxygen O	18	argon Ar
9	fluorine F	19	potassium K
10	neon Ne	20	calcium Ca

Equations. A **word equation** 文字方程式 names the reactants and products: magnesium + oxygen → magnesium oxide. A **symbol equation** 符号方程式 uses formulae, and it must be balanced, with the same number of each kind of atom on both sides.

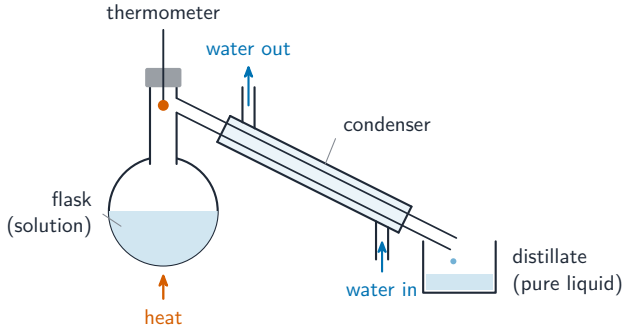
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Worked example 1: balancing. Start with $Mg + O_2 \rightarrow MgO$. There are 2 oxygen atoms on the left but only 1 on the right, so put a 2 in front of MgO. Now there are 2 Mg atoms on the right, so put a 2 in front of Mg: $2Mg + O_2 \rightarrow 2MgO$. Check: Mg 2 = 2, O 2 = 2. Never change a formula to balance an equation; MgO_2 would be a different substance.

Higher tier: ionic equations. A zinc atom loses two electrons to become a zinc ion: $Zn \rightarrow Zn^{2+} + 2e^-$. A copper ion gains two electrons: $Cu^{2+} + 2e^- \rightarrow Cu$. Adding these half equations, the electrons cancel: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$. Ions that do not change, such as sulfate in this reaction, are left out. Check that both atoms and total charge balance.

Separating mixtures

A **mixture** 混合物 contains substances that are not chemically joined, so each keeps its own properties. Separating a mixture makes no new substances. Choose the method from the property that differs between the parts.



Simple distillation: the solvent boils off, the condenser cools it back to a liquid, and the pure distillate is collected

Method	What it separates	Example
filtration 过滤	an insoluble solid from a liquid	sand from salt solution
crystallisation 结晶	a dissolved solid from its solution	salt from salt solution
simple distillation 简单蒸馏	a liquid from a solution	pure water from sea water
fractional distillation 分馏	liquids that mix, with different boiling points	ethanol from water; crude oil into fractions
chromatography 色谱法	dissolved substances that move at different speeds through paper	the dyes in an ink

Worked example 2: pure salt from rock salt. Rock salt is salt mixed with sand. Add water and stir: the salt dissolves but the sand does not. Filter: the sand stays in the filter paper and salt solution passes through. Heat the solution gently to evaporate some of the water, then leave it to cool and crystallise. Filter off the crystals and dry them. Filtration alone cannot collect the salt, because it is dissolved.

Distillation 蒸馏 separates a liquid by evaporating it and then condensing the vapour. **Worked example 3:** to collect the water from salt water, use simple distillation. The water boils, its vapour cools in the condenser and drips into the beaker as pure water, and the salt stays in the flask. In Saudi Arabia and other dry countries, large plants make drinking water from sea water in this way.

In chromatography, draw the start line in pencil, not ink, and keep it above the level of the solvent; otherwise the samples dissolve into the solvent instead of moving up the paper.

How evidence changed the model of the atom

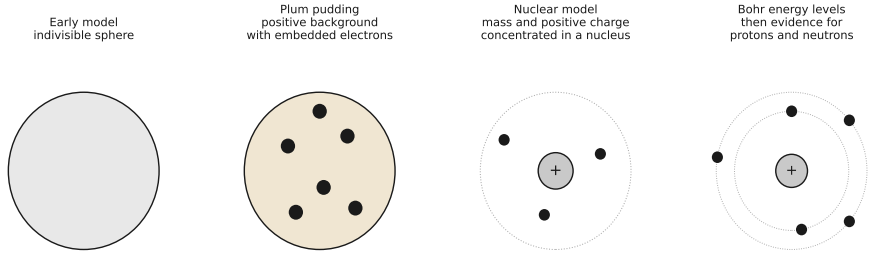
A scientific **model** 模型 explains the evidence we have. When new evidence does not fit, the model must change.

- **Around 1800**, John Dalton described atoms as tiny spheres that could not be divided.
- **In 1897**, J. J. Thomson discovered the electron, a tiny negative particle inside atoms. So atoms must have parts. He proposed the **plum pudding model** 布丁模型: a ball of positive charge with negative electrons scattered through it.
- **In 1909**, Hans Geiger and Ernest Marsden, working for Ernest Rutherford in Manchester, fired positive alpha particles at a sheet of gold foil only a few hundred atoms thick.

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What they saw	What it showed
most alpha particles went straight through	most of the atom is empty space
some were deflected	the positive charge is concentrated in one place and repels them
about 1 in 8,000 bounced back	the centre is tiny, dense and holds most of the mass

Rutherford said it was "as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you". In 1911 he proposed the **nuclear model** 核式模型: a tiny positive nucleus with electrons around it. Niels Bohr then showed (1913) that the electrons move in fixed shells, or energy levels, at particular distances from the nucleus. Later experiments showed that the positive charge of the nucleus is made of smaller particles, protons. In 1932 James Chadwick found the neutron, the uncharged particle in the nucleus.



The atom model timeline: indivisible sphere, plum pudding, nuclear, Bohr, then proton and neutron.

Worked example 4: using all the evidence. A student says, "Some alpha particles bounced back, so the atom is solid." This ignores the fact that most went straight through. Only a model with a tiny, dense nucleus inside a mostly empty atom explains both observations. A good model must fit all the evidence, not just one result.

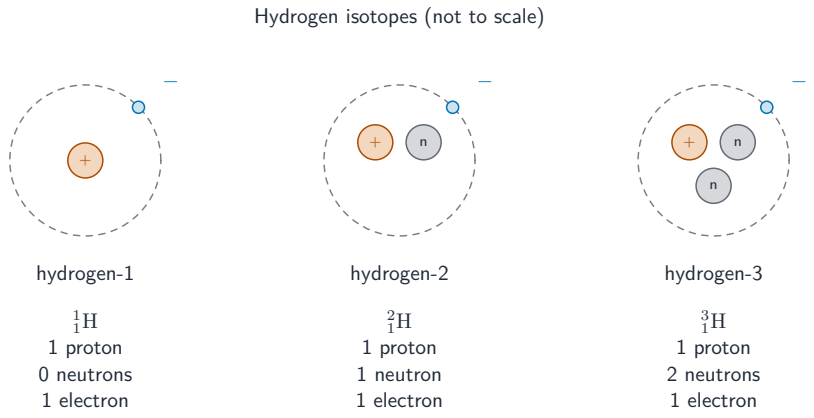
Inside the atom: particles, size and isotopes

Particle	Relative mass	Relative charge	Where it is
proton 质子	1	+1	nucleus
neutron 中子	1	0	nucleus
electron 电子	very small	−1	shells around the nucleus

The **atomic number** 原子序数 is the number of protons. It decides which element an atom is. The **mass number** 质量数 is the number of protons plus neutrons. In an atom the numbers of protons and electrons are equal, so the atom has no overall charge. An ion forms when an atom gains or loses electrons; the number of protons never changes.

> **Worked example 5: counting particles.** Chlorine-37 has atomic number 17 and mass number 37. It has 17 protons, $37 - 17 = 20$ neutrons and 17 electrons. A chloride ion, Cl^-, has gained one electron, so it has 18 electrons but still 17 protons and 20 neutrons. A calcium ion, Ca^{2+} (atomic number 20, mass number 40), has 20 protons, 20 neutrons and 18 electrons.

Isotopes. An **isotope** 同位素 is one of two or more forms of an element with the same number of protons but a different number of neutrons. Chlorine-35 has 18 neutrons and chlorine-37 has 20, but both are chlorine because both have 17 protons. Hydrogen has three isotopes.



Isotopes differ only in the number of neutrons

Size. An atom has a radius of about 0.1 nm, or 1×10^{-10} m. The nucleus has a radius less than 1/10,000 of the atom's radius, but it contains almost all the mass. If an atom were the size of a football stadium, its nucleus would be about the size of a pea on the centre spot.

Relative atomic mass. The **relative atomic mass** 相对原子质量 (Ar) of an element is the mean mass of its atoms, taking into account how common each isotope is:

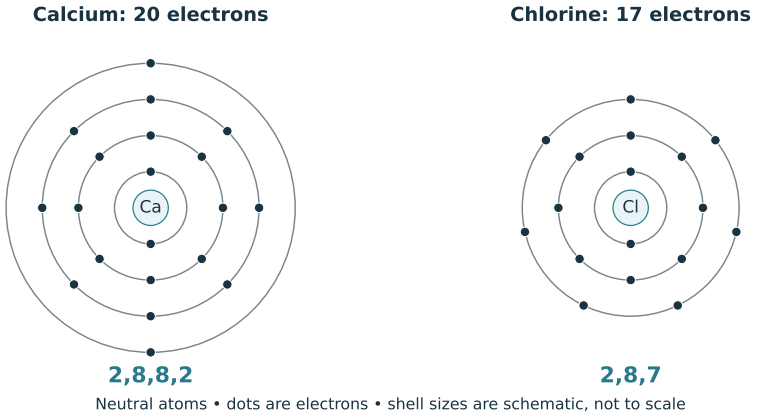
$$A_r = \frac{\sum(\text{mass number} \times \text{percentage abundance})}{100}$$

> **Worked example 6.** Chlorine is 75% chlorine-35 and 25% chlorine-37. $A_r = (35 \times 75 + 37 \times 25)/100 = (2625 + 925)/100 = 35.5$. The answer is closer to 35 because chlorine-35 is more common. No single chlorine atom has a mass of 35.5; it is an average.

Electrons and the periodic table

Electrons fill the lowest **energy levels** 能级, or shells, first. For the first twenty elements, the first shell holds up to 2 electrons, the second up to 8, and the third up to 8 before the fourth shell begins. An **outer shell electron** 最外层电子 is one in the highest occupied shell; these electrons decide how an element reacts.

> **Worked example 7.** Calcium has 20 electrons: 2 in the first shell, 8 in the second, 8 in the third and 2 in the fourth, written 2,8,8,2. Chlorine has 17 electrons: 2,8,7.



Complete neutral calcium/chlorine shell diagrams;first-twenty model.

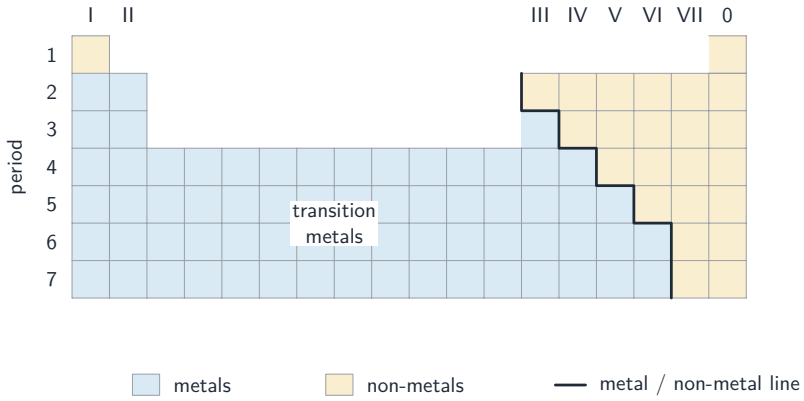
The modern **periodic table** 元素周期表 puts the elements in order of atomic number. A row is a **period** 周期: the number of occupied shells tells you the period. A column is a **group** 族: for Groups 1 to 7, the group number is the number of outer electrons. Elements in the same group have similar reactions because they have the same number of outer electrons. Group 0 elements have full outer shells: 2 electrons for helium, 8 for the others.

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Worked example 8. An element has the arrangement 2,8,1. It has 11 electrons, so 11 protons: it is sodium. It has three shells, so it is in Period 3, and one outer electron, so it is in Group 1. You can predict that it forms a +1 ion and reacts like lithium and potassium.

Building the periodic table; metals and non-metals

In the early 1800s, chemists arranged the elements in order of atomic weight, since protons were not yet known. These early tables put some elements next to others with very different properties. Mendeleev solved this in two ways: he left gaps for elements not yet discovered, and in a few places he changed the order to keep elements with similar properties in the same group. When gallium (1875) and germanium (1886) were found with the properties he had predicted, scientists accepted his table. Later, isotopes explained why ordering by atomic weight did not always work, and the modern table is ordered by atomic number.



Groups are the columns and periods are the rows; metals lie to the left of the staircase, non-metals to the right

Most elements are metals; they are on the left of the table and towards the bottom. Metals form positive ions when they react. Non-metals, on the right and towards the top, do not form positive ions. Metals usually conduct heat and electricity, are shiny, and can be bent or hammered into shape. Solid non-metals are usually dull, brittle and poor conductors, although graphite, a form of carbon, conducts electricity.

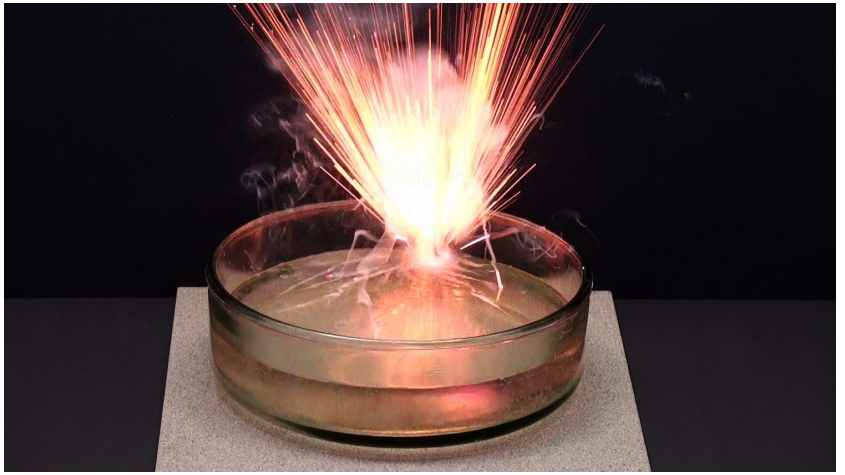
Groups 0, 1 and 7

Group 0. Each **noble gas** 稀有气体, such as helium, neon or argon, is very unreactive because its outer shell is full. The noble gases exist as single atoms. Their boiling points increase down the group as the relative atomic mass increases.

Group 1. Each **alkali metal** 碱金属 has one outer electron and forms a +1 ion. The alkali metals react with oxygen to form oxides, with chlorine to form chlorides, and with water to form a hydroxide and hydrogen:

$$4Li + O_2 \rightarrow 2Li_2O \qquad 2Na + Cl_2 \rightarrow 2NaCl \qquad 2K + 2H_2O \rightarrow 2KOH + H_2$$

In water, lithium fizzes steadily, sodium melts into a silvery ball that races across the surface, and potassium bursts into a lilac flame. Reactivity increases down the group. The outer electron is further from the nucleus and more shielded by the inner shells, so it is attracted less strongly and lost more easily.



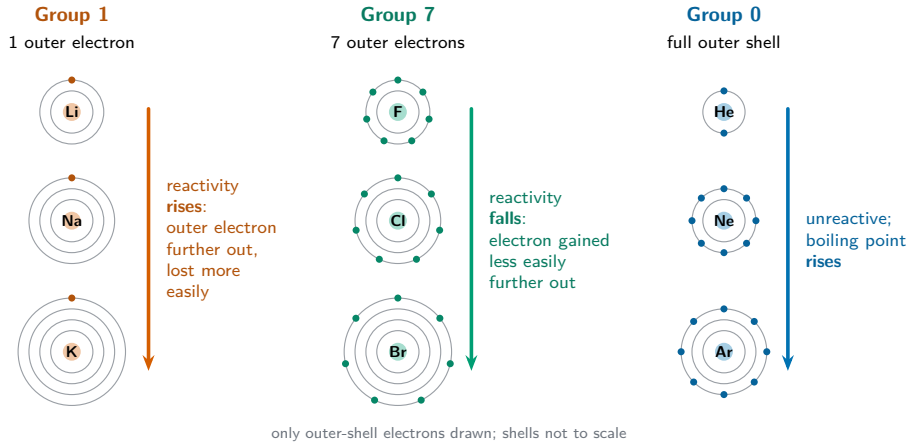
Potassium reacts violently with water, giving off hydrogen that burns with a lilac flame – reactivity increases down Group I

Group 7. Each **halogen** 卤素 has seven outer electrons and exists as molecules of two atoms: Cl_2, Br_2, I_2. At room temperature chlorine is a pale green gas, bromine a red-brown liquid and iodine a grey-black solid. Down the group, melting and boiling points increase but reactivity decreases: the outer shell is further from the nucleus, so it attracts an extra electron less strongly. Halogens form ionic compounds with metals (sodium chloride, NaCl) and covalent compounds with non-metals (hydrogen chloride, HCl).

Displacement. In a **displacement** 置换 reaction, a more reactive halogen takes the place of a less reactive halogen in a solution of its salt.

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Worked example 9. When chlorine water is added to potassium bromide solution, the solution turns orange as bromine forms: $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$. If bromine water is added to potassium chloride, nothing happens, because bromine is less reactive than chlorine. (HT) The ionic equation is $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$: chlorine gains electrons and bromide ions lose them.



Group 0, 1 and 7 trend summary: reactivity, outer electrons and boiling points.

Transition metals (chemistry only)

The **transition metals** 过渡金属 are in the central block of the periodic table; examples are chromium, manganese, iron, cobalt, nickel and copper. Compared with the Group 1 metals, they have higher melting points and densities, are stronger and harder, and are much less reactive: iron and copper do not react with cold water, while sodium reacts violently.

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Transition metals form coloured compounds: each of these solutions contains a different transition-metal ion

Transition metals have three special properties:

- they can form ions with different charges, such as Fe^{2+} and Fe^{3+};
- their compounds are often coloured: copper(II) sulfate solution is blue, iron(II) compounds are often pale green and iron(III) compounds orange-brown;
- they and their compounds are often useful as a **catalyst** 催化剂: iron in the Haber process for making ammonia, and nickel for turning vegetable oils into margarine.

Worked example 10. The sulfate ion is SO_4^{2-}. With Fe^{2+}, one of each balances the charge: $FeSO_4$, iron(II) sulfate. With Fe^{3+}, you need two iron ions (+6) and three sulfate ions (−6): $Fe_2(SO_4)_3$, iron(III) sulfate. They are different compounds of the same three elements.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.1.
- D. Mendeleev, "Die periodische Gesetzmässigkeit der chemischen Elemente", *Annalen der Chemie und Pharmacie*, supplement 8 (1871), with the predictions for eka-aluminium.
- P.-É. Lecoq de Boisbaudran, reports on gallium in *Comptes rendus* (1875–1876).
- H. Geiger and E. Marsden, "On a diffuse reflection of the α-particles", *Proceedings of the Royal Society A* (1909), and E. Rutherford, *Philosophical Magazine* (May 1911).
- J. Chadwick, "Possible existence of a neutron", *Nature* (27 February 1932).

Common errors

- Changing a formula to balance an equation; change only the numbers in front.
- Saying the atomic number is the number of protons plus neutrons; that is the mass number.
- Saying an ion forms when an atom gains or loses protons; only electrons change.
- Saying isotopes are different elements; they have the same number of protons.
- Taking a simple average of isotope masses instead of a weighted mean.
- Saying the alpha particles that bounced back show the atom is solid; most went straight through.
- Saying Group 1 reactivity increases because the outer electron is "more attracted"; it is less attracted, so it is lost more easily.

Key terms

- **Element**: a substance made of only one kind of atom, all with the same number of protons.
- **Compound**: two or more elements chemically combined in fixed proportions.
- **Mixture**: two or more substances that are not chemically combined.
- **Atomic number**: the number of protons in the nucleus of an atom.
- **Mass number**: the total number of protons and neutrons in an atom.
- **Isotope**: an atom of an element with a different number of neutrons.
- **Relative atomic mass**: the weighted mean mass of the atoms of an element.
- **Group**: a column of the periodic table; elements in it have the same number of outer electrons.

Check yourself

1. Is air an element, a compound or a mixture? Explain.
2. Which method would you use to separate the dyes in a black felt-tip pen?
3. What did the alpha particles that bounced back tell Rutherford?
4. How many protons, neutrons and electrons are in a potassium-39 atom (atomic number 19)?
5. Boron is 20% boron-10 and 80% boron-11. Calculate its relative atomic mass.
6. Write the electron arrangement of aluminium (atomic number 13), and give its group and period.
7. Explain why potassium reacts more vigorously with water than lithium does.
8. Will iodine displace chlorine from sodium chloride solution? Explain.

Answers: 1. A mixture: it contains nitrogen, oxygen and other gases that are not chemically joined. 2. Chromatography. 3. The atom has a tiny, dense, positive centre, the nucleus, that contains most of the mass. 4. 19 protons, $39 - 19 = 20$ neutrons and 19 electrons. 5. $(10 \times 20 + 11 \times 80)/100 = (200 + 880)/100 = 10.8$. 6. 2,8,3; Group 3, Period 3. 7. Potassium's outer electron is further from the nucleus and more shielded, so it is lost more easily. 8. No: iodine is less reactive than chlorine.

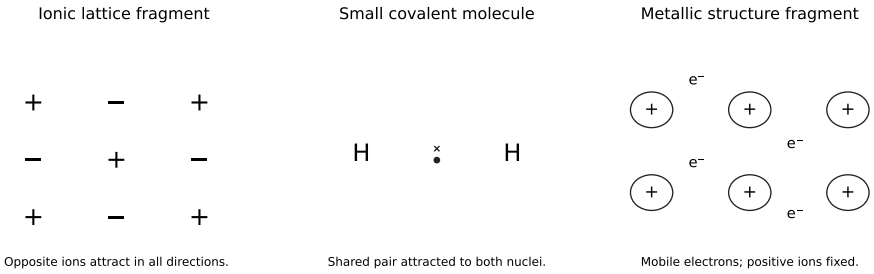
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2 · Bonding, structure and the properties of matter

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Bonding, structure and the properties of matter

Why it matters. One Friday evening in 2004, Andre Geim and Konstantin Novoselov at the University of Manchester pressed sticky tape onto a block of graphite, the grey material in a pencil, and peeled it off. They folded the tape and peeled it again, and again, until some flakes on the tape were only one atom thick. They had made **graphene** 石墨烯, a sheet of carbon atoms that is stronger than steel, conducts electricity better than copper and is almost transparent. In 2010 they won the Nobel Prize in Physics. Graphite, diamond and graphene are all pure carbon. They behave so differently because their atoms are joined in different ways. This topic explains how the bonds between particles decide the properties of every material.



Ionic lattice, covalent shared pair and metallic mobile charge carriers.

In this unit you will:

- describe ionic, covalent and metallic bonding in terms of electrons;
- draw dot-and-cross diagrams for ions and small molecules, and say what the diagrams leave out;
- explain melting points and electrical conductivity from structure and bonding;
- explain the properties of diamond, graphite, graphene and fullerenes;
- (chemistry only) calculate surface-area-to-volume ratios and weigh the uses and risks of nanoparticles.

AQA GCSE Chemistry (8462) topic 4.2, also in Combined Science: Trilogy (8464) except section 4.2.4 on nanoparticles. It is tested on Paper 1, and its ideas return in almost every later topic.

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Three kinds of strong bond

All chemical bonds come from the same force: the **electrostatic force** 静电力 of attraction between opposite charges. What changes from one kind of bond to another is where the electrons go.

Bond	Between	What the electrons do
ionic	a metal and a non-metal	the metal transfers electrons to the non-metal
covalent	non-metal atoms	atoms share pairs of electrons
metallic	atoms in a metal	outer electrons leave their atoms and move through the whole structure

Ionic bonding. In an **ionic bond** 离子键, a metal atom loses its outer electrons and becomes a positive ion; a non-metal atom gains them and becomes a negative ion. The opposite ions then attract each other. Metals in Groups 1 and 2 and non-metals in Groups 6 and 7 form ions with the electronic structure of a noble gas, so you can read the charge from the group number: Group 1 gives +1, Group 2 gives +2, Group 6 gives −2 and Group 7 gives −1.

Covalent bonding. In a **covalent bond** 共价键, two non-metal atoms share a pair of electrons. Learn the dot-and-cross diagrams for H_2, Cl_2, O_2, N_2, HCl, H_2O, NH_3 and CH_4.

Metallic bonding. In a **metallic bond** 金属键, the metal atoms lose their outer electrons, which become **delocalised electrons** 离域电子: free to move through the whole metal. A regular structure of positive ions is held together by its attraction to this "sea" of electrons.

■ Drawing electrons: two worked examples

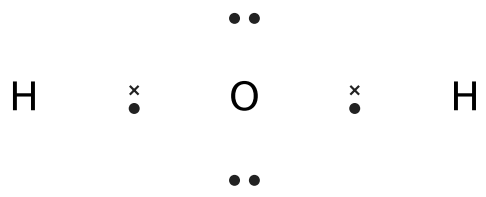
Add up the outer electrons of every atom. Use shared pairs so that each hydrogen has two electrons around it and each other atom has eight. Any electrons left over form **lone pairs** 孤电子对, pairs that belong to one atom only. Dots and crosses show where each electron came from; the electrons themselves are identical. A single bond shares two electrons, a double bond four, and a triple bond six.

Molecule	Total outer electrons	Shared pairs	Lone pairs
H_2	2	1	0
Cl_2	14	1	3 on each Cl
O_2	12	2	2 on each O
N_2	10	3	1 on each N
HCl	8	1	3 on Cl
H_2O	8	2	2 on O
NH_3	8	3	1 on N

Molecule	Total outer electrons	Shared pairs	Lone pairs
CH_4	8	4	0

Worked example 1: water. Oxygen has 6 outer electrons and each hydrogen has 1, so the total is 8. Two O–H bonds use 4 electrons. The other 4 form two lone pairs on the oxygen. Check: each hydrogen has 2 electrons around it, and oxygen has 8.

water: dots from O; crosses from H



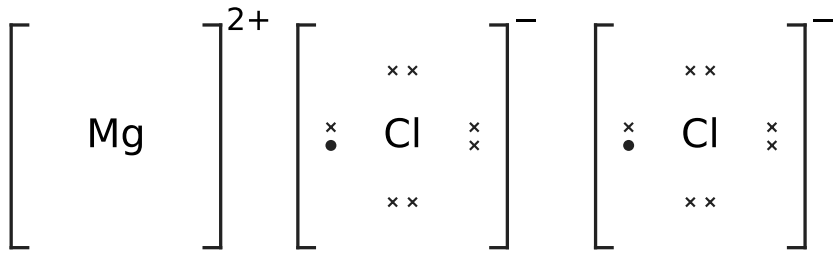
Flat electron model: positions do not establish the 3D shape.

Complete outer-electron diagram: water

Worked example 2: magnesium chloride. Magnesium is 2,8,2. It loses two electrons and becomes Mg^{2+}, which is 2,8. Each chlorine atom is 2,8,7. Each gains one electron and becomes Cl^-, which is 2,8,8. One Mg^{2+} balances the charge of two Cl^-, so the formula is $MgCl_2$. Draw square brackets around each ion with its charge outside.

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$MgCl_2$: one Mg^{2+} and two Cl^- ions



Each chloride: seven crosses + one magnesium dot.
Mg^{2+} has configuration 2,8; its old third shell is empty.

Complete outer-electron diagram: magnesium-chloride

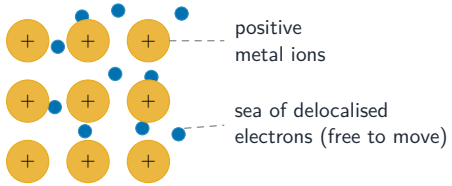
Magnesium chloride is not made of $MgCl_2$ molecules. It is an **ionic lattice** 离子晶格, a giant structure in which every ion is surrounded by ions of the opposite charge. If a drawing of a lattice shows 6 magnesium ions and 12 chloride ions, the ratio 6 : 12 simplifies to 1 : 2, which gives the formula $MgCl_2$.

A **line representation** 线式 draws each shared pair as one line and leaves out the lone pairs: water is H–O–H, oxygen is O=O and nitrogen is N≡N. A polymer is drawn as one repeating unit in brackets, with bonds that cross the brackets to show that the chain continues.

What the models leave out. Every model simplifies. A dot-and-cross diagram does not show the shape of the molecule. A ball-and-stick model shows gaps between atoms that really touch, and its sticks are not real. A flat drawing cannot fully show a 3D structure, and a drawing of a giant structure shows only a small piece of it.

How structure decides properties

States of matter. A substance melts and freezes at its melting point, and boils and condenses at its boiling point. In the particle model, particles are small solid spheres. To melt or boil a substance you must overcome the forces between its particles, so the stronger those forces, the higher the melting and boiling points. (Higher tier) The particle model has limits: it shows no forces, and real particles are not solid spheres. State symbols are (s) solid, (l) liquid, (g) gas and (aq) dissolved in water.



Metallic bonding: positive ions in a sea of delocalised electrons

The table brings the structures together. Read each row from left to right as a chain of cause and effect.

Structure	Forces to overcome on melting	Melting point	Conducts electricity?
ionic lattice (NaCl)	strong attraction between ions, in all directions	high	only when molten or dissolved, when the ions can move
small molecules (H_2O, CO_2)	only the weak **intermolecular force** 分子间作用力 between molecules; the covalent bonds inside do not break	low	no: the molecules have no overall charge
polymers (poly(ethene))	intermolecular forces between very long chains	solid at room temperature	no
giant covalent (diamond, silica)	many strong covalent bonds	very high	no, except graphite and graphene
metals (copper)	strong metallic bonding	mostly high	yes: delocalised electrons carry the charge, and heat too

Worked example: why does salt conduct only when it melts? Electric current needs a **charge carrier** 载流粒子, a charged particle that can move. Solid sodium chloride contains charged ions, but they are held in fixed positions, so it does not conduct. When it melts or dissolves, the ions are free to move and carry the current. Notice that the ions were there all the time: what changes is whether they can move.

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Worked example: water and silica. Both contain covalent bonds, but water boils at 100 °C and silica (SiO_2, sand) melts at about 1,700 °C. When water boils, the molecules separate but every O–H bond stays whole: only weak intermolecular forces are overcome. Silica is a giant covalent structure, so melting it means breaking a huge number of strong covalent bonds.

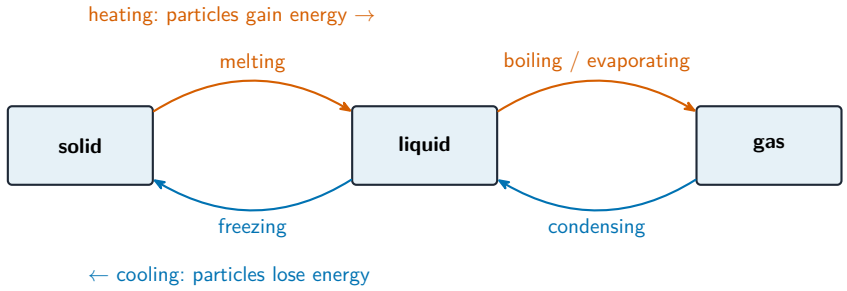
Bigger molecules have stronger intermolecular forces, so their melting and boiling points are higher. Polymer chains are very long, so the forces between them are strong enough to make most polymers solid at room temperature.

Alloys. In a pure metal, the layers of identical atoms can slide over each other, so the metal can be bent and shaped. An **alloy** 合金 is a mixture of a metal with other elements. The atoms of different sizes distort the layers and make sliding harder, so alloys are harder than pure metals. That is why gold for jewellery is mixed with copper or silver.

Predicting a state

A substance is solid below its melting point, liquid between its melting and boiling points, and gas above its boiling point.

Worked example. Substance X melts at −15 °C and boils at 65 °C. At −30 °C it is solid; at 20 °C it is liquid; at 90 °C it is a gas.



Heating: solid → liquid → gas; cooling reverses each change

In a solid the particles vibrate about fixed positions. When it melts, they gain enough energy to move past each other but stay close together. When it boils, they move far apart. The particles themselves do not get bigger. A single atom has no melting point, hardness or conductivity: these are properties of a large number of particles and the way they are arranged.

Carbon: one element, four structures

Carbon atoms can form four covalent bonds, and the way they use them gives very different materials.

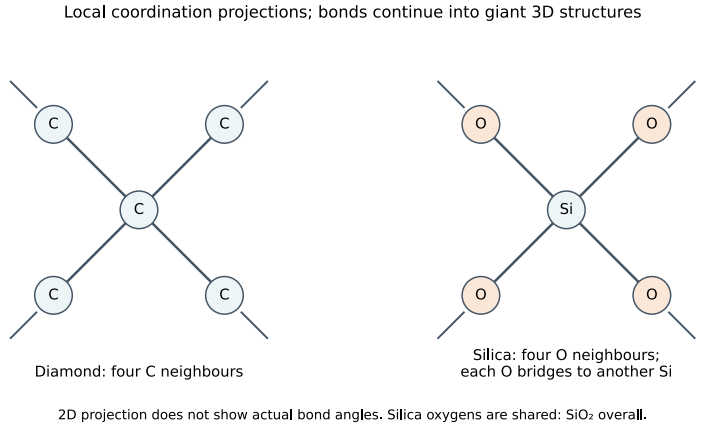
- **Diamond** 金刚石 is a **giant covalent structure** 巨型共价结构 in which every carbon atom is bonded to four others. It is very hard, has a very high melting point and does not conduct electricity, because all its outer electrons are in bonds.
- **Graphite** 石墨 is made of layers of hexagonal rings. Each carbon is bonded to three others, so one electron per atom is delocalised. There are no covalent bonds between the layers, only weak forces, so the layers slide: graphite is soft and slippery and is used in pencils and as a lubricant. The delocalised electrons let it conduct electricity.
- Graphene is a single layer of graphite. It is very strong and conducts electricity, so it is useful in electronics and composite materials.
- A **fullerene** 富勒烯 is a hollow molecule of carbon atoms in rings of five, six or seven atoms. The first discovered, buckminsterfullerene (C_{60}), was made in 1985 and is shaped like a football. A **carbon nanotube** 碳纳米管 is a cylinder-shaped fullerene that is very long compared with its width, which makes it strong and useful in electronics and for strengthening materials.



Natural graphite: those weak forces between the layers are why the specimen flakes, and why it marks paper



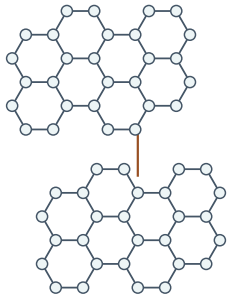
A real diamond in the rock it grew in. That whole crystal is one giant molecule – a single unbroken network of carbon atoms, each bonded to four others. Breaking it means breaking countless strong covalent bonds, which is why diamond is the hardest natural material



Diamond: each carbon bonds to four others; silica: each silicon to four oxygens, each oxygen shared

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Graphite: hexagonal covalent layers, with weaker attraction between layers



Strong covalent bonds within layers

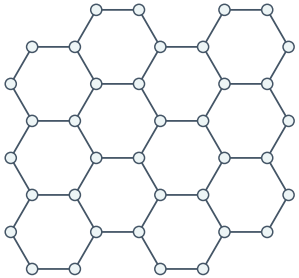
One delocalised electron per carbon

No covalent bonds between layers

Each interior carbon has three neighbours. Cut edges omit further bonds.

Graphite: layers of hexagons, three bonds per carbon, one delocalised electron each, weak forces between layers

Graphene: one hexagonal carbon layer



Strong covalent connections

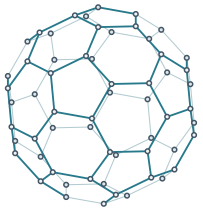
Delocalised electrons

Carbon atoms shown at vertices; cut edges continue. Single-layer schematic.

Graphene: a single layer of graphite, one atom thick

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C_{60} carbon-bond projection and schematic nanotube shape



C_{60}: spherical cage
60 carbon atoms



Nanotube: long hollow cylinder
large length/diameter ratio

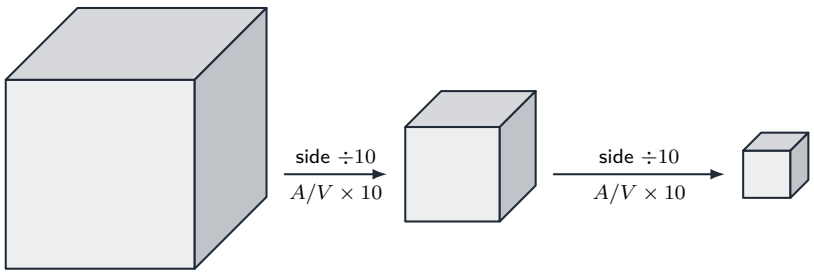
C_{60}: 60 vertices, 90 bonds; rear edges pale. Nanotube outline omits its carbon-ring bonds.

Fullerenes: C_{60} is a hollow cage of 60 carbons; a nanotube is a long hollow cylinder

> **Worked example: graphite and metals.** An exam question may ask why graphite conducts electricity "like a metal". Both have delocalised electrons that can move through the structure and carry charge. In graphite they move along the layers; in a metal they move in every direction.

Nanoparticles (chemistry only)

Nanoscience studies structures from 1 to 100 nanometres (nm) across; 1 nm is 10^{-9} m. A **nanoparticle** 纳米颗粒 contains only a few hundred atoms. A typical atom is about 0.2 nm across, so a 20 nm particle is about 100 atoms wide. Nanoparticles are smaller than fine particles (PM2.5, 100 to 2,500 nm), which are smaller than coarse particles, or dust (PM10, 2,500 to 10,000 nm, that is 2.5×10^{-6} to 1×10^{-5} m).



side 100 nm
$A/V = 6/100 = 0.06$ nm^{-1}

side 10 nm
$A/V = 6/10 = 0.6$ nm^{-1}

side 1 nm
$A/V = 6/1 = 6$ nm^{-1}

Drawings shrink by half each step for visibility; the real cubes shrink tenfold.

Model cubes within the nanoparticle range, with correctly unit-labelled area/volume quotients.

Surface area and volume

For a cube with side l, the **surface area** 表面积 is $6l^2$ and the **volume** 体积 is l^3, so the ratio of surface area to volume is $6/l$. When the side gets 10 times smaller, the ratio gets 10 times bigger.

> **Worked example.** A cube of side 100 nm has surface area $6 \times 100^2 = 60\,000$ nm^2 and volume $100^3 = 1\,000\,000$ nm^3, so the ratio is 0.06 nm^{-1}. A cube of side 10 nm has surface area 600 nm^2 and volume 1000 nm^3, so the ratio is 0.6 nm^{-1}: ten times bigger. Each small cube has less surface than the big one, but if you cut the same amount of material into small cubes, much more of it is on the surface.

Because so much of a nanoparticle is surface, it can behave differently from a lump of the same material. Silver is unreactive in a ring, but silver nanoparticles kill bacteria, so they are used in wound dressings and in some socks. Titanium dioxide nanoparticles in sun cream block ultraviolet light without leaving a white layer on the skin. Catalysts made of nanoparticles need less material for the same effect. Nanoparticles are also used in electronics and for delivering medicines.

Weighing the risks. Scientists do not yet fully know what nanoparticles do inside the body or in the environment. They are small enough to pass into cells, and silver nanoparticles washed out of socks may harm useful bacteria in water treatment. When a question gives you data, separate what the data show (for example, a nano catalyst works three times faster) from what they do not show (that it is safe), and say what further tests would help.

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Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.2.
- K. S. Novoselov, A. K. Geim and others, "Electric field effect in atomically thin carbon films", *Science* (22 October 2004).
- The Nobel Prize in Physics 2010, press release (5 October 2010).
- H. W. Kroto and others, "C_{60}: Buckminsterfullerene", *Nature* (14 November 1985).

Common errors

- Saying that covalent bonds break when a simple molecular substance melts or boils; only the weak intermolecular forces are overcome.
- Saying that solid ionic compounds do not conduct because they have no ions; they have ions, but the ions cannot move.
- Calling an ionic compound a molecule, or reading a lattice drawing such as Mg_6Cl_{12} as a formula.
- Saying graphite conducts because of its weak forces between layers; it conducts because of its delocalised electrons.
- Forgetting that diamond does not conduct: it has no delocalised electrons.
- Saying a nanoparticle is a "large atom"; it contains hundreds of atoms.

Key terms

- **Ionic bond**: the electrostatic attraction between oppositely charged ions.
- **Covalent bond**: a shared pair of electrons between two atoms.
- **Metallic bond**: the attraction between positive metal ions and delocalised electrons.
- **Delocalised electron**: an electron that is free to move through a structure.
- **Intermolecular force**: a weak force of attraction between molecules.
- **Giant covalent structure**: a large network of atoms all joined by covalent bonds.
- **Alloy**: a mixture of a metal with one or more other elements.
- **Nanoparticle**: a particle between 1 and 100 nm across.

Check yourself

1. What charge does an ion of a Group 2 metal have, and why?
2. How many lone pairs does the nitrogen atom in NH_3 have?
3. Explain why methane (CH_4) has a low boiling point.
4. Explain why molten sodium chloride conducts electricity but solid sodium chloride does not.
5. Explain why graphite is soft but diamond is very hard.

6. Why are alloys harder than pure metals?
7. (Chemistry only) A cube of side 20 nm is cut into cubes of side 2 nm. By what factor does the surface-area-to-volume ratio change?

Answers: 1. +2: the atom loses its two outer electrons to reach the electronic structure of a noble gas. 2. One (8 outer electrons: 3 shared pairs use 6, leaving one pair). 3. It is made of small molecules; boiling overcomes only the weak intermolecular forces, which needs little energy. 4. In both there are ions, but only in the molten salt can they move and carry the charge. 5. Graphite's layers are held together only by weak forces, so they slide; in diamond every carbon is joined to four others by strong covalent bonds. 6. Atoms of different sizes distort the layers, so the layers cannot slide over each other easily. 7. It increases 10 times (from $6/20 = 0.3$ nm^{-1} to $6/2 = 3$ nm^{-1}).

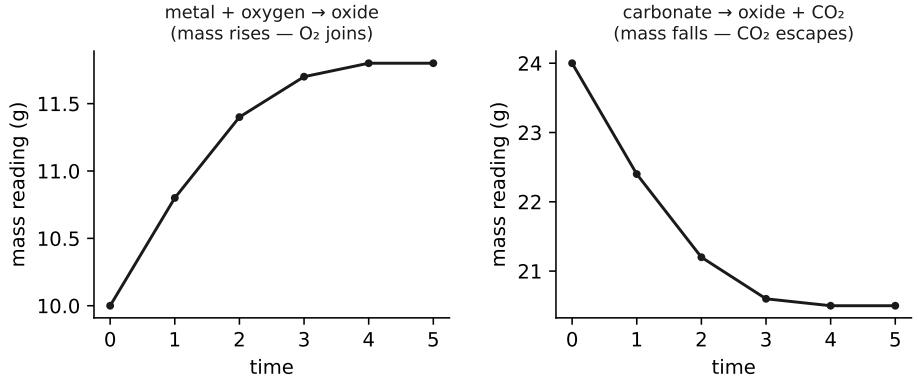
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3 · Quantitative chemistry

AQA · GCSE · Chemistry

Quantitative chemistry: counting atoms by mass

Why it matters. On 23 July 1983, Air Canada Flight 143 ran out of fuel at 12,500 metres over Manitoba. The new Boeing 767 measured its fuel in kilograms, but the ground crew used the old conversion factor for pounds: 1.77 instead of 0.80 per litre. The plane had less than half the fuel the pilots thought. Captain Robert Pearson glided it to a landing on a disused runway at Gimli, and everyone survived. The mistake was not in the arithmetic; it was in the units. Chemists face the same danger every day. This topic shows how to count atoms by weighing them, and how to keep track of every gram and every unit.



Mass-time graphs for oxidation (mass rises) and carbonate decomposition (mass falls).

In this unit you will:

- use conservation of mass to explain why some reactions seem to gain or lose mass;
- calculate relative formula masses and percentages by mass;
- (higher tier) convert between mass, moles and numbers of particles, and use moles to find reacting masses and limiting reactants;
- calculate concentrations in g/dm^3 and (higher tier) mol/dm^3;
- (chemistry only) calculate percentage yield, atom economy and gas volumes.

AQA GCSE Chemistry (8462) topic 4.3, tested on Paper 1. Combined Science: Trilogy (8464) includes sections 4.3.1 and 4.3.2 only; sections 4.3.3 to 4.3.5 are chemistry only, and much of the mole work is higher tier (HT).

Conservation of mass and equations

The law of **conservation of mass** 质量守恒 says that atoms are not made or destroyed in a chemical reaction. They are only rearranged, so the total mass of the products equals the total mass of the reactants. A symbol equation must therefore be **balanced** 配平: the same number of each kind of atom on both sides.

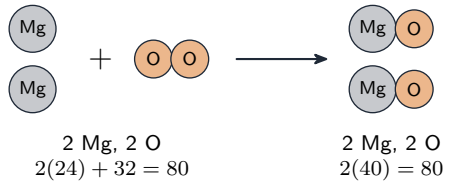
There are two kinds of number in an equation. A big number in front of a formula tells you how many units react. A **subscript** 下标, the small number inside a formula, tells you how many atoms are in one unit. To balance an equation you change the big numbers only: changing a subscript changes the substance.

Relative formula mass. The **relative formula mass** 相对式质量 (Mr) of a compound is the sum of the **relative atomic mass** 相对原子质量 (Ar) of every atom in its formula. The **percentage by mass** 元素的质量分数 of an element in a compound is:

$$\text{percentage by mass} = \frac{A_r \times \text{number of atoms of the element}}{M_r} \times 100$$

> **Worked example 1: ammonium sulfate.** $(NH_4)_2SO_4$ contains 2 N, 8 H, 1 S and 4 O. With N = 14, H = 1, S = 32 and O = 16, $M_r = 2(14)+8(1)+32+4(16) = 132$. The percentage of sulfur is $32/132 \times 100 = 24.2\%$. Remember that the 2 outside the bracket multiplies everything inside it.

> **Worked example 2: checking an equation.** In $2Mg + O_2 \rightarrow 2MgO$ there are 2 Mg and 2 O atoms on each side. Using Mg = 24 and O = 16, the reactants total $2(24)+32 = 80$ and the products total $2(40) = 80$. Multiply each Mr by the number in front of the formula before you add.



not $24 + 32 = 40$: count every atom

Count every atom: both sides total 80

Reactions with gases. Some reactions seem to break the law. When magnesium burns in air, the solid gets heavier: during **oxidation** 氧化, oxygen atoms from the air join the metal. When a **carbonate** 碳酸盐 is heated, the solid gets lighter: **thermal decomposition** 热分解 releases carbon dioxide, which escapes. For example, heating 10.0 g of calcium carbonate leaves 5.6 g of calcium oxide, and 4.4 g of carbon dioxide escapes: $5.6 + 4.4 = 10.0$ g. No mass is lost once you count the gas.

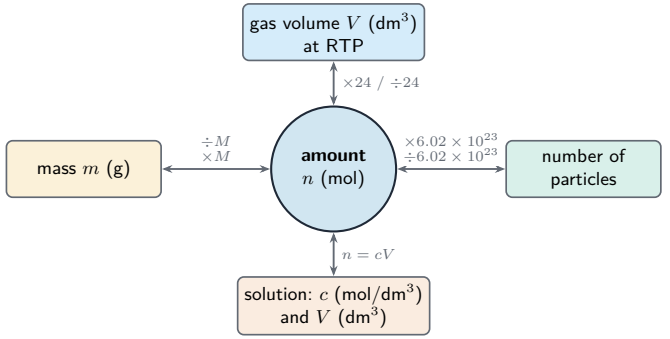
Uncertainty. Every measurement has some **uncertainty** 不确定度. Repeat a measurement, calculate the **mean** 平均值, and use the **range** 极差 of the results to show the spread.

> **Worked example 3.** Three masses are 2.40 g, 2.42 g and 2.44 g. The mean is $(2.40 + 2.42 + 2.44)/3 = 2.42$ g. The range is $2.44 - 2.40 = 0.04$ g, so the result can be written as 2.42 ± 0.02 g, using half the range. Readings that agree closely are precise, but they could all be wrong if the balance is not zeroed.

Moles and reacting masses (higher tier)

Atoms are far too small to count one by one, so chemists count them in moles. One **mole** 摩尔 of a substance contains 6.02×10^{23} particles: this number is the **Avogadro constant** 阿伏伽德罗常数. The mass of one mole in grams is equal to the Mr, so one mole of carbon atoms (Ar 12) weighs 12 g and one mole of water (Mr 18) weighs 18 g. This mass per mole is the **molar mass** 摩尔质量, in g/mol.

$$\text{moles} = \frac{\text{mass (g)}}{M_r}$$



worked chain: $CaCO_3 \rightarrow CaO + CO_2$



the equation ratio links amounts in moles, never masses

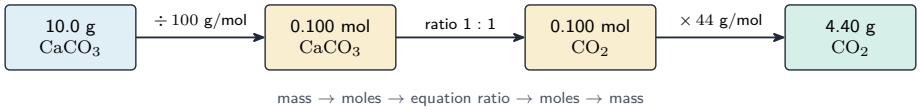
Always say what you are counting. One mole of oxygen molecules (O_2) contains two moles of oxygen atoms.

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Worked example 4: mass to particles. How many formula units are in 8.0 g of sodium hydroxide, NaOH? $M_r = 23 + 16 + 1 = 40$, so moles $= 8.0/40 = 0.20$ mol. Number of formula units $= 0.20 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$.

Equations as mole ratios. The numbers in a balanced equation give **mole ratios** 摩尔比. $Mg + 2HCl \rightarrow MgCl_2 + H_2$ means that 1 mol of magnesium reacts with 2 mol of hydrochloric acid to make 1 mol of magnesium chloride and 1 mol of hydrogen. To find a reacting mass, follow three steps: mass to moles, use the ratio, moles to mass.

Worked example 5: reacting mass. What mass of carbon dioxide is made when 10.0 g of calcium carbonate decomposes? $CaCO_3 \rightarrow CaO + CO_2$. Moles of $CaCO_3$ $= 10.0/100 = 0.100$ mol. The ratio is 1 : 1, so 0.100 mol of CO_2 forms. Mass of $CO_2 = 0.100 \times 44 = 4.40$ g.



Reacting mass: mass to moles, the equation ratio, then moles to mass

Balancing from masses. Convert each mass to moles, divide by the smallest number of moles, and use the whole-number ratio.

Worked example 6. 4.8 g of Mg reacts with 3.2 g of O_2 to make 8.0 g of MgO. Moles: Mg $4.8/24 = 0.20$; O_2 $3.2/32 = 0.10$; MgO $8.0/40 = 0.20$. Dividing by 0.10 gives 2 : 1 : 2, so the equation is $2Mg + O_2 \rightarrow 2MgO$. Use 32 for O_2, because oxygen gas is made of molecules.

Limiting reactants. A reaction stops when one reactant is used up. That reactant is the **limiting reactant** 限量反应物; the other is in **excess** 过量. Calculate the amount of product from the limiting reactant only.

Worked example 7. 4.8 g of Mg (0.20 mol) is mixed with 6.4 g of O_2 (0.20 mol). The equation needs 2 mol of Mg for every 1 mol of O_2, so 0.20 mol of Mg uses only 0.10 mol of O_2. Magnesium is limiting. It makes 0.20 mol of MgO, which is $0.20 \times 40 = 8.0$ g. The 0.10 mol of oxygen left over has a mass of 3.2 g. Check: $4.8 + 6.4 = 8.0 + 3.2 = 11.2$ g.

Concentration. The **concentration** 浓度 of a solution is the amount of **solute** 溶质, the dissolved substance, in each cubic decimetre (dm^3) of solution. 1 dm^3 = 1000 cm^3.

$$\text{concentration (g/dm}^3\text{)} = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3\text{)}}$$

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Worked example 8 (both tiers). 3.5 g of solute is dissolved to make 250 cm^3 of solution. Volume $= 250/1000 = 0.250$ dm^3, so concentration $= 3.5/0.250 = 14$ g/dm^3. A 100 cm^3 portion of it contains $14 \times 0.100 = 1.4$ g. If you double the volume with the same mass of solute, the concentration halves.

Yield and atom economy (chemistry only)

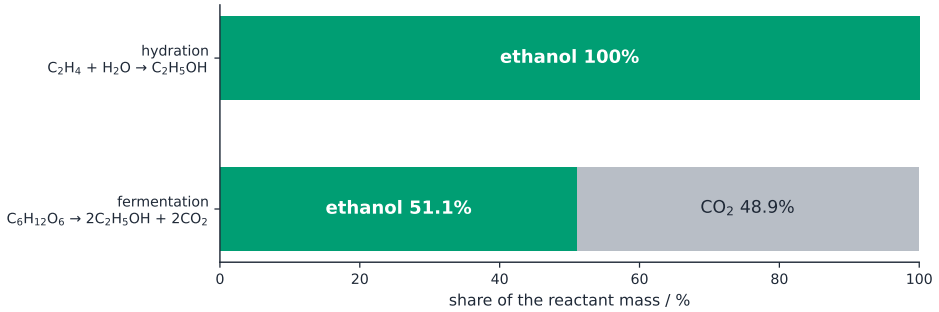
Percentage yield. You almost never get the full amount of product that the equation predicts, the **theoretical mass** 理论质量. The reaction may be **reversible** 可逆, so it does not go to completion; some product is lost when it is separated; and some reactants may react in other ways.

$$\text{percentage yield} = \frac{\text{mass of product actually made}}{\text{maximum theoretical mass}} \times 100$$

Worked example 9. The theoretical mass is 8.0 g and you collect 6.0 g of dry product. The **percentage yield** 产率 is $6.0/8.0 \times 100 = 75\%$. A yield above 100% means the product is still wet or impure, not that atoms were made.

Atom economy. **Atom economy** 原子经济 measures how much of the mass of the reactants ends up in the useful product. A high atom economy means less waste, which is better for the environment and usually cheaper.

$$\text{atom economy} = \frac{\text{total } M_r \text{ of the desired product}}{\text{total } M_r \text{ of all reactants}} \times 100$$



The displayed hydration equation has ethanol as its only product, so its theoretical atom economy is 100%

Worked example 10. In $CaCO_3 \rightarrow CaO + CO_2$, if calcium oxide (quicklime) is the useful product, the atom economy is $56/100 \times 100 = 56\%$. The rest leaves as carbon dioxide. For $2H_2 + O_2 \rightarrow 2H_2O$, the product is $2 \times 18 = 36$ and the reactants are $2 \times 2 + 32 = 36$, so the atom economy is 100%. Include the numbers in front of the formulae.

Atom economy comes from the equation; yield comes from the experiment. A reaction with 100% atom economy can still give a low yield.

Choosing a route (higher tier). Industry compares atom economy, yield, the rate of the reaction, the energy needed, the position of equilibrium, and whether a **by-product** 副产物 can be sold. Ethanol can be made by hydrating ethene (100% atom economy, but from crude oil, at high temperature and pressure) or by fermenting sugar (51% atom economy, but from a renewable crop, at low temperature). Neither number alone decides which is better.

Concentrations in mol/dm³ (chemistry only, higher tier)

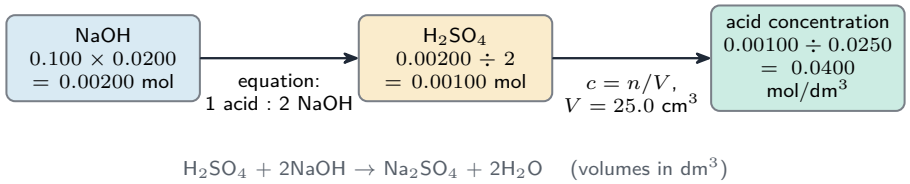
Molar concentration 物质的量浓度 is the number of moles of solute in each dm^3 of solution:

$$\text{moles} = \text{concentration (mol/dm}^3\text{)} \times \text{volume (dm}^3\text{)}$$

To change it into **mass concentration** 质量浓度, multiply by the Mr: 0.20 mol/dm^3 NaOH is $0.20 \times 40 = 8.0$ g/dm^3.

In a **titration** 滴定, you measure the exact volumes of an acid and an alkali that react. If you know one concentration, you can calculate the other, using the mole ratio in the equation.

Worked example 11. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$. 25.0 cm^3 of sulfuric acid reacts with 20.0 cm^3 of 0.100 mol/dm^3 sodium hydroxide. Moles of NaOH $= 0.100 \times 0.0200 = 0.00200$ mol. The ratio is 1 : 2, so moles of acid $= 0.00200/2 = 0.00100$ mol. Concentration of acid $= 0.00100/0.0250 = 0.0400$ mol/dm^3.



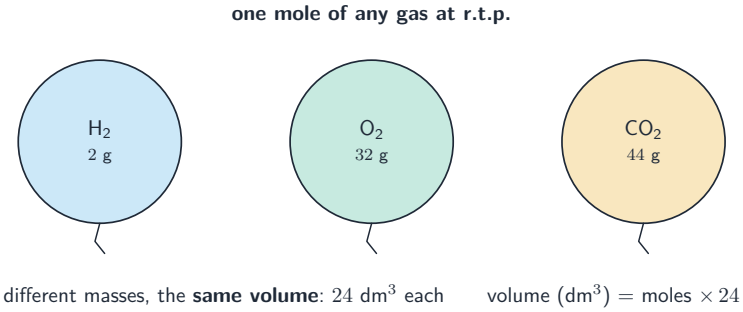
The equation, not the volumes, sets the mole ratio

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Gas volumes (chemistry only, higher tier)

Equal volumes of different gases at the same temperature and pressure contain the same number of molecules. At **room temperature and pressure** 室温室压 (20 °C and 1 atmosphere), one mole of any gas takes up 24 dm^3.

$$\text{volume of gas (dm}^3\text{)} = \text{moles} \times 24$$



Different masses, the same volume: one mole of any gas fills 24 dm^3 at room conditions

Worked example 12. What volume does 4.4 g of carbon dioxide take up at room temperature and pressure? Moles $= 4.4/44 = 0.100$ mol, so volume $= 0.100 \times 24 = 2.40$ dm^3, or 2400 cm^3.

Because volume is proportional to moles, the volumes of reacting gases are in the same ratio as the numbers in the equation.

Worked example 13. $N_2 + 3H_2 \rightarrow 2NH_3$. 60 cm^3 of nitrogen needs $3 \times 60 = 180$ cm^3 of hydrogen and makes $2 \times 60 = 120$ cm^3 of ammonia. If only 150 cm^3 of hydrogen is available, hydrogen is limiting: it uses $150/3 = 50$ cm^3 of nitrogen, makes $150 \times 2/3 = 100$ cm^3 of ammonia, and 10 cm^3 of nitrogen is left. The ratio rule applies only to gases, measured at the same temperature and pressure.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.3.
- G. H. Lockwood, *Final Report of the Board of Inquiry Investigating the Circumstances of an Accident Involving the Air Canada Boeing 767 Aircraft C-GAUN that Effected an Emergency Landing at Gimli, Manitoba* (Government of Canada, April 1985).
- A. Lavoisier, *Traité élémentaire de chimie* (1789), where the law of conservation of mass was first stated clearly.

- International Bureau of Weights and Measures, *The International System of Units* (9th edition, 2019), on the mole and the Avogadro constant.

Common errors

- Changing a subscript to balance an equation; change only the numbers in front of formulae.
- Forgetting to multiply by the number in front of a formula when adding up masses.
- Saying mass is "lost" when a carbonate is heated; the carbon dioxide escapes into the air.
- Using 16 instead of 32 for oxygen gas, which is made of O_2 molecules.
- Using cm^3 in a formula that needs dm^3; divide by 1000 first.
- Calculating the product from the reactant in excess instead of the limiting reactant.
- Assuming the moles of acid equal the moles of alkali in every titration; use the ratio in the equation.

Key terms

- **Conservation of mass**: in a chemical reaction, no atoms are made or destroyed, so total mass does not change.
- **Relative formula mass (Mr)**: the sum of the relative atomic masses of the atoms in a formula.
- **Mole**: the amount of substance that contains 6.02×10^{23} particles.
- **Limiting reactant**: the reactant that is used up first and so limits the amount of product.
- **Concentration**: the amount of solute in each dm^3 of solution.
- **Percentage yield**: the mass of product made as a percentage of the maximum theoretical mass.
- **Atom economy**: the mass of the desired product as a percentage of the total mass of the reactants.

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Check yourself

1. Calculate the Mr of calcium hydroxide, $Ca(OH)_2$ (Ca = 40, O = 16, H = 1).
2. Why does the mass of an open flask fall when marble chips react with acid?
3. (HT) How many moles are in 9.0 g of water?
4. (HT) 2.4 g of magnesium reacts with excess hydrochloric acid: $Mg + 2HCl \rightarrow MgCl_2 + H_2$. What mass of hydrogen is made?
5. 5.0 g of salt is dissolved to make 200 cm^3 of solution. What is the concentration in g/dm^3?
6. (Chemistry only) The theoretical yield is 12.0 g and 9.0 g is made. What is the percentage yield?
7. (Chemistry only, HT) What volume does 0.25 mol of oxygen occupy at room temperature and pressure?

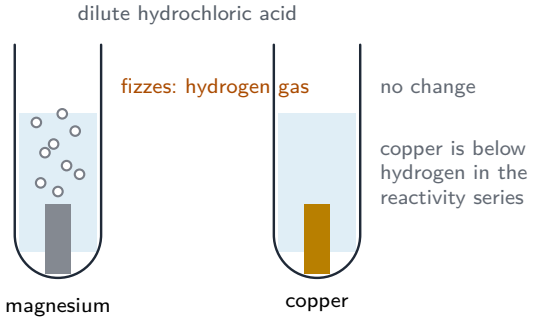
Answers: 1. $40 + 2(16 + 1) = 74$. 2. Carbon dioxide gas is made and escapes from the flask. 3. $9.0/18 = 0.50$ mol. 4. Moles of Mg $= 2.4/24 = 0.10$; ratio 1 : 1, so 0.10 mol of H_2; mass $= 0.10 \times 2 = 0.20$ g. 5. $5.0/0.200 = 25$ g/dm^3. 6. $9.0/12.0 \times 100 = 75\%$. 7. $0.25 \times 24 = 6.0$ dm^3.

4 · Chemical changes

AQA · GCSE · Chemistry

Chemical changes: choose the reaction from the evidence

Why it matters. In 1884 the builders of the Washington Monument put a small pyramid of aluminium on its top. It weighed 2.8 kg and was one of the largest pieces of aluminium in the world: the metal cost more than silver, and the French emperor Napoleon III kept aluminium plates for his most important guests. Aluminium is the most common metal in the Earth's crust, but it holds on to oxygen so strongly that heating its ore with carbon does not release it. Then, in 1886, Charles Martin Hall in the United States and Paul Héroult in France each found the same answer: dissolve the ore in molten cryolite and pass an electric current through it. Within a few years the price had fallen by more than 90%. This topic explains which reactions happen by themselves, and how electricity can force the ones that do not.



Magnesium reacts with dilute acid; copper does not

In this unit you will:

- place metals in the reactivity series from how they react with water, acids and other metal compounds;
- explain oxidation and reduction as gain or loss of oxygen and (higher tier) of electrons;
- make a pure, dry salt and work out its formula;
- explain pH, neutralisation and (higher tier) strong and weak acids, and (chemistry only) carry out and calculate a titration;
- predict the products of electrolysis and explain how aluminium is extracted.

AQA GCSE Chemistry (8462) topic 4.4, tested on Paper 1. It contains two required practicals for all students (making a salt and electrolysis) and one for chemistry only (titration). Higher tier (HT) content is labelled.

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The reactivity of metals

When metals react, their atoms lose electrons and form a **positive ion** 正离子. The more easily a metal does this, the more reactive it is. The **reactivity series** 活动性顺序 puts the metals in order: potassium, sodium, lithium, calcium, magnesium, zinc, iron, copper, from most to least reactive. Carbon and hydrogen are not metals, but they are included as reference points.

more reactive

potassium	cold water: hydroxide + hydrogen
sodium	cold water: hydroxide + hydrogen
lithium	cold water: hydroxide + hydrogen
calcium	cold water: hydroxide + hydrogen
magnesium	cold water: extremely slow
CARBON	reference for oxide reduction
zinc	dilute HCl / H_2SO_4: salt + hydrogen
iron	dilute HCl / H_2SO_4: salt + hydrogen
HYDROGEN	reference for acid / cathode products
copper	no reaction with cold water or dilute HCl

Carbon and hydrogen are non-metals. Compare metals under matching conditions.

Reactivity increases upwards; carbon and hydrogen are reference points.

With water. Potassium, sodium and lithium react with cold water to make hydrogen and a **soluble** 可溶 **hydroxide** 氢氧化物. Lithium floats and fizzes; sodium melts into a ball and moves about; potassium bursts into a lilac flame. Calcium also fizzes and makes calcium hydroxide, which turns the water cloudy. Magnesium reacts only very slowly with cold water, and zinc, iron and copper show no reaction.

With dilute acid. Magnesium, zinc and iron react with dilute hydrochloric acid or sulfuric acid to make a **salt** 盐 and hydrogen. Magnesium fizzes fast, zinc more slowly, and iron very slowly. Copper does not react, because it is below hydrogen. The metals above magnesium react too violently to test safely with acid in school.

Displacement. In a **displacement** 置换 reaction, a more reactive metal takes the place of a less reactive metal in a **compound** 化合物. When zinc is put in blue copper(II) sulfate solution, a brown coating of copper forms and the blue colour fades:

$$\mathrm{Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)}$$

Worked example 1: ranking an unknown metal. Metal X displaces copper from copper sulfate but does not displace zinc from zinc sulfate. So X is less reactive than zinc but more reactive than copper: it could be iron. One test places X between two metals; you need more tests to find its exact position.

Oxidation and reduction. When a metal reacts with oxygen it forms an **oxide** 氧化物: $2Mg + O_2 \rightarrow 2MgO$. Gaining oxygen is **oxidation** 氧化; losing oxygen is **reduction** 还原.

Extracting metals. Only very unreactive metals, such as gold, are found as the metal itself. Most metals are found in the ground as compounds, often oxides, in rocks called ores. A metal less reactive than carbon can be extracted by heating its oxide with carbon, which removes the oxygen. A metal more reactive than carbon, such as aluminium, must be extracted by electrolysis.

Worked example 2: iron from its ore. $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$. Iron oxide loses oxygen, so it is reduced. Carbon gains oxygen, so it is oxidised. Check the atoms: 4 Fe, 6 O and 3 C on each side.

Higher tier: electron transfer. Oxidation and reduction can also be described with electrons. Each **electron** 电子 carries a negative charge. Oxidation is loss of electrons; reduction is gain of electrons (remember OIL RIG: oxidation is loss, reduction is gain). In the zinc and copper sulfate reaction, the sulfate ions do not change, so leave them out:

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

$$Zn \rightarrow Zn^{2+} + 2e^- \qquad Cu^{2+} + 2e^- \rightarrow Cu$$

Zinc atoms lose electrons and are oxidised; copper ions gain them and are reduced. An equation like this, which shows only the particles that change, is an ionic equation. In the reaction of magnesium with hydrochloric acid, magnesium atoms lose two electrons and two H^+ ions gain them; the chloride ions do not change.

Acids, alkalis and salts

An acid in water releases hydrogen ions, H^+. A **base** 碱 is a substance that neutralises an acid; a base that dissolves in water is an alkali, and its solution contains hydroxide ions, OH^-. In **neutralisation** 中和, an acid reacts with a base to make a salt and water. The ionic equation is always:

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

Naming salts. The first part of the name comes from the metal, base or carbonate; the second part comes from the acid. Hydrochloric acid makes a **chloride** 氯化物, nitric acid makes a **nitrate** 硝酸盐, and sulfuric acid makes a **sulfate** 硫酸盐.

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Reaction	Products	Example
acid + metal	salt + hydrogen	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$
acid + metal hydroxide	salt + water	$NaOH + HCl \rightarrow NaCl + H_2O$
acid + metal oxide	salt + water	$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
acid + **carbonate** 碳酸盐	salt + water + carbon dioxide	$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$

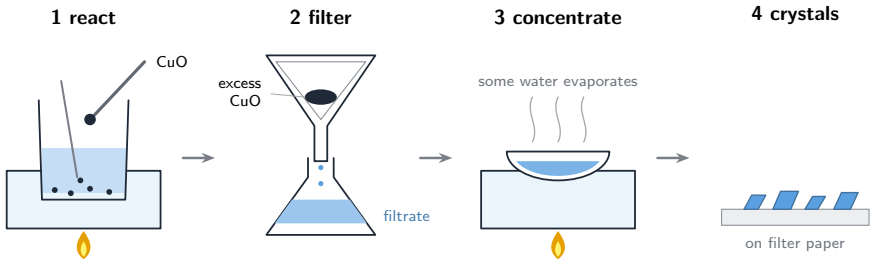
Worked example 3: the formula of a salt. Calcium nitrate contains Ca^{2+} and NO_3^- ions. Two nitrate ions (2×-1) are needed to balance one calcium ion (+2), so the formula is $Ca(NO_3)_2$, not $CaNO_3$.

■ Required practical: making copper sulfate crystals

The aim is a pure, dry sample of a soluble salt from an **insoluble** 不溶 base.

1. Warm dilute sulfuric acid gently. Add copper(II) oxide a little at a time, stirring, until some black solid stays at the bottom. This excess shows that all the acid has reacted.
2. **Filter** 过滤 the mixture. The extra copper oxide stays in the filter paper; the blue **filtrate** 滤液 is copper sulfate solution.
3. Heat the filtrate gently in an evaporating basin over a water bath until it is more concentrated, but do not boil it dry.
4. Leave it to cool, so that the salt can **crystallise** 结晶.
5. Filter off the crystals (each **crystal** 晶体 has flat faces and straight edges), wash them with a little cold water, and dry them between filter papers.

Wear eye protection, and take care with hot apparatus. Using an excess of the base is the key idea: it makes sure no acid is left in the salt, and the extra solid is easy to remove because it does not dissolve.



$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
filtering cannot remove dissolved acid, so use excess oxide; do not boil the solution dry

Choose separation from solubility: filter excess solid, then crystallise the dissolved salt.

The pH scale

The pH scale runs from 0 to 14. A solution with pH below 7 is acidic, pH 7 is neutral, and above 7 is alkaline. **Universal indicator** 通用指示剂 changes colour to show the approximate pH; a pH probe gives a more exact reading. Adding water to an acid makes it less concentrated, so its pH rises towards 7, but it never becomes alkaline.

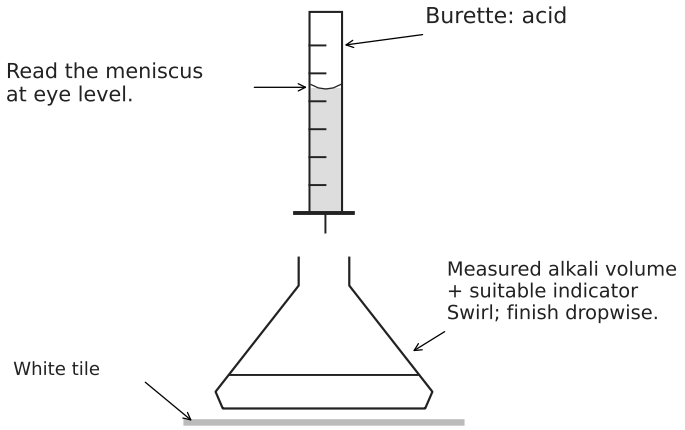
Required practical (chemistry only): titration

A **titration** 滴定 finds exactly how much acid reacts with a known volume of alkali.

1. Use a **pipette** 移液管 with a pipette filler to put 25.0 cm^3 of alkali in a **conical flask** 锥形瓶. Add a few drops of indicator, and stand the flask on a white tile.
2. Fill a **burette** 滴定管 with acid. Read the bottom of the **meniscus** 弯月面, the curved surface of the liquid, at eye level.
3. Add the acid while swirling the flask, slowly and then drop by drop near the end, until the indicator just changes colour. Record the final reading. The volume added is the titre.
4. Do a rough titration first, then repeat accurately until you have **concordant titres** 相符滴定体积, results within 0.10 cm^3 of each other. Calculate the mean of these and leave out the rough result.

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Titre = final burette reading − initial burette reading



A burette delivers acid into a measured alkali sample; read both burette levels.

Worked example 4 (higher tier). 20.0 cm^3 of 0.100 mol/dm^3 sodium hydroxide reacts with 25.0 cm^3 of sulfuric acid. Find the concentration of the acid in mol/dm^3 and in g/dm^3 (Mr of H_2SO_4 = 98).

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

- Moles of NaOH = $0.100 \times 0.0200 = 0.00200$ mol.
- The ratio is 1 : 2, so moles of $H_2SO_4 = 0.00200/2 = 0.00100$ mol.
- Concentration = $0.00100/0.0250 = 0.0400$ mol/dm^3.
- In g/dm^3: $0.0400 \times 98 = 3.92$ g/dm^3.

Higher tier: strong and weak acids

A **strong acid** 强酸, such as hydrochloric, nitric or sulfuric acid, is completely **ionised** 电离 in water: every molecule releases its H^+. A **weak acid** 弱酸, such as ethanoic acid (in vinegar), citric acid (in lemons) or carbonic acid (in fizzy drinks), is only partly ionised. At the same concentration, a strong acid has a lower pH.

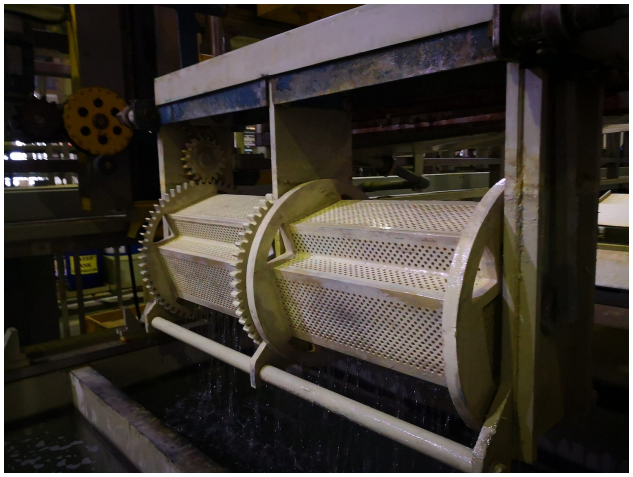
Strong and weak are not the same as concentrated and dilute. Strength tells you what fraction of the acid is ionised; concentration tells you how much acid is in each dm^3. A dilute strong acid is still a strong acid.

Each fall of one pH unit means ten times more H^+ ions. **Worked example 5:** pH 2 has $10 \times 10 \times 10 = 1000$ times the H^+ concentration of pH 5.

Electrolysis

An **ionic compound** 离子化合物 conducts electricity when its ions can move: when it is **molten** 熔融 or dissolved in water. A liquid or solution that conducts in this way is an **electrolyte** 电解质. In a solid **lattice** 晶格 the ions are fixed and cannot carry charge.

In **electrolysis** 电解, an electric current breaks down an electrolyte. A positive ion, or **cation** 阳离子, moves to the negative electrode, the **cathode** 阴极, and gains electrons. A negative ion, or **anion** 阴离子, moves to the positive electrode, the **anode** 阳极, and loses electrons. In the liquid the ions carry the charge; in the wires the electrons do.

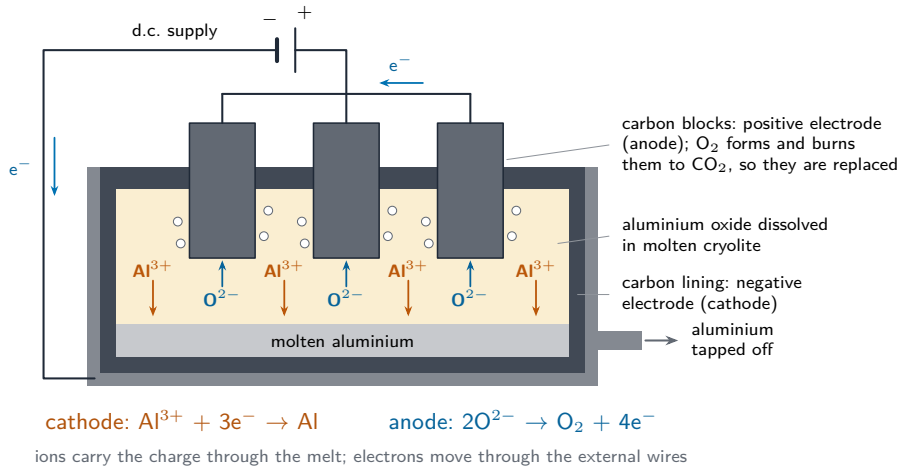


Electrolysis is used to electroplate an object with a thin, shiny layer of metal such as chromium

Molten compounds. With an **inert electrode** 惰性电极, one that does not react, a molten compound of two elements splits into them. Molten lead bromide gives lead at the cathode and bromine at the anode; molten zinc chloride gives zinc and chlorine.

Extracting aluminium. Aluminium is more reactive than carbon, so it is extracted by electrolysis. Aluminium oxide melts at over 2,000 °C, so it is dissolved in molten **cryolite** 冰晶石, which lowers the working temperature to about 950 °C and saves energy. Aluminium forms at the cathode and collects as a liquid at the bottom of the cell. Oxygen forms at the carbon anodes and reacts with them to make carbon dioxide, so the anodes burn away and must be replaced. The process uses a great deal of electricity, which is why aluminium smelters are often built near cheap hydroelectric power.

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The aluminium extraction cell: ions carry charge through the melt, electrons through the wires; half-equations at each electrode.

Solutions in water. Water also provides H^+ and OH^- ions, so there are two possible products at each electrode.

- At the cathode, the metal forms if it is less reactive than hydrogen (such as copper). Otherwise hydrogen forms.
- At the anode, if the solution contains a **halide** 卤化物 ion (chloride, bromide or iodide), the **halogen** 卤素 forms. Otherwise oxygen forms.

Electrolyte	Cathode	Anode
sodium chloride solution	hydrogen	chlorine
copper(II) chloride solution	copper	chlorine
sodium sulfate solution	hydrogen	oxygen
copper(II) sulfate solution	copper	oxygen

> **Worked example 6.** Predict the products of electrolysing potassium bromide solution. Potassium is more reactive than hydrogen, so hydrogen forms at the cathode. Bromide is a halide, so bromine forms at the anode.

■ Required practical: electrolysis of solutions

Before testing, write a hypothesis for each solution using the rules above. Use inert electrodes, a low-voltage d.c. supply, and keep the electrodes, their spacing and the time the same for each solution. Watch each electrode, then test the products:

- hydrogen burns with a squeaky pop when a **lighted splint** 燃着的木条 is held near it;
- oxygen relights a **glowing splint** 带火星的木条;
- chlorine bleaches damp **litmus paper** 石蕊试纸;
- copper appears as a pink-brown coating on the cathode.

Wear eye protection, collect only small volumes of gas, keep the room well ventilated, and never smell chlorine. Wash the electrodes between solutions.

■ Higher tier: half equations

A **half equation** 半方程 shows what happens at one electrode. Balance the atoms first, then add electrons to balance the charge. Electrons appear on the left for reduction (at the cathode) and on the right for oxidation (at the anode).

$$Al^{3+} + 3e^- \rightarrow Al \qquad 2O^{2-} \rightarrow O_2 + 4e^-$$

$$Cu^{2+} + 2e^- \rightarrow Cu \qquad 2Cl^- \rightarrow Cl_2 + 2e^-$$

$$2H^+ + 2e^- \rightarrow H_2 \qquad 4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

> **Worked example 7: checking a half equation.** In $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$, the left has 4 O, 4 H and a total charge of -4. The right has $2 + 2 = 4$ O, 4 H, and a charge of -4 from the four electrons. Both atoms and charge balance.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.4, and the required practicals handbook.
- Smithsonian Institution, "The Aluminum Apex of the Washington Monument" (1884 capstone, 100 troy ounces).
- C. M. Hall, US Patent 400,766, "Process of Reducing Aluminium from its Fluoride Salts by Electrolysis" (filed July 1886), and P. Héroult, French patent 175,711 (April 1886).
- Royal Society of Chemistry, *Practical Chemistry*: "Preparing a salt" and "Electrolysis of solutions".

Common errors

- Saying copper reacts with dilute acid; it is below hydrogen and does not.
- Writing salt formulae without balancing the charges, such as $CaNO_3$ instead of $Ca(NO_3)_2$.

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- Boiling the copper sulfate solution dry; the crystals then decompose and are not pure.
- Confusing strong with concentrated, or weak with dilute.
- Saying electrons flow through the electrolyte; ions carry the charge in the liquid.
- Predicting sodium at the cathode for sodium chloride solution; hydrogen forms because sodium is more reactive than hydrogen.
- Mixing up the electrodes: cations go to the cathode (negative), anions to the anode (positive).

Key terms

- **Reactivity series**: a list of metals in order of how readily they react.
- **Displacement**: a reaction in which a more reactive metal replaces a less reactive one in a compound.
- **Oxidation**: gain of oxygen or (higher tier) loss of electrons.
- **Reduction**: loss of oxygen or (higher tier) gain of electrons.
- **Neutralisation**: the reaction of an acid with a base to form a salt and water.
- **Strong acid**: an acid that is completely ionised in water.
- **Electrolysis**: the breaking down of an electrolyte by an electric current.
- **Half equation**: an equation that shows the change at one electrode, with electrons.

Check yourself

1. Metal Y reacts with dilute acid but does not displace iron from iron sulfate. Where could Y be in the reactivity series?
2. Name the salt made from zinc oxide and hydrochloric acid, and write a word equation.
3. Why is excess copper oxide added when making copper sulfate crystals?
4. (HT) What is the difference between a dilute strong acid and a concentrated weak acid?
5. Predict the products at each electrode when copper(II) chloride solution is electrolysed.
6. Why is aluminium oxide dissolved in cryolite?
7. (HT) Write the half equation for the formation of bromine at the anode.

Answers: 1. Below iron but above hydrogen (it reacts with acid), so between iron and hydrogen. 2. Zinc chloride; zinc oxide + hydrochloric acid → zinc chloride + water. 3. To make sure all the acid reacts; the extra solid is insoluble, so it is easily filtered off. 4. A dilute strong acid has little acid per dm^3, but all of it is ionised; a concentrated weak acid has a lot of acid per dm^3, but only a small fraction is ionised. 5. Copper at the cathode, chlorine at the anode. 6. It lowers the melting point, so less energy is needed to keep the electrolyte molten. 7. $2Br^- \rightarrow Br_2 + 2e^-$.

5 · Energy changes

AQA · GCSE · Chemistry

Energy changes: follow the energy transfer

Why it matters. The Apollo spacecraft that took astronauts to the Moon got their electricity from three hydrogen–oxygen fuel cells. Each **fuel cell** 燃料电池 combined the two gases to make electricity, and the only product, pure water, was used for drinking. On 13 April 1970, two days into the flight of Apollo 13, an oxygen tank exploded. Without oxygen the fuel cells stopped, and with them most of the spacecraft's power. The three astronauts survived by moving into the small lunar module and using its batteries, while engineers on Earth worked out how to stretch every amp-hour until they splashed down four days later. Whether a reaction gives out energy or takes it in, and how that energy can be turned into electricity, is the subject of this topic.

In this unit you will:

- classify reactions as exothermic or endothermic from temperature changes, and judge products such as hand warmers;
- carry out the required practical on temperature changes in reacting solutions;
- draw and read reaction profiles, including activation energy;
- (higher tier) calculate energy changes from bond energies;
- (chemistry only) explain how cells, batteries and hydrogen fuel cells produce electricity.

AQA GCSE Chemistry (8462) topic 4.5, tested on Paper 1. Section 4.5.1 is also in Combined Science: Trilogy (8464); cells and fuel cells (4.5.2) are chemistry only. Bond energy calculations and fuel-cell half equations are higher tier.

Exothermic and endothermic reactions

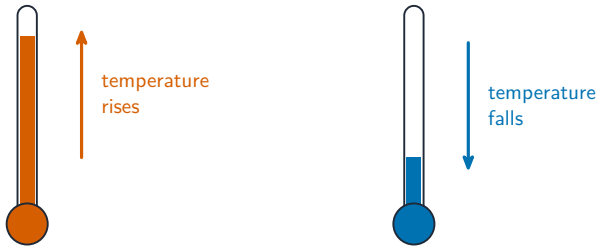
Energy is conserved in every reaction: it is not made or destroyed, only moved between the chemicals and their **surroundings** 周围环境.

An **exothermic** 放热 reaction transfers energy to the surroundings, so the temperature rises. **Combustion** 燃烧, many **oxidation** 氧化 reactions and **neutralisation** 中和 are exothermic. A **hand warmer** 暖手宝 uses an exothermic reaction: iron powder slowly oxidises in air and keeps your hands warm for hours.

An **endothermic** 吸热 reaction takes in energy from the surroundings, so the temperature falls. **Thermal decomposition** 热分解 is endothermic, and so is the reaction between citric acid and sodium hydrogencarbonate. An **instant cold pack** 速冷袋 for sports injuries uses an endothermic change: squeezing the pack breaks an inner bag, and ammonium nitrate dissolves in water and cools it.



An instant cold pack uses an endothermic reaction that takes in heat.



exothermic

heat given out, so the surroundings get hotter ($\Delta H < 0$)

endothermic

heat taken in, so the surroundings get colder ($\Delta H > 0$)

An exothermic reaction warms the surroundings; an endothermic reaction cools them

A temperature change alone does not prove a chemical reaction happened. **Dissolving** 溶解 is a physical change, but it can also make a solution warmer or colder.

> **Worked example 1.** A mixture changes from 21.0 °C to 15.5 °C. The temperature change is final − initial = 15.5 − 21.0 = −5.5 °C. The mixture got colder, so the reaction is endothermic: the chemicals took energy from the water around them.

Judging a product. A good hand warmer must get warm enough, stay warm long enough, and be safe and cheap. A good cold pack must cool quickly and be ready when someone is injured. When a question gives you data on temperature, time, cost and whether the pack can be reused, use it to compare the products for their purpose. A reaction being exothermic does not by itself make it a good hand warmer.

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Required practical: temperature changes

The aim is to see how one variable, such as the volume of alkali added to an acid, affects the temperature change.

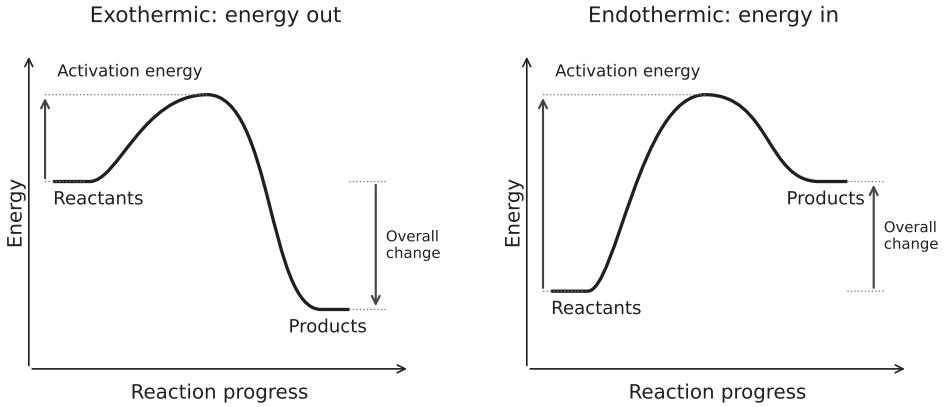
1. Put 30 cm^3 of dilute hydrochloric acid in a polystyrene cup standing in a beaker, so it cannot fall over. Measure its temperature.
2. Add 5 cm^3 of sodium hydroxide solution, put on the lid, stir, and record the highest temperature.
3. Keep adding 5 cm^3 portions, recording the total volume added and the highest temperature each time, up to 40 cm^3.
4. Repeat the whole experiment with fresh solutions and calculate means.

The polystyrene cup and lid reduce energy transfer to the air, but they do not stop it completely. Keep the acid volume, concentrations and starting temperature the same. Wear eye protection.

Plot temperature against the volume of alkali added. The temperature rises while the acid is being neutralised, then falls as more cold alkali is added after all the acid has reacted. Draw a **best-fit line** 最佳拟合线 through each part of the graph; where the two lines cross shows the volume of alkali that exactly neutralised the acid, which may lie between two of your readings.

Reaction profiles

Particles react only when they collide with enough energy. The minimum energy needed is the **activation energy** 活化能. A **reaction profile** 反应能量图 shows the energy of the reactants and products, and the "hill" between them.



Exothermic and endothermic reaction profiles with activation energy and overall change.

To draw one, label the vertical axis "energy" and the horizontal axis "progress of reaction" (not time). Draw the reactant and product levels, joined by a smooth curve that rises above both. Then add two arrows: activation energy from the reactants up to the peak, and the overall energy change from the reactants to the products.

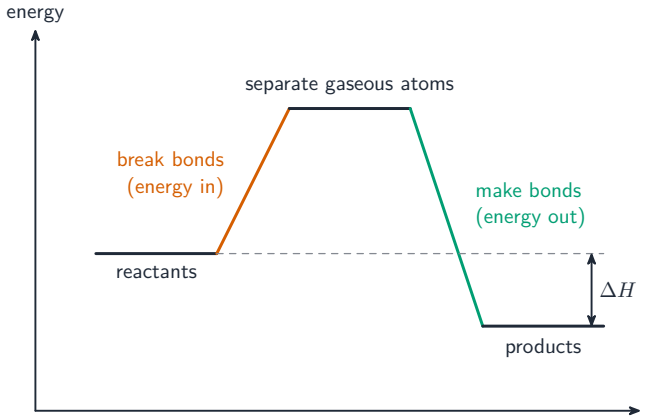
- In an exothermic reaction, the products are lower than the reactants.
- In an endothermic reaction, the products are higher than the reactants.
- Both kinds of reaction have an activation energy.

A **catalyst** 催化剂 provides a different route with a lower activation energy, so the peak is lower. The reactant and product levels, and so the overall energy change, stay the same.

> **Worked example 2.** On a profile, the reactants are at 40 units, the peak at 110 and the products at 15. Activation energy $= 110 - 40 = 70$ units. Overall change $= 15 - 40 = -25$ units. The products are lower, so the reaction is exothermic.

Bond energies (higher tier)

Breaking bonds takes in energy; making bonds gives out energy. Each **bond energy** 键能 is the energy needed to break one **mole** 摩尔 of that bond, in kJ/mol.



Breaking bonds takes energy in; making bonds gives energy out. ΔH is the difference between the two

overall energy change = energy to break bonds − energy released making bonds

If more energy is released making new bonds than was needed to break the old ones, the result is negative and the reaction is exothermic. If breaking needs more than making releases, the result is positive and the reaction is endothermic.

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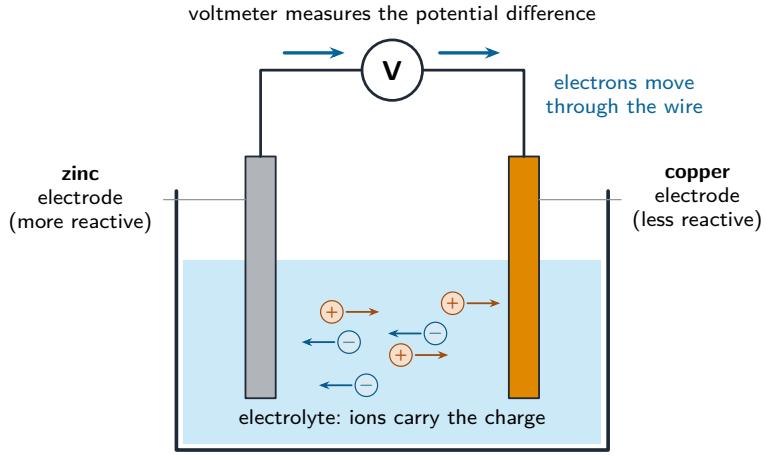
Worked example 3. Calculate the energy change for $2H_2 + O_2 \rightarrow 2H_2O$, using H–H = 436, O=O = 498 and O–H = 464 kJ/mol.

- Bonds broken: two H–H and one O=O **double bond** 双键: $2(436) + 498 = 1370$ kJ/mol.
- Bonds made: each water molecule has two O–H bonds, so two molecules have four: $4(464) = 1856$ kJ/mol.
- Overall change $= 1370 - 1856 = -486$ kJ/mol. The reaction is exothermic.

Count the bonds from structural formulae, and remember that the number in front of a formula multiplies every bond in that molecule.

Cells and batteries (chemistry only)

A **cell** 电池 uses a chemical reaction to produce electricity. A simple cell has two different metals as each **electrode** 电极, dipped in an **electrolyte** 电解质, and joined by wires in an **external circuit** 外电路. Electrons flow through the wire; ions carry the charge through the electrolyte. A **voltmeter** 电压表 measures the **potential difference** 电势差, or voltage.



voltage depends on the two metals, the electrolyte and the conditions

Simple zinc/copper cell with electrolyte and voltmeter; electron and ion paths distinguished.

The voltage depends on the two metals and the electrolyte. The bigger the difference in reactivity between the two metals, the bigger the voltage.

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Worked example 4. In the same electrolyte, a zinc–copper cell gives 1.10 V and an iron–copper cell gives 0.78 V. Zinc is more reactive than iron, so it is further from copper in the reactivity series and gives the higher voltage. To make this a fair test, only one electrode was changed; changing the electrolyte too would make the result impossible to explain.

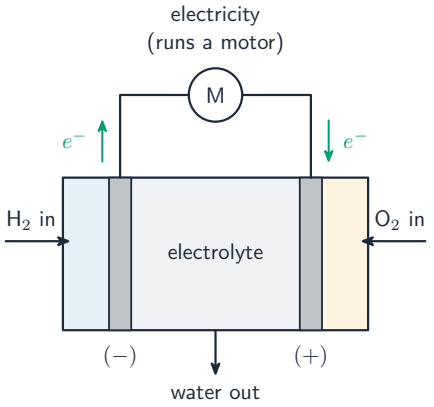
A **battery** 电池组 is two or more cells joined in **series** 串联, which adds their voltages. Three 1.2 V cells in series give $3 \times 1.2 = 3.6$ V.

In a **non-rechargeable** 不可充电 cell, such as an **alkaline battery** 碱性电池, the reaction stops when one of the reactants is used up. In a **rechargeable** 可充电 cell, such as the lithium-ion battery in a phone, an external current reverses the reactions so that the cell can be used again. Never try to recharge a non-rechargeable cell.

Hydrogen fuel cells (chemistry only)

A fuel cell is supplied with a fuel, such as hydrogen, and with oxygen from the air. It produces electricity for as long as the supply continues. In a hydrogen fuel cell, hydrogen is **oxidised** 被氧化 to make water:

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



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

A hydrogen–oxygen fuel cell turns chemical energy straight into electricity, with water as the only product

	Hydrogen fuel cell	Rechargeable battery
product	only water	none during use
refilling	a few minutes, but hydrogen stations are rare	charging takes much longer
storage	hydrogen is a flammable gas, stored under high pressure	the battery is heavy
lifetime	works while fuel is supplied	loses capacity after many charges

Neither choice is free of pollution. Most hydrogen today is made from natural gas, which releases carbon dioxide, and batteries need metals such as lithium and cobalt that must be mined. Use the data in a question to compare them for a stated purpose.

Higher tier: half equations. In a cell that supplies electricity, the electrode where oxidation happens is the **anode** 阳极 (negative) and the electrode where reduction happens is the **cathode** 阴极 (positive). In an acidic fuel cell:

$$\text{at the anode: } 2H_2 \rightarrow 4H^+ + 4e^-$$

$$\text{at the cathode: } O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$$

Hydrogen loses electrons and is oxidised; oxygen gains electrons and is **reduced** 被还原. Each line is a **half-equation** 半方程. Adding the two cancels the four electrons and four H^+ ions, leaving the overall equation. These electrode signs are the opposite of those in **electrolysis** 电解, where an external supply drives the reaction.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.5, and the required practicals handbook.
- NASA, *Apollo 13 Mission Operations Report* (1970) and the *Report of the Apollo 13 Review Board* (15 June 1970).
- NASA Glenn Research Center, "Fuel cells: a history" (Apollo command and service module power system).
- International Energy Agency, *Global Hydrogen Review 2024*, on how hydrogen is produced.

Common errors

- Saying that an endothermic reaction "gives out cold"; it takes in energy from the surroundings.

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- Drawing the activation energy arrow from the products, or the overall change from the peak.
- Saying a catalyst changes the overall energy change; it only lowers the activation energy.
- Saying bond making takes in energy; making bonds releases energy.
- Forgetting to multiply bonds by the numbers in front of formulae.
- Saying fuel cells produce no pollution at all; making the hydrogen can release carbon dioxide.

Key terms

- **Exothermic**: describes a reaction that transfers energy to the surroundings.
- **Endothermic**: describes a reaction that takes in energy from the surroundings.
- **Activation energy**: the minimum energy that particles need to react when they collide.
- **Reaction profile**: a diagram of the energy of reactants and products during a reaction.
- **Bond energy**: the energy needed to break one mole of a particular bond.
- **Cell**: a device that uses a chemical reaction to produce a potential difference.
- **Fuel cell**: a cell supplied with fuel and oxygen, which it oxidises to produce electricity.

Check yourself

1. The temperature of a mixture rises from 19 °C to 31 °C. Is the reaction exothermic or endothermic? What is the temperature change?
2. Why is a polystyrene cup with a lid used in the required practical?
3. On a reaction profile, the reactants are at 60 units, the peak at 150 and the products at 90. Find the activation energy and the overall change, and classify the reaction.
4. What effect does a catalyst have on a reaction profile?
5. (HT) For $H_2 + Cl_2 \rightarrow 2HCl$, use H–H = 436, Cl–Cl = 243 and H–Cl = 432 kJ/mol to calculate the energy change.
6. (Chemistry only) Why does a non-rechargeable cell eventually stop working?
7. (Chemistry only) Give one advantage and one disadvantage of hydrogen fuel cells compared with rechargeable batteries for cars.

Answers: 1. Exothermic; $31 - 19 = +12$ °C. 2. To reduce energy transfer to or from the surroundings, so the measured temperature change is more accurate. 3. Activation energy $= 150 - 60 = 90$ units; overall change $= 90 - 60 = +30$ units; endothermic. 4. It lowers the peak (the activation energy); the reactant and product levels do not change. 5. Broken $= 436 + 243 = 679$; made $= 2 \times 432 = 864$; change

= 679 − 864 = −185 kJ/mol (exothermic). 6. One of the reactants is used up. 7. For example: advantage, refuelling takes minutes and the only product is water; disadvantage, hydrogen is hard to store and few filling stations exist.

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6 · The rate and extent of chemical change

AQA · GCSE · Chemistry

The rate and extent of chemical change

Why it matters. On the evening of 2 May 1878, the Washburn "A" Mill in Minneapolis, then the largest flour mill in the United States, exploded. The blast was heard 15 kilometres away, destroyed the mill and the buildings around it, and killed 18 workers. Flour is just food, and a sack of it burns only slowly. But when fine flour dust fills the air, every tiny grain is surrounded by oxygen, and a single spark can make the whole cloud burn almost at once. The same substance reacted in a fraction of a second instead of hours, because its surface area had changed. This topic explains what makes reactions fast or slow, and why some reactions never quite finish.

In this unit you will:

- calculate mean rates of reaction from tables and graphs, and draw tangents to find instantaneous rates (higher tier);
- use collision theory to explain how concentration, pressure, surface area, temperature and catalysts change the rate;
- plan and interpret the required practical on concentration and rate, by two methods;
- describe reversible reactions and dynamic equilibrium;
- (higher tier) predict how changing conditions moves the position of an equilibrium.

AQA GCSE Chemistry (8462) topic 4.6, also in Combined Science: Trilogy (8464), tested on Paper 2. Calculating gradients of tangents, rates in mol/s and predicting equilibrium shifts are higher tier (HT).

Measuring the rate of a reaction

The **rate of reaction** 反应速率 tells you how fast reactants are used up or products are made. The **mean rate** 平均速率 over a period of time is:

$$\text{mean rate} = \frac{\text{quantity of reactant used or product formed}}{\text{time taken}}$$

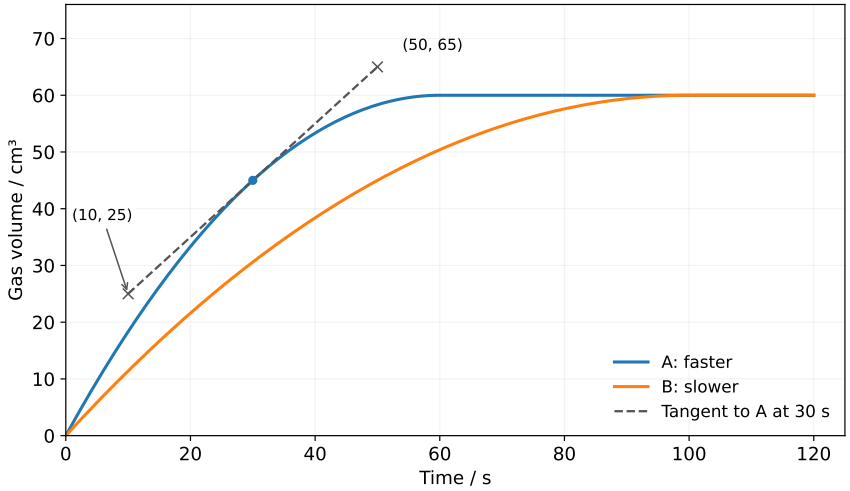
The quantity can be a mass (rate in g/s) or a gas volume (cm^3/s); at higher tier it can also be an amount in moles (mol/s).



A gas syringe measures the volume of gas made over time, which gives the rate of reaction

> **Worked example 1.** The volume of gas rises from 12 cm^3 at 10 s to 36 cm^3 at 30 s. Mean rate $= (36 - 12)/(30 - 10) = 24/20 = 1.2$ cm^3/s.

On a graph of product against time, the line rises steeply at first and then levels off as a reactant runs out. The steeper the line, the faster the reaction. The height where the line levels off shows how much product was made in total; that depends on the amount of the limiting reactant, not on the rate.



Gas volume against time for three concentrations, with a tangent showing instantaneous rate.

Tangents. The rate at one moment is the **instantaneous rate** 瞬时速率. To find it, draw a **tangent** 切线: a straight line that just touches the curve at that time and follows its direction. Make it long, so that you can read two points far apart on it. The **gradient** 斜率 of the tangent is the rate at that moment.

> **Worked example 2 (HT).** A tangent drawn at 30 s passes through (10 s, 25 cm^3) and (50 s, 65 cm^3). Gradient $= (65 - 25)/(50 - 10) = 40/40 = 1.0$ cm^3/s. That is the rate at 30 s.

> **Worked example 3 (HT).** The amount of a reactant falls from 0.018 mol to 0.006 mol in 40 s. The amount used is 0.012 mol, so the mean rate $= 0.012/40 = 0.00030$ mol/s.

What changes the rate?

Collision theory 碰撞理论 says that particles react only when they collide, and only if they collide with at least the **activation energy** 活化能. Anything that makes successful collisions happen more often makes the reaction faster.

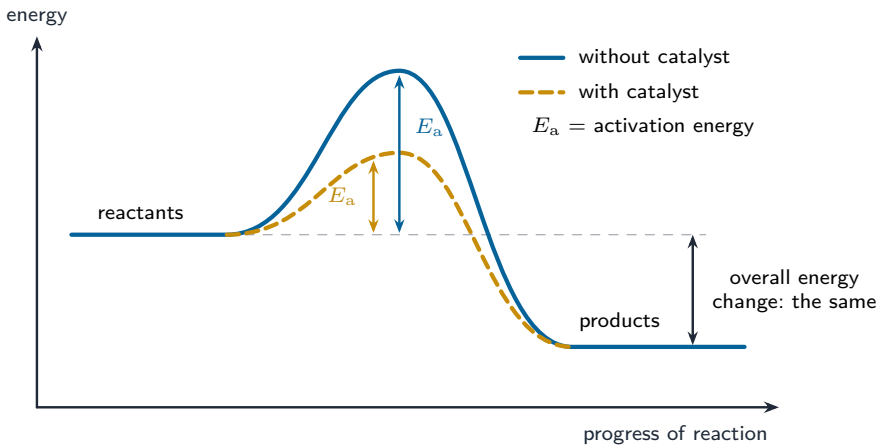
Change	Why the rate increases
higher **concentration** 浓度 of a solution	more particles in the same volume, so more frequent collisions
higher **pressure** 压强 of a gas	particles closer together, so more frequent collisions
smaller pieces of a solid	larger **surface-area-to-volume ratio** 表面积体积比: more particles exposed and able to collide
higher temperature	particles move faster, so they collide more often, and more of the collisions have enough energy
adding a **catalyst** 催化剂	provides a different route with a lower activation energy

Temperature has two effects, and the second is the bigger one: more of the collisions have enough energy to react. Concentration does not change the **kinetic energy** 动能 of the particles; it only makes collisions more frequent.

> **Worked example 4: the flour mill.** One cube of flour 1 cm across has a surface area of 6 cm^2. Grind it into dust grains 0.01 mm across and the same flour has a total surface area of about 6,000 cm^2: a thousand times more surface for oxygen to reach. That is why flour dust can explode while a sack of flour burns slowly.

Catalysts. A catalyst speeds up a reaction but is not used up, so it does not appear in the overall equation. Different reactions need different catalysts. Each **enzyme** 酶 is a biological catalyst: for example, amylase in your saliva speeds up the breakdown of starch.

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Catalysed and uncatalysed reaction profiles – lower activation energy, same overall change.

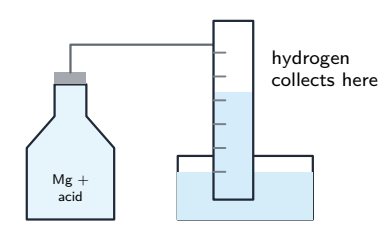
On a reaction profile, the catalysed route has a lower peak, but it starts and ends at the same energy levels. A catalyst makes the reaction reach its end sooner; it does not make more product.

Required practical: concentration and rate

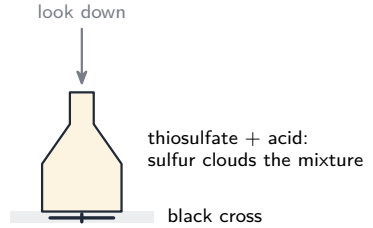
Before you start, write a **hypothesis** 假设, for example: "If the acid is more concentrated, the reaction will be faster, because there will be more frequent collisions."

Method 1: measuring a gas. Add a 3 cm length of cleaned **magnesium ribbon** 镁带 to 50 cm^3 of **dilute hydrochloric acid** 稀盐酸 in a conical flask. Quickly fit a bung with a **delivery tube** 导管 that leads into an upside-down **measuring cylinder** 量筒 full of water (or use a **gas syringe** 气体注射器). Record the volume of hydrogen every 10 seconds until it stops changing. Repeat with acid of a different concentration, keeping the length of magnesium, the volume of acid and the starting temperature the same. Plot both curves on the same axes: the more concentrated acid gives a steeper curve.

Keep flames away from hydrogen, wear eye protection, and never seal a flask that is making gas. If you fit the bung slowly, gas escapes at the start and your results will be too low.



gas-volume method: read the volume every 10 s until constant



turbidity method: time until the cross disappears

Schematic: collect and read the hydrogen volume over time, or time the cloudiness until the cross can no longer be seen

Method 2: watching for cloudiness. Sodium thiosulfate reacts with hydrochloric acid to make **sulfur** 硫, which makes the solution cloudy. This cloudiness is called **turbidity** 浑浊度. Draw a black cross on paper and stand a conical flask on it. Put 50 cm^3 of **sodium thiosulfate** 硫代硫酸钠 solution in the flask, add 10 cm^3 of dilute hydrochloric acid, swirl, and time how long it takes until you can no longer see the cross. To change the concentration, mix the 40 g/dm^3 stock solution with water, always to a total of 50 cm^3:

Stock solution (cm^3)	Water (cm^3)	Concentration before adding acid (g/dm^3)
10	40	8
20	30	16
30	20	24
40	10	32
50	0	40

Keeping the total volume the same means the depth of liquid you look through is the same each time. Repeat each concentration and calculate the mean time. Because the "end" is always the same amount of cloudiness, 1/time can be used as a measure of the rate.

> **Worked example 5.** One mixture takes 60 s and another takes 30 s. Then $1/60 = 0.017\ s^{-1}$ and $1/30 = 0.033\ s^{-1}$, so the second reaction is twice as fast.

The reaction gives off **sulfur dioxide** 二氧化硫, which can cause breathing problems, especially for people with asthma. Work with good **ventilation** 通风, wear eye protection, and do not lean over the flask. If one result is very different from the others, an **anomaly** 异常值, check for a reason before you leave it out, and say that you did.

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Reversible reactions

In a **reversible reaction** 可逆反应, the products can react to form the original reactants again. The symbol ⇌ shows this:

$$\text{ammonium chloride} \rightleftharpoons \text{ammonia} + \text{hydrogen chloride}$$

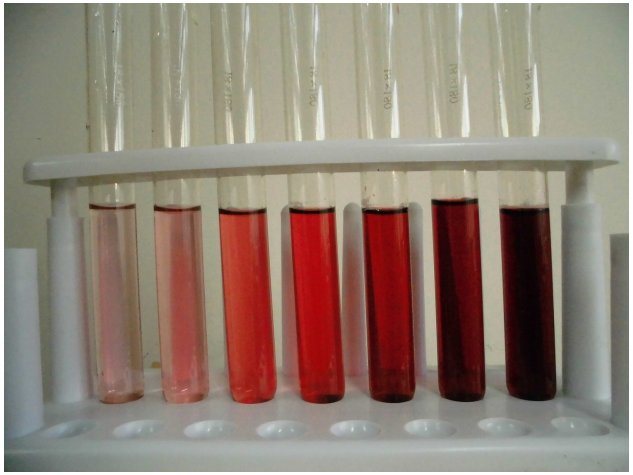
Heating **ammonium chloride** 氯化铵 breaks it down into two gases; when they cool, they join to form ammonium chloride again.

If a reaction is exothermic in one direction, it is endothermic in the other, and the same amount of energy is transferred. **Hydrated copper sulfate** 水合硫酸铜 is blue. Heating it drives off water and leaves white **anhydrous** 无水 copper sulfate; this **dehydration** 脱水 is endothermic. Adding water turns it blue again and makes it hot; this **hydration** 水合 is exothermic. White anhydrous copper sulfate is used as a test for water.

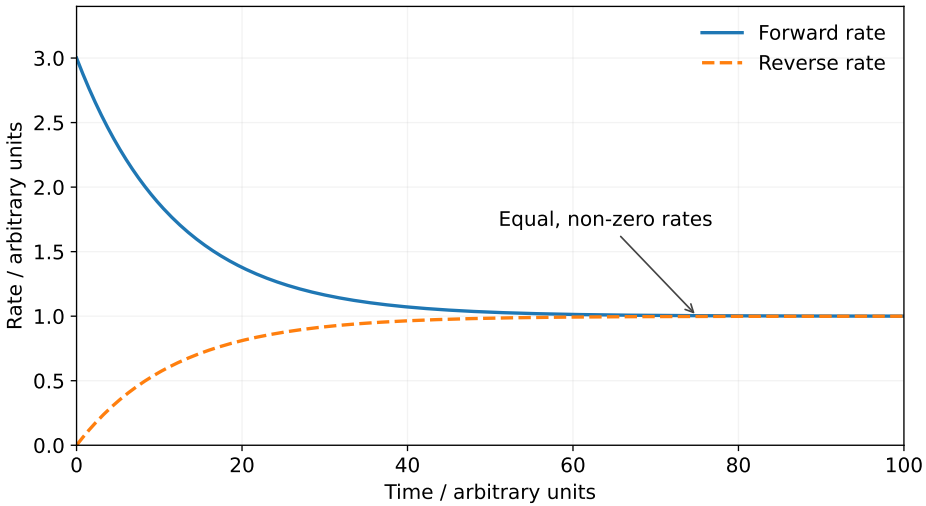
> **Worked example 6.** If dehydrating a sample takes in 72 kJ, rehydrating the same sample gives out 72 kJ.

Dynamic equilibrium

In a **closed system** 封闭体系, nothing can get in or out. When a reversible reaction happens in a closed system, it eventually reaches **dynamic equilibrium** 动态平衡: the forward and reverse reactions go on at exactly the same rate. The amounts of reactants and products then stay constant, although they are usually not equal.



Cobalt chloride changes colour as its equilibrium shifts.



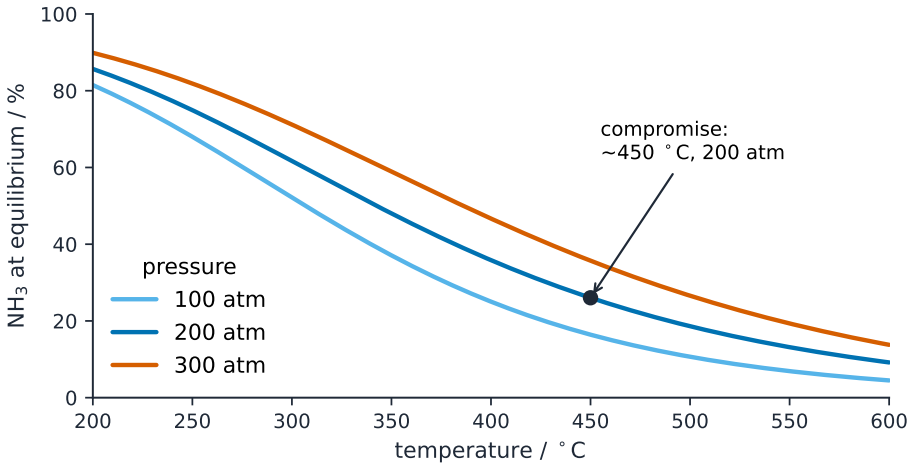
Equal continuing forward and reverse rates in a closed system.

At the start, only reactants are present, so the forward reaction is fast. As products build up, the reverse reaction speeds up and the forward reaction slows down, until the two rates are equal. The reactions do not stop; they balance. Think of an escalator that you walk down at the same speed as it moves up: you stay in the same place, but you are still moving.

Changing an equilibrium (higher tier)

The **equilibrium position** 平衡位置 describes how much product there is compared with reactant. The **Le Chatelier principle** 勒夏特列原理 says that if you change the conditions of a system at equilibrium, the system responds in a way that opposes the change.

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exothermic: lower T ↑ yield · higher P ↑ yield

Less ammonia at a higher temperature (the forward reaction is exothermic); more at a higher pressure (fewer gas molecules on the right)

Concentration. If you add more of a reactant, more product is made until equilibrium is reached again. If you remove product as it forms, more product is made. Adding product, or removing reactant, has the opposite effect.

Temperature. Raising the temperature favours the endothermic direction; lowering it favours the exothermic direction. So if the forward reaction is exothermic, heating gives less product at equilibrium, and cooling gives more.

Pressure (for gases). Raising the pressure favours the side with fewer gas molecules. Count the numbers in front of the gas formulae.

Worked example 7: making ammonia. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. The forward reaction is exothermic.

- There are 4 gas molecules on the left and 2 on the right, so high pressure gives more ammonia.
- Low temperature gives a higher **yield** 产率 of ammonia, but then the reaction is very slow. Industry uses a compromise of about 450 °C, with an iron catalyst to speed it up.
- A catalyst does not change the equilibrium position; it only helps the system reach equilibrium sooner.

Worked example 8. In $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, there are 2 gas molecules on each side, so changing the pressure does not move the equilibrium.

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Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.6, and the required practicals handbook (Required practical 5).
- Minnesota Historical Society, "Washburn A Mill explosion, 2 May 1878", and Mill City Museum, Minneapolis.
- US Chemical Safety Board, *Combustible Dust Hazard Study* (2006).
- H. L. Le Chatelier, "Sur un énoncé général des lois des équilibres chimiques", *Comptes rendus* 99 (1884).

Common errors

- Saying that a higher temperature only makes particles collide more often; more collisions also have enough energy to react.
- Saying that a catalyst gives more product; it only makes the reaction faster.
- Joining two points on the curve instead of drawing a tangent that touches it at one point.
- Saying that reactions stop at equilibrium; both reactions continue at the same rate.
- Saying that the amounts of reactants and products are equal at equilibrium; they are constant, not equal.
- Counting all molecules, including solids, when predicting the effect of pressure; count gas molecules only.

Key terms

- **Rate of reaction**: how fast a reactant is used up or a product is formed.
- **Collision theory**: particles react only when they collide with enough energy.
- **Activation energy**: the minimum energy that colliding particles need to react.
- **Catalyst**: a substance that speeds up a reaction without being used up.
- **Reversible reaction**: a reaction in which the products can react to form the reactants.
- **Dynamic equilibrium**: the state of a reversible reaction in a closed system when the forward and reverse rates are equal.
- **Le Chatelier principle**: a system at equilibrium responds to oppose any change in conditions.

Check yourself

1. 0.60 g of magnesium is used up in 40 s. What is the mean rate?
2. Explain, using collision theory, why powdered marble reacts faster with acid than marble chips.
3. Give two reasons why raising the temperature increases the rate.

4. In the cloudiness method, why is the total volume kept at 50 cm^3?
5. What is meant by "dynamic" in dynamic equilibrium?
6. (HT) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; the forward reaction is exothermic. What happens to the amount of SO_3 if the pressure is increased? And if the temperature is increased?

Answers: 1. 0.60/40 = 0.015 g/s. 2. The powder has a larger surface area to volume ratio, so more particles are exposed and collisions with acid particles are more frequent. 3. Particles move faster and collide more often; a greater proportion of collisions have at least the activation energy. 4. So that only the concentration changes, and the depth of liquid you look through stays the same. 5. The forward and reverse reactions are still happening, at the same rate. 6. Higher pressure: more SO_3 (3 gas molecules become 2). Higher temperature: less SO_3 (the endothermic reverse reaction is favoured).

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7 · Organic chemistry

AQA · GCSE · Chemistry

Organic chemistry: choose the reaction

Why it matters. On 27 March 1933, two chemists at ICI in Northwich, England, Eric Fawcett and Reginald Gibson, heated ethene gas with benzaldehyde at about 1,400 times atmospheric pressure. They were hoping for a new liquid. When they opened the vessel, they found a small amount of white, waxy solid on the walls. A tiny leak had let in a trace of oxygen, which started the ethene molecules joining into long chains. They had made poly(ethene), or polythene. By 1939 it was insulating the cables of British radar sets; today the world makes over 100 million tonnes of it a year, for bags, bottles and pipes. This topic follows carbon compounds from crude oil to fuels, to plastics, and to the polymers in your own body.

In this unit you will:

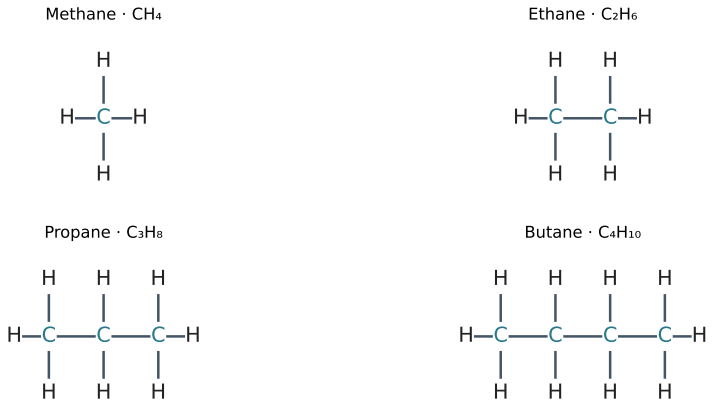
- explain how crude oil is separated into fractions, and how the properties of hydrocarbons change with chain length;
- balance combustion equations and explain why cracking is needed;
- (chemistry only) describe the reactions of alkenes, alcohols and carboxylic acids;
- (chemistry only) draw addition polymers from their monomers and back again;
- (chemistry only, higher tier) explain condensation polymers, and recognise the monomers of natural polymers.

AQA GCSE Chemistry (8462) topic 4.7, tested on Paper 2. Section 4.7.1 (crude oil, fuels and cracking) is also in Combined Science: Trilogy (8464). Sections 4.7.2 and 4.7.3 are chemistry only, and condensation polymers and amino acids are higher tier (HT).

Crude oil and the alkanes

Crude oil 原油 is a finite resource found in rocks. It formed over millions of years from the remains of plankton buried in mud, so it is not being replaced as fast as we use it. Crude oil is a mixture of a very large number of compounds. Most of them are each a **hydrocarbon** 碳氢化合物: a compound of hydrogen and carbon only.

Most of the hydrocarbons in crude oil are alkanes. Each **alkane** 烷烃 has only single bonds and follows the general formula C_nH_{2n+2}. Carbon atoms always form four bonds and hydrogen atoms one. Learn the first four: methane CH_4, ethane C_2H_6, propane C_3H_8 and butane C_4H_{10}.



Fully displayed hydrogen family; stated arrangements.

Worked example 1. An alkane has 6 carbon atoms. Its number of hydrogen atoms is $2 \times 6 + 2 = 14$, so its formula is C_6H_{14}.

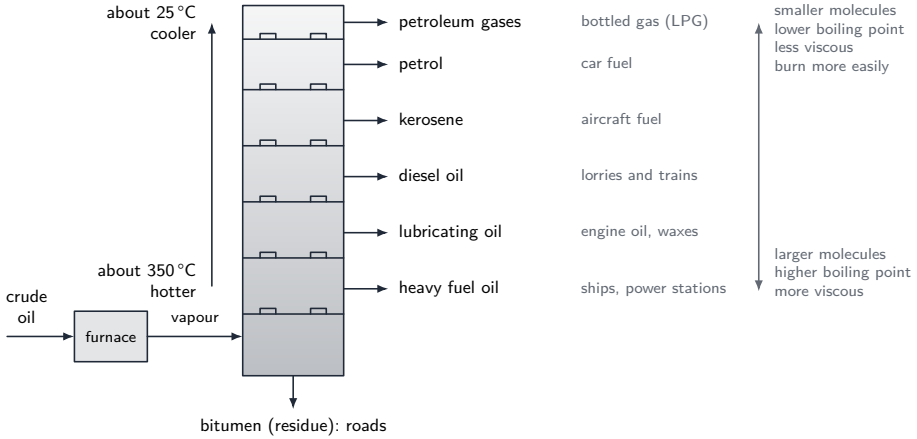
Fractional distillation and fuels

Fractional distillation 分馏 separates crude oil into parts with similar boiling points. The oil is heated until most of it evaporates, and the vapour enters a tall column that is hot at the bottom and cooler at the top. As the vapours rise, each hydrocarbon condenses when it reaches a level cooler than its boiling point. Large molecules, with high boiling points, condense near the bottom; small molecules rise higher; the smallest leave at the top as gases. Each **fraction** 馏分 is a mixture of hydrocarbons with similar numbers of carbon atoms. Distillation is a physical process: no bonds inside the molecules are broken.

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An oil refinery separates crude oil into useful fuels by fractional distillation.



A fractionating column: small molecules condense high, large viscous fractions low.

Properties change with size. As hydrocarbon molecules get bigger:

- the boiling point increases;
- the **viscosity** 黏度 increases: the liquid is thicker and flows less easily;
- the **flammability** 可燃性 decreases: it is harder to set alight.

The fractions give us fuels: liquefied petroleum gas, petrol, kerosene for aircraft, diesel oil and heavy fuel oil for ships. They also give **feedstock** 化工原料, the starting materials for making solvents, lubricants, detergents and polymers.

Worked example 2. Fraction P flows through a funnel in 8 s and catches fire easily; fraction Q takes 24 s and is hard to light. P has the smaller molecules, so it would be a better fuel for a camping stove, which needs to light quickly.

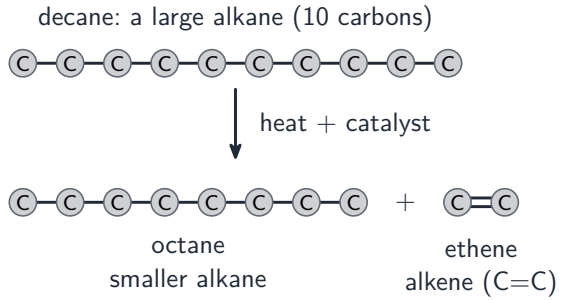
Combustion. In complete combustion, the carbon in a hydrocarbon is oxidised to carbon dioxide and the hydrogen to water, and energy is released. Balance carbon first, then hydrogen, then oxygen.

Worked example 3. Propane, C_3H_8, has 3 C atoms and 8 H atoms, so it makes $3CO_2$ and $4H_2O$. These contain $3 \times 2 + 4 = 10$ oxygen atoms, which need $5O_2$:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Cracking and alkenes

Crude oil contains more large hydrocarbons than we need, and fewer small ones such as petrol. **Cracking** 裂解 breaks long hydrocarbon molecules into shorter, more useful ones. In catalytic cracking, the vapour is passed over a hot catalyst; in steam cracking, it is mixed with steam at a very high temperature. Cracking makes smaller alkanes, which are useful fuels, and alkenes, which are used to make polymers and other chemicals.



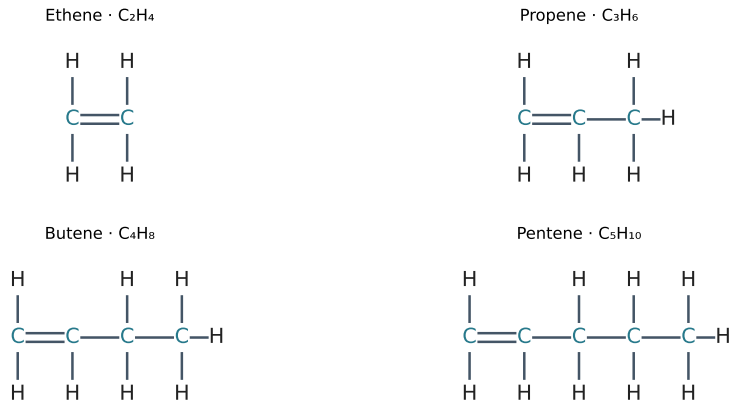
$$C_{10}H_{22} \longrightarrow C_8H_{18} + C_2H_4$$

Cracking breaks one large alkane into a smaller alkane plus an alkene (with a C=C), using heat and a catalyst

Worked example 4. Balance $C_{10}H_{22} \rightarrow C_8H_{18} + X$. X has $10 - 8 = 2$ carbon atoms and $22 - 18 = 4$ hydrogen atoms, so X is C_2H_4, ethene.

An **alkene** 烯烃 contains a carbon–carbon double bond, C=C. Alkenes are **unsaturated** 不饱和的: they have two fewer hydrogen atoms than the alkane with the same number of carbons, and their general formula is C_nH_{2n}. The first four are ethene C_2H_4, propene C_3H_6, butene C_4H_8 and pentene C_5H_{10}.

Testing for alkenes. Shake the hydrocarbon with orange bromine water. An alkene turns it colourless; an alkane leaves it orange.



Fully displayed alkene family; stated arrangements.

Reactions of alkenes (chemistry only)

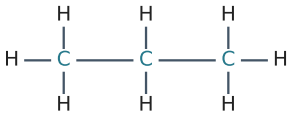
The C=C double bond is the **functional group** 官能团 of the alkenes: the group of atoms that decides how they react. Alkenes burn, but with smoky flames, because combustion in air is often incomplete. With enough oxygen they burn completely: $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$.

In an **addition** 加成 reaction, the double bond opens and atoms join to the two carbons that were double-bonded. Nothing is lost.

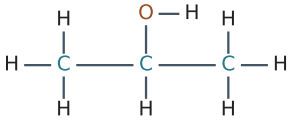
- With hydrogen, using a nickel catalyst at about 60 °C, an alkene becomes an alkane.
- With steam, at a high temperature and pressure with a phosphoric acid catalyst, it becomes an alcohol: H adds to one carbon and OH to the other.
- With chlorine, bromine or iodine, one halogen atom adds to each carbon. This is why alkenes decolourise bromine water.

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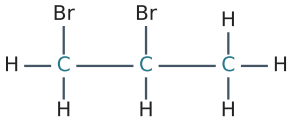
Propene + H_2 → propane · heated nickel



Propene + steam → shown OH-position product · acid catalyst



Propene + Br_2 → dibromo product · no H removed



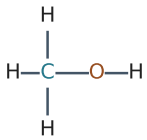
Displayed products from propene, with all atoms retained.

Worked example 5. Propene, $CH_2{=}CH{-}CH_3$, reacts with bromine. The double bond opens and one bromine atom joins each of the first two carbons: $CH_2Br{-}CHBr{-}CH_3$, 1,2-dibromopropane. Check that every carbon still has four bonds.

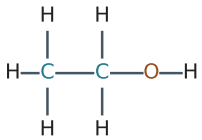
Alcohols (chemistry only)

Each **alcohol** 醇 contains the –OH group. Learn methanol, ethanol, propanol and butanol.

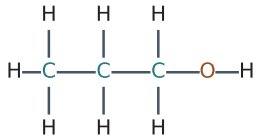
Methanol



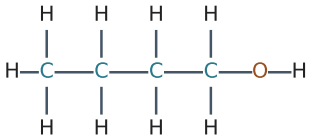
Ethanol



Propanol · primary form shown



Butanol · primary form shown



Fully displayed alcohol family; stated arrangements.

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These four alcohols:

- react with sodium, fizzing as hydrogen is given off;
- burn in air, releasing energy, so they are used as fuels;
- dissolve in water to form a neutral solution (the smaller ones mix completely);
- are oxidised to carboxylic acids by oxidising agents: ethanol becomes ethanoic acid, which is why wine turns to vinegar if it is left open.

Alcohols are used as fuels and as solvents, for example in hand gel and perfume.

Worked example 6. Balance the combustion of ethanol, C_2H_5OH. Products: $2CO_2$ and $3H_2O$, which contain 7 oxygen atoms. Ethanol already has 1, so 6 more are needed: $3O_2$.

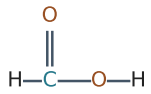
$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

Fermentation. **Fermentation** 发酵 turns sugar solution into ethanol using yeast. The conditions are warm (about 30 °C), with no oxygen, and in water. Enzymes in the yeast convert glucose into ethanol and carbon dioxide: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$. If it is too cold the enzymes work slowly; if it is too hot they are destroyed. An airlock lets the carbon dioxide out but stops air getting in. The product is a dilute solution of ethanol, which must be distilled to make it more concentrated.

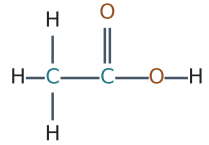
Carboxylic acids and esters (chemistry only)

Each **carboxylic acid** 羧酸 contains the –COOH group. The carbon in this group counts as part of the chain. The first four are methanoic, ethanoic, propanoic and butanoic acid; ethanoic acid is the acid in vinegar.

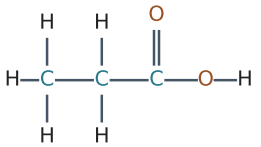
Methanoic acid



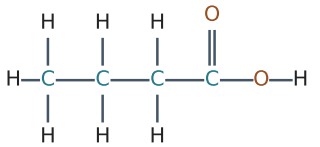
Ethanoic acid



Propanoic acid



Butanoic acid



Fully displayed acid family; stated arrangements.

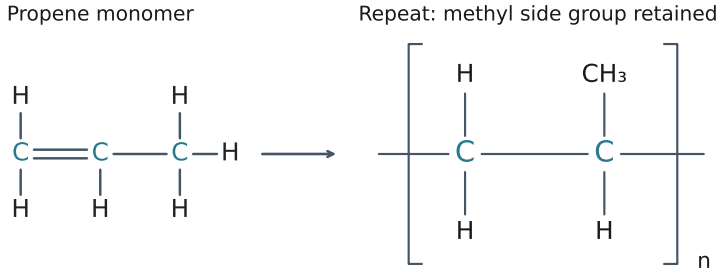
Carboxylic acids dissolve in water to make acidic solutions. They react with carbonates

to make a salt, water and carbon dioxide, so they fizz. With an alcohol, and an acid catalyst, they form an **ester** 酯 and water. Ethanoic acid and ethanol make **ethyl ethanoate** 乙酸乙酯, a sweet-smelling liquid used as a solvent in nail-varnish remover and in flavourings.

Higher tier: weak acids. Carboxylic acids are weak acids: in water, only a small fraction of their molecules release H^+ ions. So at the same concentration, ethanoic acid has a higher pH than hydrochloric acid, a strong acid that is completely ionised. Weak does not mean dilute.

Addition polymers (chemistry only)

Alkene molecules can join together into very long chains. Each small molecule is a **monomer** 单体, and this kind of reaction is **addition polymerisation** 加成聚合. The double bonds open and the monomers join end to end. All the atoms of the monomer stay in the polymer, and nothing else is made.



C_3H_6 per repeat · all atoms retained · no small by-product

Addition polymer: the C=C opens and the CH_3 side group hangs off every repeat unit

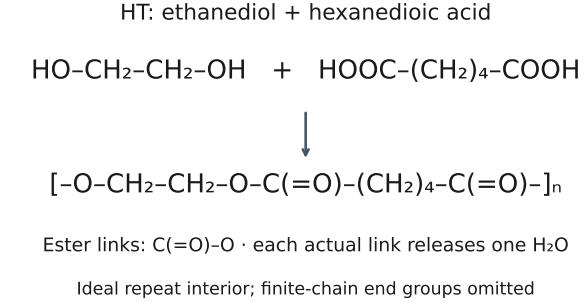
Worked example 7: drawing a polymer. For propene, $CH_2{=}CH{-}CH_3$:

1. Draw the two carbons of the double bond, now joined by a single bond.
2. Copy everything else that was attached to them: three H atoms and the CH_3 group.
3. Add a bond sticking out on each side, draw brackets across these bonds, and write n after the brackets.

The **repeating unit** 重复单元 is $-CH_2-CH(CH_3)-$, and the polymer is poly(propene). To find the monomer from a polymer, work backwards: take the two carbons in the repeating unit and put the double bond back between them.

Condensation polymers and natural polymers (chemistry only)

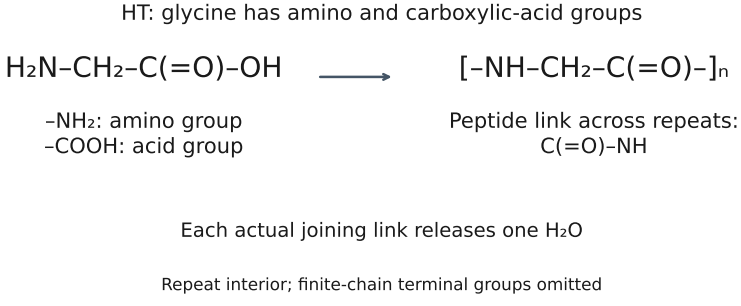
Higher tier: condensation polymers. In **condensation polymerisation** 缩聚, each monomer has two functional groups, one at each end. When two monomers join, a small molecule, usually water, is given off. A diol (two –OH groups) and a dicarboxylic acid (two –COOH groups) join to form a **polyester** 聚酯, with ester links between them. The polyester PET is used for drinks bottles and clothing fibres.



Condensation polymer: each ester link between diol and dicarboxylic acid releases one H_2O

Worked example 8 (HT). A short chain is made from 3 diol molecules and 3 dicarboxylic acid molecules, joined in a line. Six molecules in a line have 5 links between them, so 5 water molecules are given off.

Higher tier: amino acids. An **amino acid** 氨基酸 has two different functional groups: an amino group ($-NH_2$) and a carboxylic acid group (–COOH). The simplest, glycine, is $H_2N{-}CH_2{-}COOH$. Amino acids join by condensation to form a **polypeptide** 多肽, losing water at each link. Different amino acids combine in different orders to make proteins.



One amino acid with both groups links to itself through peptide links, releasing H_2O each time

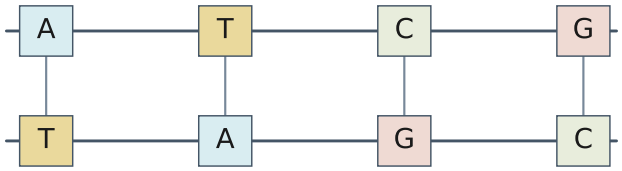
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Worked example 9 (HT). Four glycine molecules (Mr 75) join in a chain, with 3 links, losing 3 water molecules (Mr 18). The Mr of the chain is $4 \times 75 - 3 \times 18 = 300 - 54 = 246$.

Natural polymers. Many polymers are made by living things:

- DNA is two polymer chains wound into a **double helix** 双螺旋; each chain is made from four different monomers, each a **nucleotide** 核苷酸. DNA carries the genetic instructions for living things.
- Proteins are polymers of amino acids.
- Starch and cellulose are both polymers of glucose, but the glucose units are joined differently, so starch stores energy in plants while cellulose makes their cell walls.

DNA: two chains · four nucleotide types · flattened before twisting



Each base-labelled block represents a whole nucleotide; not to molecular scale

Two chains held by base pairs: A pairs with T and C pairs with G

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.7.
- ICI, laboratory notebook of E. W. Fawcett and R. O. Gibson (27 March 1933), and Royal Society of Chemistry, "The discovery of polythene" (National Historic Chemical Landmark, Northwich).
- J. R. Whinfield and J. T. Dickson, British patent 578,079 (1941), the first polyester (PET).
- J. D. Watson and F. H. C. Crick, "Molecular structure of nucleic acids", *Nature* (25 April 1953).

Common errors

- Saying fractional distillation breaks molecules; that is cracking. Distillation only separates them.
- Saying a fraction is a single compound; it is a mixture with similar boiling points.
- Saying bigger hydrocarbons are more flammable; they are less flammable.

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- Losing atoms when drawing an addition product; every atom of both reactants stays.
- Putting every carbon of the monomer into the polymer backbone; side groups like CH_3 hang off the chain.
- Saying condensation polymerisation gives off nothing; a small molecule such as water is lost at each link.
- Forgetting that fermentation needs no oxygen.

Key terms

- **Hydrocarbon**: a compound made of hydrogen and carbon only.
- **Alkane**: a saturated hydrocarbon with only single bonds, general formula C_nH_{2n+2}.
- **Alkene**: a hydrocarbon with a C=C double bond, general formula C_nH_{2n}.
- **Cracking**: breaking long hydrocarbon molecules into shorter, more useful ones.
- **Functional group**: the group of atoms that decides how a compound reacts.
- **Monomer**: a small molecule that can join with others to make a polymer.
- **Addition polymerisation**: the joining of alkene monomers by opening their double bonds.
- **Condensation polymerisation**: the joining of monomers with two functional groups, giving off a small molecule.

Check yourself

1. Write the formula of the alkane with 8 carbon atoms.
2. Why do the fractions condense at different heights in the column?
3. Balance the complete combustion of butane, C_4H_{10}.
4. Give two reasons why cracking is carried out.
5. How would you show that a gas from cracking contains an alkene?
6. (Chemistry only) Name the product when ethene reacts with steam, and give the conditions.
7. (Chemistry only) Draw the repeating unit of poly(chloroethene), made from $CH_2{=}CHCl$.
8. (Chemistry only, HT) Name the small molecule lost when a polyester forms.

Answers: 1. C_8H_{18}. 2. The column is hotter at the bottom; each hydrocarbon condenses where the temperature falls below its boiling point, and larger molecules have higher boiling points. 3. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$. 4. To make more small molecules such as petrol, which are in high demand, and to make alkenes for polymers. 5. Bubble it through bromine water: it turns from orange to colourless. 6. Ethanol; high temperature and pressure with a phosphoric acid catalyst. 7. $-CH_2-CHCl-$ inside brackets with continuation bonds, and n outside. 8. Water.

8 · Chemical analysis

AQA · GCSE · Chemistry

Chemical analysis: choosing evidence

Why it matters. In October 1858 a sweet-maker in Bradford, England, made a batch of peppermint humbugs. Sugar was expensive, so, like many sweet-makers at the time, he mixed in a cheap white powder called "daff", usually plaster of Paris. But the chemist's assistant who sold him the daff had picked up the wrong barrel: it was arsenic trioxide. About twenty people died and more than two hundred became ill. A white powder looks like any other white powder. Only a chemical test can tell you what is really there. Ten years later, the Pharmacy Act of 1868 made it a legal requirement to record every sale of poison. This topic is about the tests that chemists use to find out what a substance is, and whether it is pure.

In this unit you will:

- explain what "pure" means in chemistry, and use melting points as evidence of purity;
- explain what a formulation is;
- carry out paper chromatography and calculate Rf values;
- carry out the tests for four gases;
- (chemistry only) identify metal ions and negative ions with flame tests and precipitate tests;
- (chemistry only) explain the advantages of instrumental methods such as flame emission spectroscopy.

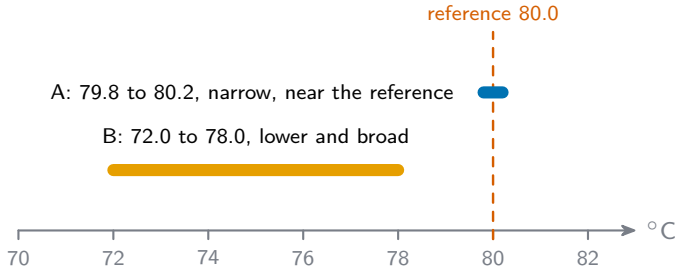
AQA GCSE Chemistry (8462) topic 4.8, tested on Paper 2. Purity, formulations, chromatography and the gas tests (4.8.1 and 4.8.2) are also in Combined Science: Trilogy (8464); the ion tests and instrumental methods (4.8.3) are chemistry only. It contains required practical 6 (chromatography) and, for chemistry only, required practical 7 (identifying ions).

In **chemical analysis** 化学分析, keep two things apart. An **observation** 观察结果 is what you see, such as "a white precipitate formed". A **conclusion** 结论 is what the observation tells you, such as "chloride ions are present".

Purity and formulations

In everyday language, "pure" means nothing has been added: "pure orange juice". In chemistry, a **pure substance** 纯净物 is a single **element** 元素 or a single compound, with nothing else mixed in. Orange juice is a **mixture** 混合物 of water, sugars, acids and much more, however natural it is.

A pure substance melts and boils at one exact temperature. An **impurity** 杂质 usually lowers the **melting point** 熔点 and makes the substance melt over a range of temperatures. So a melting point is a quick test of purity.



Sample A melts sharply at the reference; sample B melts lower and over a wide range, which supports impurity

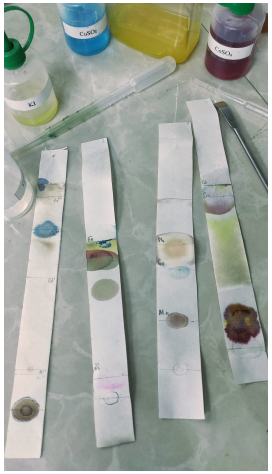
> **Worked example 1.** Pure aspirin melts at about 136 °C. A student's sample A melts between 135.5 and 136.5 °C; sample B melts between 124 and 132 °C. A is nearly pure: it melts sharply close to the true value. B is impure: it melts lower and over a wide range.

A **formulation** 配方 is a mixture that has been designed as a useful product. Each part has a job, and is added in a measured amount. A paint contains a **pigment** 颜料 for colour, a **binder** 黏合剂 to stick it to the surface, and a solvent to make it spread. Medicines, fuels, cleaning products, foods, each **alloy** 合金 and each **fertiliser** 化肥 are formulations. A paracetamol tablet contains only about 500 mg of the medicine itself; the rest is there to bind the tablet together and help it break up in the stomach.

Chromatography

Chromatography 色谱法 separates the substances in a mixture. In paper chromatography, the paper is the **stationary phase** 固定相, which does not move, and the **solvent** 溶剂 is the **mobile phase** 流动相, which moves up through the paper. Each substance is shared between the two phases in a different way. A substance that is more attracted to the solvent moves further up the paper.

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On real chromatograms, each substance in the mixture rises a different distance, leaving a separate coloured spot

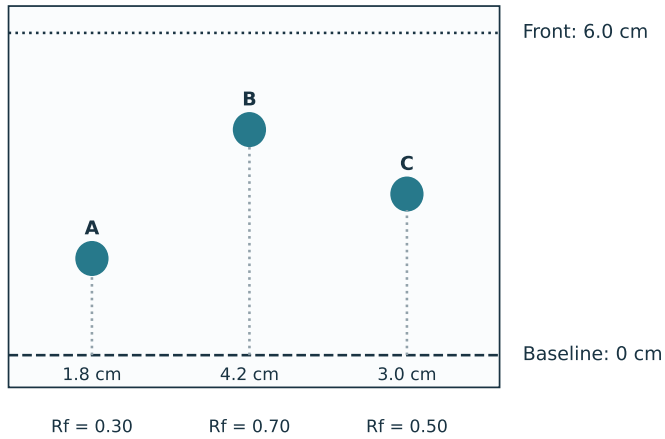
Required practical: paper chromatography

1. Draw a pencil line, the **baseline** 基线, about 1.5 cm from the bottom of the paper. Use pencil because ink would dissolve and run.
2. Put a small, concentrated spot of each sample and known reference on the line.
3. Stand the paper in a beaker with solvent below the baseline, so the spots do not dissolve into the solvent.
4. Let the solvent rise. Take the paper out before the solvent reaches the top, and mark the **solvent front** 溶剂前沿 straight away, before it dries.

For each spot, calculate the **Rf value** 比移值:

$$R_f = \frac{\text{distance from baseline to centre of spot}}{\text{distance from baseline to solvent front}}$$

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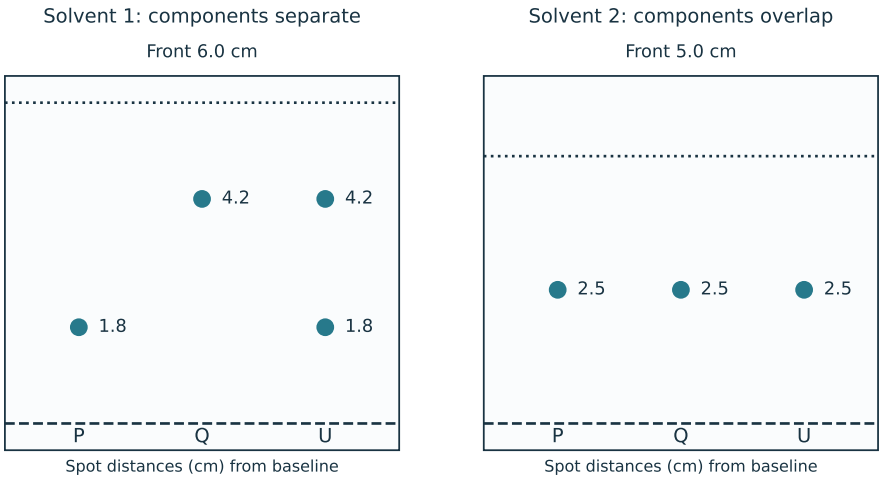
Rf = distance moved by spot ÷ distance moved by solvent: 1.8 ÷ 6.0 = 0.30, 4.2 ÷ 6.0 = 0.70, 3.0 ÷ 6.0 = 0.50

Rf has no units, and it is always less than 1, because no spot can move further than the solvent.

> **Worked example 2.** A spot moves 3.6 cm and the solvent front moves 6.0 cm. $R_f = 3.6/6.0 = 0.60$.

Interpreting a chromatogram. A pure substance gives one spot in any solvent. A mixture gives two or more spots, as long as the substances separate. If an unknown spot has the same Rf as a reference in the same solvent, they may be the same substance.

> **Worked example 3.** In solvent 1, dye P has Rf 0.30, dye Q has Rf 0.70, and an unknown ink U gives two spots at 0.30 and 0.70. So U contains P and Q. In solvent 2, P and Q both have Rf 0.50, and U gives one spot at 0.50. This single spot does not mean U is pure: in this solvent, P and Q just move together. A good chemist tries more than one solvent.

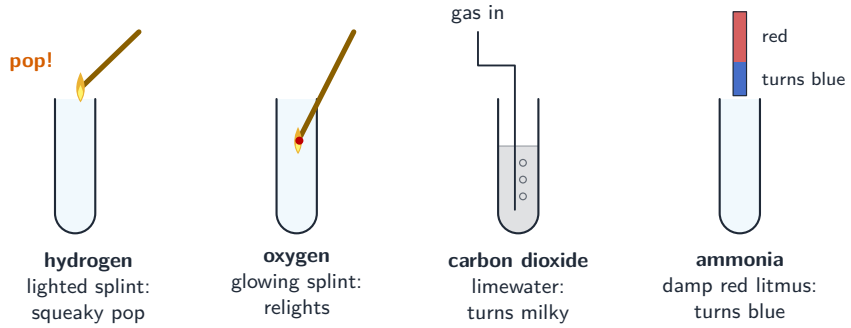


Dyes P and Q separate in solvent 1 but travel together in solvent 2, so ink U shows one spot.

Tests for gases

Gas	Test	Result
hydrogen	hold a burning splint at the mouth of a **test tube** 试管 of the gas	it burns with a squeaky pop
oxygen	put a glowing splint into the gas	the splint relights
carbon dioxide	bubble the gas through **limewater** 澄清石灰水 (calcium hydroxide solution)	the limewater turns milky (cloudy)
chlorine	hold damp litmus paper in the gas	the paper is bleached white

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Testing gases: hydrogen pops, oxygen relights a glowing splint

Worked example 4. A gas makes a glowing splint burst back into flame. It is oxygen. Notice the splint must be glowing, not burning: a burning splint cannot "relight". Chlorine is toxic, so it is tested only in small amounts in a fume cupboard.

Identifying metal ions (chemistry only)

Flame tests. For a **flame test** 焰色试验, dip a clean wire loop in the solid, hold it in a hot Bunsen flame, and observe the colour. Clean the loop between tests to avoid **contamination** 污染. In a mixture, one colour can hide another: the strong yellow of sodium can mask the others.

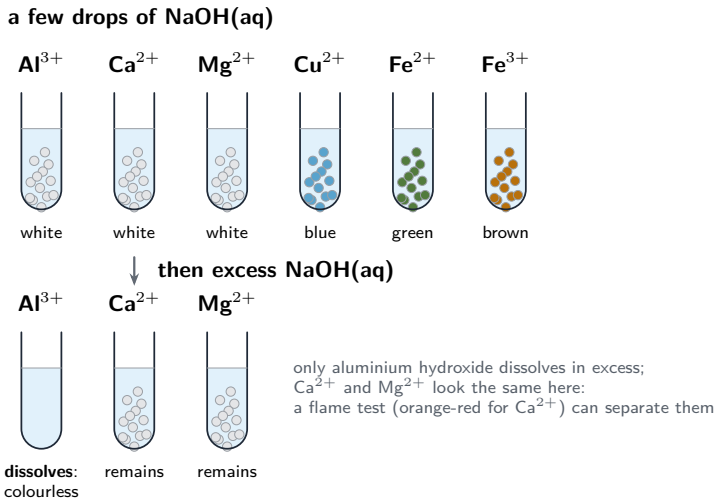


Flame tests: each metal ion gives its own colour because heat excites its electrons and, as they drop back down, they emit light of specific wavelengths

Metal ion	Flame colour
lithium Li^+	crimson
sodium Na^+	yellow
potassium K^+	lilac
calcium Ca^{2+}	orange-red
copper Cu^{2+}	green

Sodium hydroxide test. Add a few drops of sodium hydroxide solution to a solution of the compound. Many metal ions form an insoluble hydroxide, which appears as a solid **precipitate** 沉淀. Its colour identifies the metal. To tell the white precipitates apart, add an **excess** 过量 of sodium hydroxide.

Metal ion	Precipitate with a few drops	With excess sodium hydroxide
Al^{3+}	white	dissolves
Ca^{2+}	white	stays
Mg^{2+}	white	stays
Cu^{2+}	blue	stays
Fe^{2+}	green	stays
Fe^{3+}	brown	stays



NaOH precipitates: white for Al^{3+}, Ca^{2+}, Mg^{2+} (only Al^{3+} dissolves in excess), blue Cu^{2+}, green Fe^{2+}, brown Fe^{3+}

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> **Worked example 5.** An unknown gives a white precipitate that stays in excess sodium hydroxide, so it could be Ca^{2+} or Mg^{2+}. A flame test gives orange-red, so it is calcium. If there is no flame colour, it is magnesium.

The equations balance the hydroxide ions against the charge on the metal ion:

$$CuCl_2(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + 2NaCl(aq)$$

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

A 2+ ion needs two OH^- ions; a 3+ ion, such as Fe^{3+} or Al^{3+}, needs three.

Identifying negative ions (chemistry only)

Ion	Test	Result
carbonate CO_3^{2-}	add dilute acid; bubble the gas through limewater	**effervescence** 冒泡; the limewater turns milky
chloride Cl^-	add dilute nitric acid, then **silver nitrate** 硝酸银 solution	white precipitate
bromide Br^-	add dilute nitric acid, then silver nitrate solution	cream precipitate
iodide I^-	add dilute nitric acid, then silver nitrate solution	yellow precipitate
sulfate SO_4^{2-}	add dilute hydrochloric acid, then **barium chloride** 氯化钡 solution	white precipitate



The silver halide test: AgCl is white, AgBr cream and AgI yellow

The acid removes carbonate ions, which would otherwise also give a precipitate. Use nitric acid, not hydrochloric acid, in the halide test, because hydrochloric acid contains chloride ions and would always give a white precipitate. Use a fresh sample of the unknown for each test, so that the chemicals from one test do not affect the next.

> **Worked example 6.** A white solid gives a lilac flame, and its solution gives a cream precipitate with nitric acid and silver nitrate. It contains K^+ and Br^-: it is potassium bromide.

■ Required practical: identifying ions

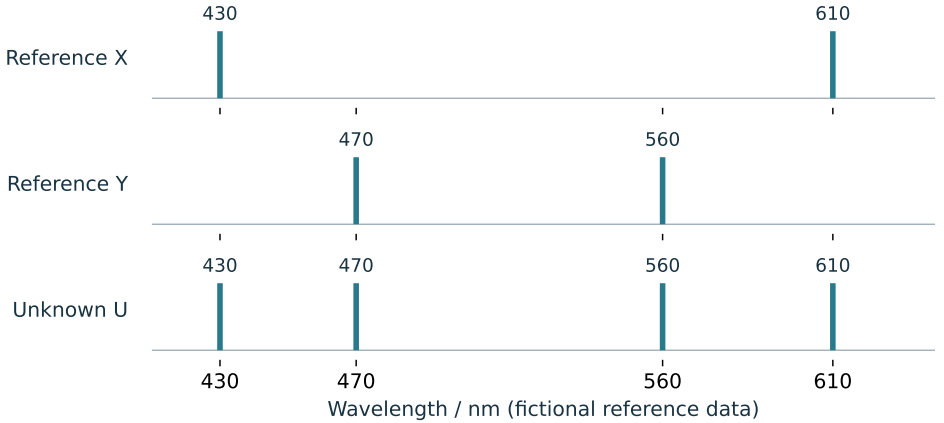
You test known and unknown ionic compounds with the flame test and the precipitate tests above. Use small amounts, clean apparatus and fresh samples. Record in three columns: the test, the observation, and what it shows.

Instrumental methods (chemistry only)

An **instrumental method** 仪器分析法 uses a machine to analyse a sample. Instruments are **accurate** 准确, giving results close to the true value; they have high **sensitivity** 灵敏度, detecting very small amounts; and they are fast. They must be set up with known samples first, a process called **calibration** 校准.

Flame emission spectroscopy. **Spectroscopy** 光谱法 studies light. In **flame emission spectroscopy** 火焰发射光谱法, the sample is put into a flame and its light passes through a **spectroscope** 分光镜, which splits it into a **line spectrum** 线状光谱: a set of bright lines, each at a particular **wavelength** 波长. Each metal ion has its own pattern of lines, like a barcode. Unlike a flame test, this method can identify several metals in a mixture, and the brightness of the lines shows the concentration. Doping tests on athletes and tests for lead in drinking water use instruments like this.

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Metal references X (430/610 nm) and Y (470/560 nm); unknown contains both.

> **Worked example 7.** Metal X gives lines at 430 and 610 nm; metal Y gives lines at 470 and 560 nm. An unknown shows all four lines, so it contains both X and Y. A missing line does not always mean a metal is absent; it may be below the **detection limit** 检测限 of the instrument.

> **Worked example 8: using a calibration.** In a calibration, a solution of 2.0 mg/dm^3 gives a reading of 20 and 6.0 mg/dm^3 gives 60, on a straight line. An unknown gives 35. Its concentration is $2.0 + \frac{35-20}{60-20} \times (6.0 - 2.0) = 2.0 + 1.5 = 3.5$ mg/dm^3.

A simple chemical test gives a **qualitative** 定性 answer: what is there. An instrument can also tell you how much is there, which matters when the question is whether water is safe to drink.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.8, and the required practicals handbook (Required practicals 6 and 7).
- J. C. Whorton, *The Arsenic Century: How Victorian Britain was Poisoned at Home, Work, and Play* (2010), on the Bradford sweets poisoning of 1858.
- *Pharmacy Act 1868* (31 & 32 Vict. c. 121), on the sale of poisons.
- G. Kirchhoff and R. Bunsen, "Chemische Analyse durch Spectralbeobachtungen", *Annalen der Physik* (1860), the origin of flame spectroscopy.

Common errors

- Saying a natural product with nothing added is "pure"; in chemistry, pure means one substance only.
- Drawing the baseline in ink, or putting it below the solvent level.
- Measuring to the edge of a spot, or from the bottom of the paper, when calculating Rf.
- Testing for oxygen with a burning splint; it must be glowing.
- Using hydrochloric acid in the halide test; it adds chloride ions.
- Confusing copper's green flame with its blue hydroxide precipitate, or the green Fe^{2+} precipitate with the brown Fe^{3+} one.
- Writing an observation when the question asks for a conclusion, or the other way round.

Key terms

- **Pure substance**: a single element or compound, not mixed with any other substance.
- **Formulation**: a mixture designed as a useful product, with measured amounts of each part.
- **Chromatography**: a method of separating substances that move at different speeds through a stationary phase.
- **Rf value**: the distance moved by a substance divided by the distance moved by the solvent.
- **Precipitate**: an insoluble solid that forms when two solutions react.
- **Flame emission spectroscopy**: an instrumental method that identifies metal ions from the lines in the light they give out in a flame.

Check yourself

1. Is sea water a pure substance? Explain.
2. A solid melts between 61 °C and 66 °C. Its pure form melts at 69 °C. What does this show?
3. A spot moves 2.4 cm and the solvent moves 8.0 cm. Calculate the Rf value.
4. Why must the baseline be above the level of the solvent?
5. Describe the test for carbon dioxide and its result.
6. (Chemistry only) A solution gives a brown precipitate with sodium hydroxide. Which ion is present?
7. (Chemistry only) Describe how to test a solution for sulfate ions.
8. (Chemistry only) Give two advantages of flame emission spectroscopy over a flame test.

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Answers: 1. No: it is a mixture of water and many dissolved salts. 2. It is impure: it melts below the true melting point and over a range. 3. $2.4/8.0 = 0.30$. 4. So that the spots do not dissolve into the solvent in the beaker. 5. Bubble the gas through limewater; it turns milky. 6. Iron(III), Fe^{3+}. 7. Add dilute hydrochloric acid, then barium chloride solution; a white precipitate shows sulfate. 8. For example: it can identify several metals in a mixture; it can measure concentration; it is more sensitive.

9 · Chemistry of the atmosphere

AQA · GCSE · Chemistry

Chemistry of the atmosphere: explain the evidence

Why it matters. In March 1958 a young American chemist, Charles David Keeling, began measuring carbon dioxide in the air near the top of Mauna Loa, a volcano in Hawaii, 3,400 metres up and far from any city. His first readings were about 313 parts per million (ppm). He soon noticed that the level rose and fell each year as the forests of the northern hemisphere grew in summer and lost their leaves in winter. But underneath this yearly breathing, the level kept climbing. The measurements have never stopped: in May 2024 they reached 427 ppm, higher than at any time in at least 800,000 years. This topic explains how Earth's atmosphere formed, how carbon dioxide affects the climate, and which pollutants come from burning fuels.

In this unit you will:

- give the composition of today's atmosphere;
- explain one theory of how the early atmosphere formed and changed, and why the evidence is limited;
- explain how oxygen increased and carbon dioxide decreased;
- explain the greenhouse effect, and judge evidence about climate change and its effects;
- calculate and compare carbon footprints;
- predict the products of burning fuels, and describe the effects of the pollutants and how to reduce them.

AQA GCSE Chemistry (8462) topic 4.9, also in Combined Science: Trilogy (8464), tested on Paper 2. The whole topic is the same for both tiers.

The atmosphere today

For about the last 200 million years, the atmosphere has been roughly four-fifths (about 80%) nitrogen and one-fifth (about 20%) oxygen. The rest is small amounts of other gases: argon and the other **noble gases** 稀有气体, carbon dioxide (about 0.04%), and a varying amount of **water vapour** 水蒸气.

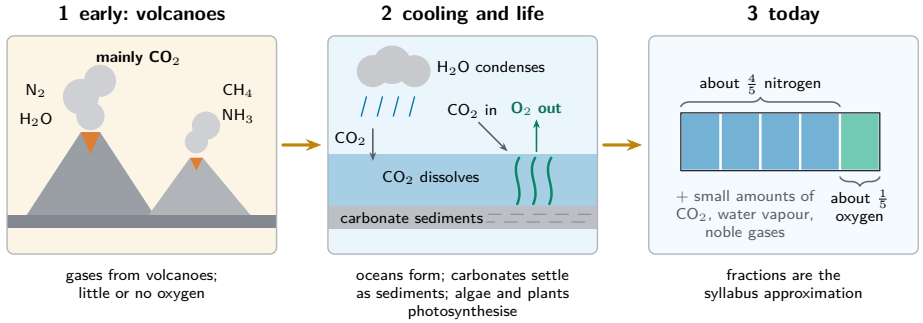
> **Worked example 1.** Estimate the volumes of nitrogen and oxygen in 250 cm^3 of air. Nitrogen $= \frac{4}{5} \times 250 = 200$ cm^3; oxygen $= \frac{1}{5} \times 250 = 50$ cm^3.

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The early atmosphere

Earth formed about 4.6 billion years ago. Nobody was there to measure the air, and rocks from that time have been changed or destroyed, so the evidence is limited. Scientists have theories, which change as new evidence is found.

One theory is that during the first billion years, intense **volcanic activity** 火山活动 released the gases that formed the early atmosphere. It was mainly carbon dioxide, with little or no oxygen, rather like the atmospheres of Mars and Venus today. Volcanoes also released nitrogen, which built up over time, water vapour, and small amounts of **methane** 甲烷 and ammonia.



Qualitative atmosphere sequence: volcanic carbon dioxide, then oceans, sediments and photosynthesis; only today's air carries the syllabus fractions.

As Earth cooled, the water vapour turned into liquid water by **condensation** 凝结 and formed the oceans. Carbon dioxide dissolved in the oceans, and carbonates precipitated out and settled on the sea floor as **sediment** 沉积物, which removed carbon dioxide from the air.

> **Worked example 2: judging the evidence.** Volcanoes today give off carbon dioxide, nitrogen and steam. This supports the volcanic theory, because it shows a possible source of those gases. But it does not tell us the exact composition of the air four billion years ago. A good answer gives both the support and the limit.

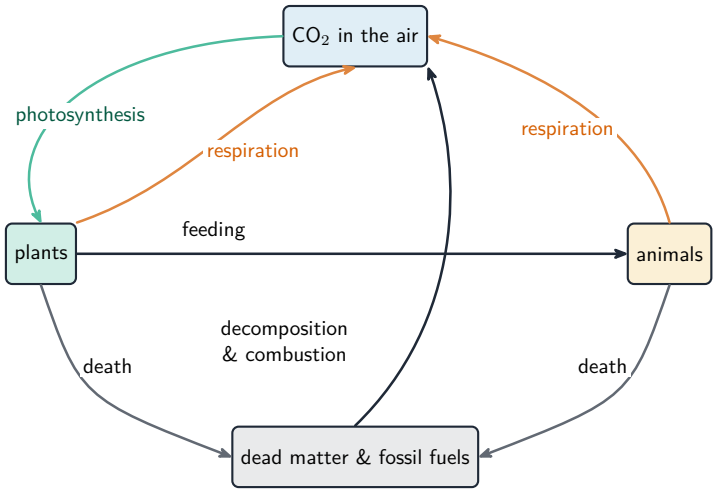
How oxygen increased and carbon dioxide decreased

Oxygen. About 2.7 billion years ago, **algae** 藻类 began to produce oxygen by **photosynthesis** 光合作用. Plants evolved later. Over the next billion years, the oxygen level slowly rose until animals could evolve.

$$6CO_2 + 6H_2O \xrightarrow{\text{light}} C_6H_{12}O_6 + 6O_2$$

Carbon dioxide. Several processes reduced the carbon dioxide in the air: it dissolved

in the oceans and formed carbonate sediments; algae and plants used it in photosynthesis to make **biomass** 生物量; and some of that carbon was buried and locked away in rocks for millions of years.



The carbon cycle: photosynthesis removes CO_2; respiration, combustion and decay return it

- **Limestone** 石灰岩 formed from the calcium carbonate shells and skeletons of sea creatures, which settled on the sea floor and became rock. The White Cliffs of Dover are made of chalk, a kind of limestone, formed from tiny shells in this way.
- **Coal** 煤 formed from plants that were buried in swamps before they could decay, then squeezed and heated over millions of years.
- **Crude oil** 原油 and **natural gas** 天然气 formed from the remains of tiny sea creatures, buried in mud and changed by heat and pressure.

Coal, crude oil and natural gas are each a **fossil fuel** 化石燃料. When we burn them, we return carbon to the air as carbon dioxide in a few hundred years that took millions of years to lock away.

Living things also release carbon dioxide by **respiration** 呼吸作用. The amount in the air depends on the balance between processes that remove it and processes that add it.

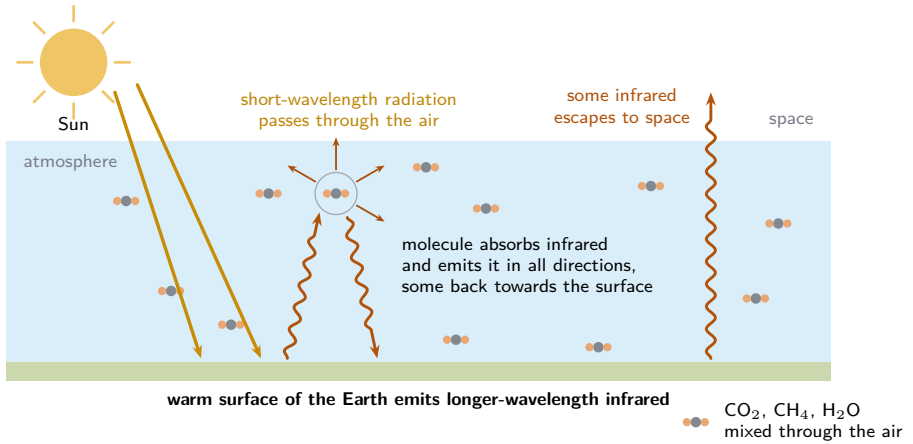
The greenhouse effect

Each **greenhouse gas** 温室气体, such as water vapour, carbon dioxide or methane, keeps the Earth warmer than it would otherwise be. Without any greenhouse effect, the average surface temperature would be about −18 °C instead of about 15 °C, too cold for most life.

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This is how the **greenhouse effect** 温室效应 works:

1. Short-wavelength radiation from the Sun passes through the atmosphere and warms the Earth's surface.
2. The warm surface gives out longer-wavelength **infrared radiation** 红外辐射.
3. Greenhouse gases absorb some of this infrared radiation and give it out again in all directions, including back towards the surface.
4. So less energy escapes to space, and the surface stays warmer.



Short-wavelength input and longer-wave infrared absorption/emission; gases mixed through air.

More greenhouse gases mean that more infrared is absorbed and sent back, so the Earth warms until it reaches a new balance between energy in and energy out. The gases are mixed through the air; they are not a solid layer, and they do not reflect radiation like a mirror.

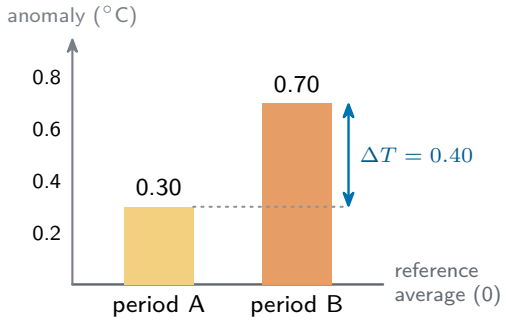
Human activities and climate evidence

Carbon dioxide is increased by burning fossil fuels and by **deforestation** 森林砍伐, which means fewer trees to remove carbon dioxide, and often burning or rotting wood. **Methane** is increased by cattle farming (cows release methane as they digest grass), by rice fields, and by organic waste rotting in **landfill** 垃圾填埋场 sites.

Judging evidence. Most scientists agree that human activities are increasing the average global temperature. Their evidence comes from many independent measurements: temperature records, ice cores, sea levels and the Keeling curve. Their work goes through **peer review** 同行评审: other experts check the methods and conclusions before a study is published. Peer review does not make a study perfect, but it

catches many mistakes. Climate is very complex, so models use simplifications. This means there is uncertainty about the exact size of future effects, but uncertainty about how much is not evidence that there is no effect. Be careful with headlines based on a single year or a single place.

> **Worked example 3.** A graph shows a **temperature anomaly** 温度距平 of 0.30 °C for period A and 0.70 °C for period B. An anomaly is the difference from a reference average, not the actual temperature. The change is $0.70 - 0.30 = 0.40$ °C: the world was warmer in period B. NASA's record puts the year 2023 about 1.2 °C above its 1951–1980 average.



fictional classroom data; differences from a reference, not absolute temperatures

Anomalies are differences from a reference average, so the 0.40 degree rise compares two period means, not absolute temperatures

Effects of climate change

Global climate change 全球气候变化 means long-term changes in patterns of weather, not the weather on one day. Possible effects include:

- rising sea levels, which can flood low-lying coasts and islands; sea level rises because water expands as it warms (**thermal expansion** 热膨胀) and because ice on land melts;
- more frequent and more severe storms, heatwaves and floods;
- changes in the amount and timing of rainfall, which can cause droughts or damage crops;
- changes to habitats, so that some species must move or die out, and pests and diseases can spread to new areas.

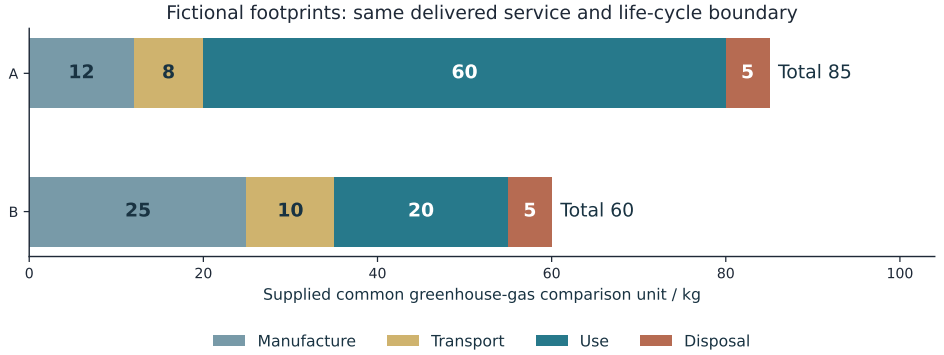
When you discuss an effect, think about its scale (how many people, how large an area, over what time) and its risk (how likely, and how serious). Bangladesh, where

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millions of people live less than one metre above sea level, faces a much larger risk from rising seas than a mountainous country.

Carbon footprints

A **carbon footprint** 碳足迹 is the total amount of carbon dioxide and other greenhouse gases emitted over the full **life cycle** 生命周期 of a product, service or event: getting the raw materials, making it, transporting it, using it and disposing of it. To compare two products fairly, they must give the same service, over the same time, with the same **system boundary** 系统边界, that is, the same stages counted.



Equal-service life-cycle totals: A85kg and B60kg in supplied common unit.

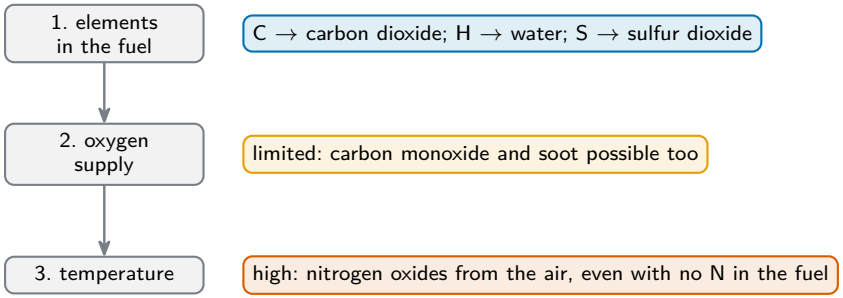
> **Worked example 4.** Two options give the same service. Option A emits 12 kg in manufacture, 8 kg in transport, 60 kg in use and 5 kg at disposal. Option B emits 25, 10, 20 and 5 kg.

- Total for A $= 12 + 8 + 60 + 5 = 85$ kg; total for B $= 25 + 10 + 20 + 5 = 60$ kg.
- B is lower overall, even though it costs more carbon to make.
- Percentage reduction $= (85 - 60)/85 \times 100 = 29.4\%$.

Reducing footprints. Carbon dioxide emissions can be cut by using renewable energy, more efficient machines, better **insulation** 隔热 for buildings, and public transport. Methane emissions can be cut by reducing food waste, collecting gas from landfill sites, and changing farming. There are limits: people may not change their habits, some countries cannot afford new technology, and governments may disagree about who should act first.

Pollutants from burning fuels

Most fuels contain carbon and hydrogen, and some contain sulfur. To predict the products of burning, check three things in order: what elements are in the fuel, how much oxygen there is, and how hot it burns.



Check in order: what the fuel contains, how much oxygen there is, and how hot it burns

- Complete combustion of a **pure hydrocarbon** 纯烃 in plenty of oxygen gives carbon dioxide and water: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
- **Incomplete combustion** 不完全燃烧, in too little oxygen, also gives **carbon monoxide** 一氧化碳 and **soot** 烟灰, solid carbon particles. **Particulates** 颗粒物 are small solid particles and unburnt hydrocarbons released into the air.
- Sulfur in the fuel burns to form sulfur dioxide: $S + O_2 \rightarrow SO_2$.
- In the high temperature of an engine, nitrogen and oxygen from the air react to form **nitrogen oxides** 氮氧化物 (NO_x). The nitrogen comes from the air, not the fuel.

Worked example 5. A town switches to sulfur-free diesel. This reduces sulfur dioxide, but it does not reduce nitrogen oxides, because their nitrogen comes from the air in the hot engine.

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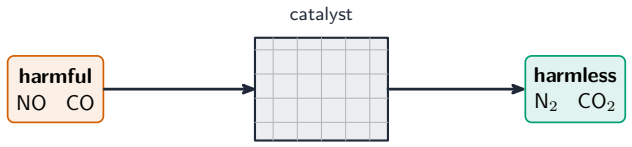
Pollutant	Effect	How to reduce it
carbon monoxide	**toxic** 有毒: it stops the blood carrying oxygen; it has no colour or smell	keep appliances serviced, provide good ventilation, fit carbon monoxide alarms
sulfur dioxide	breathing problems; **acid rain** 酸雨	remove sulfur from fuels; treat power-station gases with calcium oxide
nitrogen oxides	breathing problems; acid rain	catalytic converters in cars
particulates	breathing problems and heart disease; **global dimming** 全球变暗, less sunlight reaching the ground	particle filters; cleaner engines



Burning fossil fuels pollutes the air and changes the climate.

Acid rain damages trees, makes lakes too acidic for fish, and wears away limestone buildings and statues. In London in December 1952, smoke and sulfur dioxide from coal fires formed a thick smog that is thought to have killed about 4,000 people in a week, which led to the Clean Air Act of 1956.

A **catalytic converter** 催化转换器 in a car exhaust uses carbon monoxide to remove nitrogen monoxide: $2CO + 2NO \rightarrow 2CO_2 + N_2$. Two harmful gases become carbon dioxide and nitrogen.



$$2NO + 2CO \rightarrow N_2 + 2CO_2$$

A catalytic converter uses carbon monoxide to reduce nitrogen monoxide to nitrogen, forming carbon dioxide

Carbon monoxide, CO, and carbon dioxide, CO_2, are easy to confuse. CO is a poisonous gas from incomplete combustion; CO_2 is not poisonous at normal levels, but it is a greenhouse gas.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.9.
- C. D. Keeling, "The concentration and isotopic abundances of carbon dioxide in the atmosphere", *Tellus* (1960), and Scripps Institution of Oceanography / NOAA, Mauna Loa CO_2 record (May 2024 monthly mean).
- NASA Goddard Institute for Space Studies, *GISTEMP* global temperature analysis (2023 anomaly).
- Intergovernmental Panel on Climate Change, *Sixth Assessment Report: Summary for Policymakers* (2021).
- UK Met Office, "The Great Smog of 1952".

Common errors

- Saying the greenhouse effect is caused by the hole in the ozone layer; they are different problems.
- Saying greenhouse gases reflect heat like a mirror; they absorb and re-emit infrared radiation.
- Saying the greenhouse effect is entirely harmful; without it the Earth would be too cold for life.
- Confusing carbon monoxide (poisonous) with carbon dioxide (a greenhouse gas).
- Saying the nitrogen in nitrogen oxides comes from the fuel; it comes from the air.
- Counting only the use stage of a carbon footprint and forgetting manufacture, transport and disposal.
- Saying melting sea ice raises sea level; floating ice does not. Melting land ice and thermal expansion do.

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Key terms

- **Photosynthesis**: the process in which plants and algae use light to make glucose and oxygen from carbon dioxide and water.
- **Fossil fuel**: coal, crude oil or natural gas, formed from the remains of living things over millions of years.
- **Greenhouse gas**: a gas, such as carbon dioxide, methane or water vapour, that absorbs and emits infrared radiation.
- **Carbon footprint**: the total greenhouse gases emitted over the full life cycle of a product, service or event.
- **Incomplete combustion**: burning in too little oxygen, which can produce carbon monoxide and soot.
- **Acid rain**: rain made acidic by dissolved sulfur dioxide or nitrogen oxides.
- **Global dimming**: the reduction in sunlight reaching the Earth's surface caused by particulates.

Check yourself

1. Roughly what fraction of today's atmosphere is oxygen?
2. Why is the evidence about Earth's early atmosphere limited?
3. Give two ways in which carbon dioxide was removed from the early atmosphere.
4. Explain how carbon dioxide in the atmosphere keeps the Earth warm.
5. Give two human activities that increase methane in the atmosphere.
6. Option C has a footprint of 30 + 5 + 40 + 5 kg; option D has 20 + 5 + 70 + 5 kg. Which is lower, and by how much?
7. Why do car engines produce nitrogen oxides even when the fuel contains no nitrogen?
8. Why is carbon monoxide especially dangerous?

Answers: 1. About one-fifth (20%). 2. Nobody measured it, and rocks from that time have been changed or destroyed over 4.6 billion years. 3. For example: it dissolved in the oceans and formed carbonate sediments; algae and plants used it in photosynthesis. 4. The surface gives out infrared radiation; carbon dioxide absorbs some of it and re-emits it in all directions, including back to the surface, so less energy escapes to space. 5. Cattle farming; rotting waste in landfill sites (or rice fields). 6. C: $30 + 5 + 40 + 5 = 80$ kg; D: $20 + 5 + 70 + 5 = 100$ kg; C is lower by 20 kg. 7. The high temperature makes nitrogen and oxygen from the air react. 8. It is toxic, and it has no colour or smell, so people cannot tell it is there.

10 · Using resources

AQA · GCSE · Chemistry

Using resources: choose the route by the job it must do

Why it matters. Singapore has almost no rivers, lakes or ground water of its own, and for decades it bought much of its water from Malaysia. In February 2003 it opened its first NEWater plants. They take treated sewage, filter it through membranes so fine that they stop bacteria and viruses, push it through reverse-osmosis membranes that hold back dissolved salts, and finally shine ultraviolet light through it to kill anything left. The water that comes out is cleaner than the tap water in most cities. Today NEWater can supply up to 40% of Singapore's needs, mostly to factories that need very pure water. This topic is about getting what we need from Earth's resources, from safe water to metals, materials and fertilisers, and about counting the cost.

In this unit you will:

- tell finite and renewable resources apart, and explain sustainable development;
- explain how potable water is produced from fresh water, sea water and waste water;
- (higher tier) describe phytomining and bioleaching for extracting copper;
- carry out a life cycle assessment and compare reducing, reusing and recycling;
- (chemistry only) explain corrosion and its prevention, and choose alloys, ceramics, polymers and composites for a job;
- (chemistry only) describe the Haber process and the production of NPK fertilisers, and (higher tier) explain the choice of conditions.

AQA GCSE Chemistry (8462) topic 4.10, tested on Paper 2. Sections 4.10.1 and 4.10.2 are also in Combined Science: Trilogy (8464); sections 4.10.3 and 4.10.4 are chemistry only. It contains required practical 8 (analysis and purification of water samples).

Resources and sustainable development

We use Earth's resources for warmth, shelter, food and transport. Natural resources, supplemented by farming, give us food, timber, clothing and fuels. Chemistry also makes synthetic products, such as polyester fibres, that can replace natural ones.

A **finite resource** 有限资源 will run out, because it is not replaced as fast as we use it: metal ores and fossil fuels are examples. A **renewable resource** 可再生资源, such as timber from a managed forest, can be replaced in a reasonable time. But renewable does not mean unlimited: if trees are cut faster than they grow, the forest still shrinks.

Sustainable development 可持续发展 meets the needs of people today without stopping future generations from meeting their own needs. Chemists help by improving crop yields, making materials last longer, and designing processes that use less energy and produce less waste.

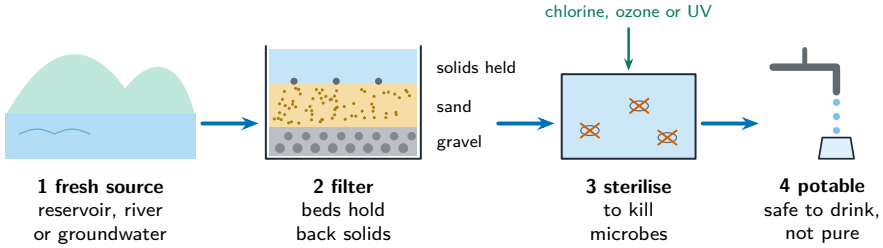
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Potable water

Potable water 饮用水 is water that is safe to drink. It is not pure in the chemical sense, because it contains small amounts of dissolved salts, but the levels of salts and microbes are low enough to be safe.

From fresh water. In the UK, rain collects in the ground, lakes and rivers. To make it potable:

1. choose a suitable source of fresh water;
2. pass it through filter beds to remove solid particles;
3. **sterilise** 消毒 it with chlorine, ozone or ultraviolet light to kill microbes.



salt water (sea water) is a different route: **distillation** or **reverse osmosis** removes the dissolved salts, but both need a lot of energy, so they are used only where fresh water is scarce.

Fresh-water filtration and sterilisation, distinct from desalination.

From sea water. Where fresh water is scarce, as in the Middle East, **desalination** 海水淡化 removes the salt from sea water. In distillation, the water is boiled and the steam condensed, leaving the salts behind. In **reverse osmosis** 反渗透, high pressure forces water through a membrane that holds back the salts. Both methods need large amounts of energy, so they are expensive.

■ Required practical: analysing and purifying water

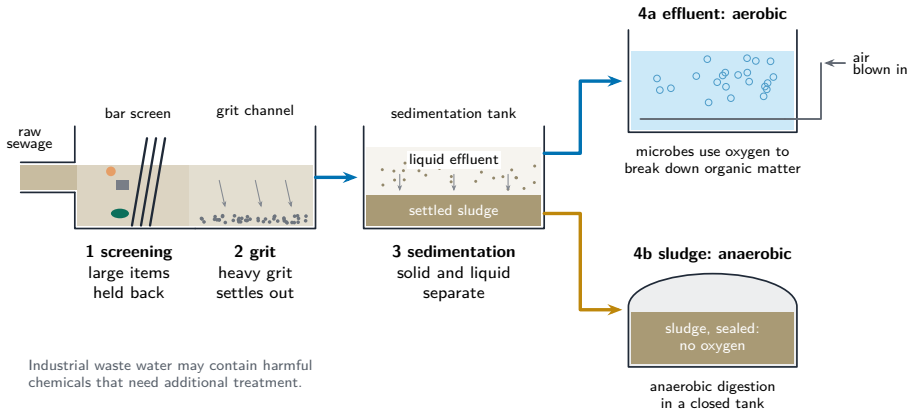
Test samples of water (for example tap water, rain water and sea water). Measure the pH with a pH meter or universal indicator. To find the mass of dissolved solids, weigh an empty evaporating basin, add a measured volume of water, heat it gently until all the water has evaporated, let it cool, and weigh again. Then purify salt water by distillation, collecting the condensed water in a cooled test tube. Never drink laboratory samples.

> **Worked example 1.** 50.0 cm^3 of a water sample leaves 0.120 g of solid. 50.0 cm^3 = 0.0500 dm^3, so the dissolved solids are 0.120/0.0500 = 2.40 g/dm^3.

Waste water. Sewage from homes and farms contains organic matter and harmful

microbes; waste water from factories may also contain harmful chemicals. Sewage treatment has these stages:

- screening removes large objects, and grit settles out;
- **sedimentation** 沉降 in large tanks separates it into **sludge** 污泥, which sinks, and **effluent** 出水, the liquid on top;
- the sludge goes through **anaerobic digestion** 厌氧消化: bacteria break it down without oxygen, producing methane gas that can be burned for energy;
- the effluent goes through **aerobic biological treatment** 好氧生物处理: air is bubbled through it so that other bacteria break down the remaining organic matter.



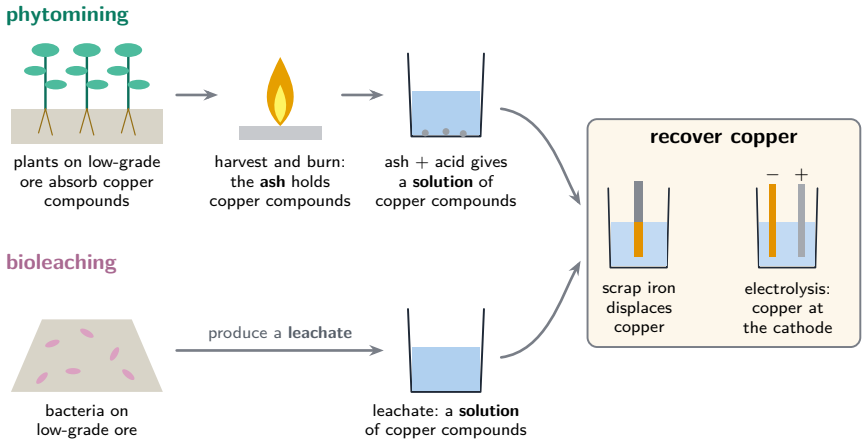
Sedimentation splits the sewage: liquid effluent is treated aerobically with air blown in, settled sludge anaerobically in a sealed digester.

Ground water is usually the easiest to make potable, because it has been filtered through rock. Waste water needs more treatment, and sea water needs the most energy.

Extracting copper with living things (higher tier)

High-grade copper ores are running out, so copper is now also extracted from low-grade ores and waste heaps, which contain only small amounts of copper.

- **Phytomining** 植物采矿 uses plants to absorb metal compounds from the soil. The plants are harvested and burned, and the ash contains the metal compounds.
- **Bioleaching** 生物浸出 uses bacteria to produce a solution, called a **leachate** 浸出液, that contains metal compounds.



HT biological extraction followed by copper recovery.

The copper is then obtained from the compounds by electrolysis or by displacement with scrap iron:

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

These methods avoid digging up, moving and dumping huge amounts of rock, but they are slow. Today about a fifth of the world's copper comes from bioleaching.

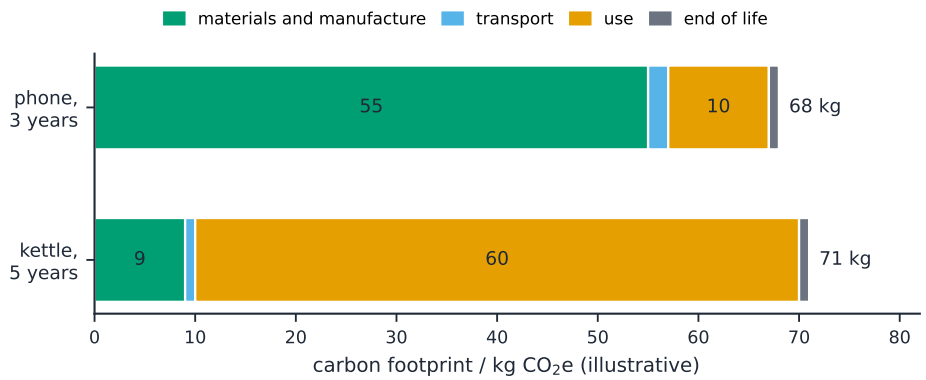
Life cycle assessment

A **life cycle assessment** 生命周期评估 (LCA) looks at the environmental impact of a product at every stage of its life:

1. extracting and processing the raw materials;
2. manufacturing and packaging;
3. use and operation during its lifetime;
4. disposal at the end of its life.

Transport between stages is included too.

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Where the footprint comes from depends on the product (round numbers for teaching)

The use of water, energy and resources, and the amount of some wastes, can be measured fairly easily. But giving values to pollution effects, such as how bad one gas is compared with another, needs judgements. So an LCA is not completely objective, and companies can use selective or shortened LCAs to support claims for advertising.

> **Worked example 2: paper or plastic?** Making, transporting and disposing of a paper bag uses 12 units of energy, and it is used once. A reusable plastic bag uses 30 units to make and 1 unit per use for cleaning. After n uses, the paper bags have used $12n$ units and the plastic bag $30+n$ units. These are equal when $12n = 30+n$, so $11n = 30$ and $n = 2.7$. From the third shopping trip onwards, the reusable bag uses less energy, as long as it really is reused.

Reduce, reuse and recycle

Using less material reduces the energy needed to extract and process it and the waste produced. Glass bottles can be reused after washing; or the glass can be crushed and melted to make new bottles. Metals can be melted and recycled; scrap steel is added to iron from the blast furnace, which reduces the amount of iron ore that must be mined. Recycling aluminium uses only about 5% of the energy needed to extract new aluminium from its ore.

Recycling still uses energy for collection, sorting, transport and melting. Mixed or dirty materials are harder to recycle.

> **Worked example 3.** Recycling a batch of metal uses 8 MJ; making the same amount from ore uses 50 MJ. The saving is $50 - 8 = 42$ MJ, or $42/50 \times 100 = 84\%$.

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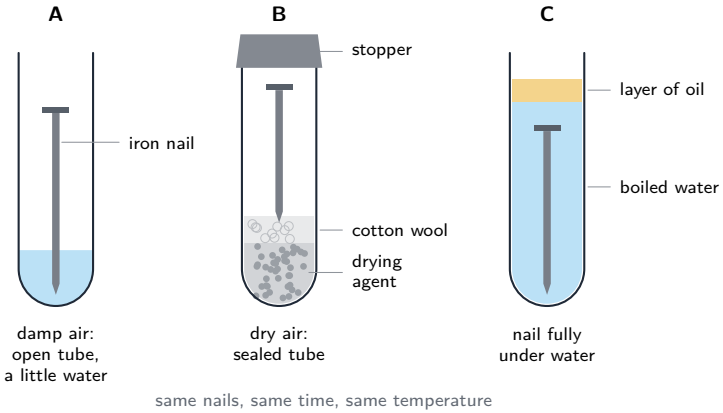
Corrosion and its prevention (chemistry only)

Corrosion 腐蚀 is the destruction of a material by reactions with substances in its environment. The corrosion of iron is called rusting, and it needs both oxygen and water.



Sea water and air have badly rusted this iron chain – the rust is hydrated iron(III) oxide

You can show this with three iron nails in test tubes: one in air and a little water, one in dry air with a drying agent, and one in boiled water (which has no dissolved air) under a layer of oil. Only the nail with both air and water rusts.



Water and oxygen controls for iron rusting: same nails, time and temperature in each tube.

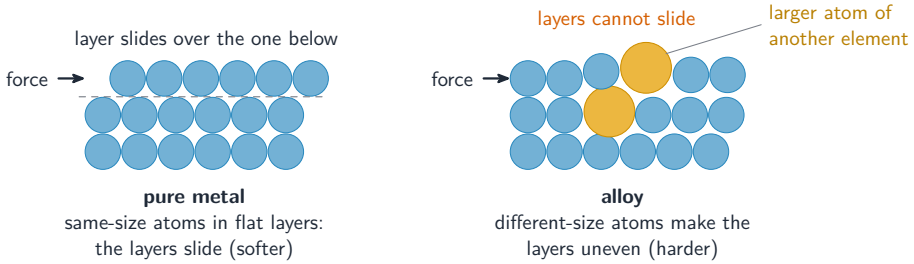
Preventing corrosion.

- A barrier stops air and water reaching the metal: paint, oil or grease, or electroplating with another metal.
- **Sacrificial protection** 牺牲保护 uses a more reactive metal. Galvanised iron is coated with zinc. Zinc reacts in place of the iron, so even if the coating is scratched, the iron does not rust. Ships' hulls have blocks of zinc or magnesium fixed to them for the same reason.

Aluminium is quite reactive, but it does not corrode badly, because it forms a thin layer of aluminium oxide that sticks to the surface and protects the metal underneath.

Alloys (chemistry only)

Most metals we use are alloys. In an **alloy** 合金, atoms of different sizes distort the layers of metal atoms, so they slide over each other less easily. This makes alloys harder than pure metals.



Different-sized atoms disrupt the layers, so they slide less easily: one reason many alloys are harder than the pure metal

Alloy	Made of	Used for
bronze	copper and tin	statues, medals, bells
brass	copper and zinc	door handles, musical instruments, water taps
gold alloys	gold with silver, copper and zinc	jewellery (harder than pure gold)
high carbon steel	iron with more carbon	cutting tools: strong but brittle
low carbon steel	iron with less carbon	car bodies: soft and easily shaped
stainless steel	iron with chromium and nickel	cutlery: hard and does not rust
aluminium alloys	aluminium with other metals	aircraft: low density

The purity of gold is measured in carats. 24 **carat** 开 gold is pure gold, so 18 carat gold is $18/24 = 75\%$ gold. **Worked example 4.** A 12 g ring of 18 carat gold contains $12 \times 18/24 = 9$ g of gold.

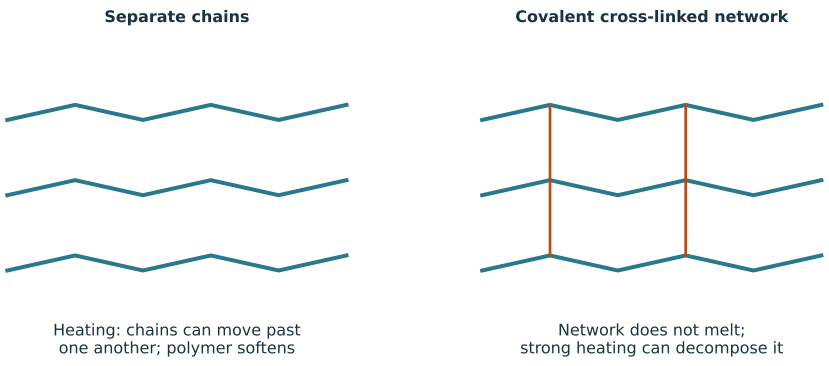
Ceramics, polymers and composites (chemistry only)

Ceramics. Most glass is soda-lime glass, made by heating sand, sodium carbonate and limestone. Borosilicate glass, made from sand and boron trioxide, melts at a higher temperature, so it is used for oven dishes and laboratory glassware. Clay ceramics, such as pottery and bricks, are made by shaping wet clay and heating it in a furnace.

Polymers. The properties of a polymer depend on its monomer and on the conditions used to make it. Low density poly(ethene) has branched chains that cannot pack closely, so it is soft and flexible, good for bags. High density poly(ethene), made with a catalyst at a lower temperature, has straighter chains that pack closely, so it is stronger, good for bottles and pipes.

A **thermosoftening polymer** 热塑性聚合物 is made of separate chains held by weak intermolecular forces. When heated, the forces are overcome and the polymer softens, so it can be remoulded and recycled. A **thermosetting polymer** 热固性聚合物 has strong covalent cross-links between the chains. It does not melt when heated, so it is used for things like saucepan handles and electrical plugs.

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Separate polymer chains and a covalent cross-linked network.

Composites. A **composite** 复合材料 is made of two materials: a **reinforcement** 增强体, such as fibres or fragments, held together by a **matrix** 基体, or binder. Concrete is gravel and sand in cement; fibreglass is glass fibres in a polymer resin; carbon fibre composite is carbon fibres in a resin and is strong and very light, so it is used in racing bikes and aircraft.

The Haber process (chemistry only)

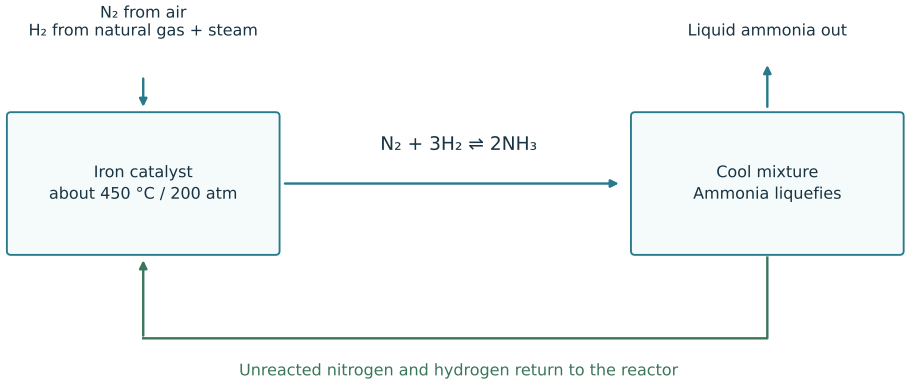
The **Haber process** 哈伯法 makes ammonia, which is used to produce nitrogen-based fertilisers. Fritz Haber first demonstrated it in 1909, and Carl Bosch scaled it up for industry. Without it, the world could not grow enough food for its present population.



Ammonia for fertiliser is made by the Haber process in large industrial plants like this one.

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- Nitrogen comes from the air; hydrogen is made from natural gas.
- The purified gases are passed over an iron catalyst at about 450 °C and 200 atmospheres pressure.
- The reaction is reversible: $N_2 + 3H_2 \rightleftharpoons 2NH_3$.
- The gas mixture is cooled, so the ammonia liquefies and is removed. The unreacted nitrogen and hydrogen are recycled back to the reactor.



Haber conditions, ammonia removal and a connected recycle loop.

Higher tier: choosing the conditions. The forward reaction is exothermic.

- A low temperature would give a higher yield, but the reaction would be too slow. 450 °C is a compromise between a reasonable yield and a fast rate.
- A high pressure gives a higher yield, because there are 4 molecules of gas on the left and 2 on the right. It also increases the rate. But very high pressures need expensive, thick-walled equipment and a lot of energy, so 200 atmospheres is a compromise with cost and safety.
- The iron catalyst speeds up the reaction but does not change the yield.

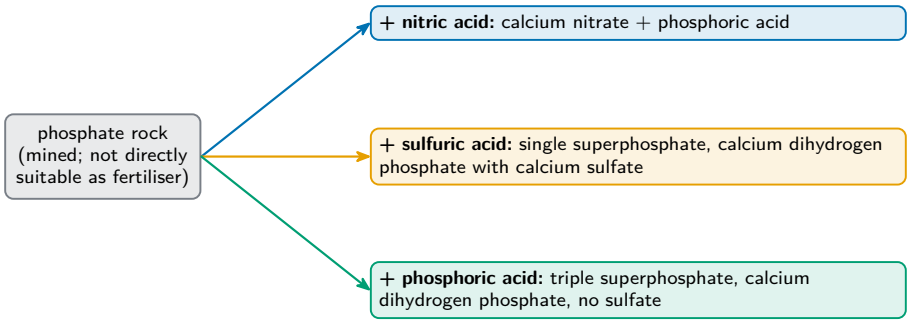
> **Worked example 5.** 2 mol of nitrogen and 6 mol of hydrogen could, in theory, make 4 mol of ammonia. If only 1 mol of nitrogen reacts as the gases pass through the reactor once, it uses 3 mol of hydrogen and makes 2 mol of ammonia. The remaining 1 mol of nitrogen and 3 mol of hydrogen are recycled, so in the end almost all of them are converted.

NPK fertilisers (chemistry only)

An **NPK fertiliser** 氮磷钾化肥 is a formulation that supplies the three elements plants need most: nitrogen (N), phosphorus (P) and potassium (K).

- Nitrogen: ammonia from the Haber process is used to make nitric acid and ammonium salts. Ammonia reacts with nitric acid to make ammonium nitrate: $NH_3 + HNO_3 \rightarrow NH_4NO_3$.
- Potassium: potassium chloride and potassium sulfate are mined.
- Phosphorus: phosphate rock is mined, but it cannot be used directly because it does not dissolve. It is treated with acid to make soluble salts.

Phosphate rock treated with	Products
nitric acid	phosphoric acid and calcium nitrate
sulfuric acid	single superphosphate (a mixture of calcium phosphate and calcium sulfate)
phosphoric acid	triple superphosphate (calcium phosphate)



Three acid treatments of phosphate rock and what each gives

In industry, these salts are made in huge continuous plants from raw materials. In a school laboratory, a small batch of ammonium sulfate can be made by titrating ammonia solution with dilute sulfuric acid and crystallising the salt.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.10, and the required practicals handbook (Required practical 8).
- PUB, Singapore's National Water Agency, "NEWater" (first plants opened February 2003; up to 40% of current demand).
- F. Haber, Nobel Lecture, "The synthesis of ammonia from its elements" (2 June 1920).
- V. Smil, *Enriching the Earth: Fritz Haber, Carl Bosch, and the Transformation of World Food Production* (2001).
- International Aluminium Institute, on the energy saving of recycling aluminium.

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Common errors

- Saying potable water is pure water; it contains dissolved substances at safe levels.
- Saying filtration removes dissolved salts or kills microbes; it removes solid particles only.
- Saying iron rusts in water alone or in air alone; it needs both.
- Saying a scratched painted surface still protects iron; only sacrificial protection keeps working at a scratch.
- Saying thermosoftening polymers have no forces between chains; they have weak intermolecular forces, which are overcome on heating.
- Saying the Haber process uses the conditions for maximum yield; the conditions are a compromise with rate and cost.
- Saying phytomining produces pure copper directly; the ash must still be processed.

Key terms

- **Finite resource**: a resource that will run out because it is not replaced as fast as it is used.
- **Sustainable development**: meeting the needs of the present without harming the ability of future generations to meet theirs.
- **Potable water**: water that is safe to drink.
- **Life cycle assessment**: an assessment of the environmental impact of a product at each stage of its life.
- **Corrosion**: the destruction of a material by chemical reactions with its environment.
- **Sacrificial protection**: protecting a metal by attaching a more reactive metal that corrodes instead.
- **Composite**: a material made of a reinforcement held in a matrix.
- **Haber process**: the industrial manufacture of ammonia from nitrogen and hydrogen.

Check yourself

1. Is wood a finite or a renewable resource? Explain.
2. Why is sea water not made potable in the UK, even though there is plenty of it?
3. What is the difference between sludge and effluent in sewage treatment, and how is each treated?
4. Give one reason why a life cycle assessment is not completely objective.
5. (Chemistry only) Why does galvanised iron not rust even when the zinc coating is scratched?

6. (Chemistry only) Why is 9 carat gold harder than 24 carat gold? What percentage of 9 carat gold is gold?
7. (Chemistry only) Why can a thermosoftening polymer be recycled by melting, but a thermosetting polymer cannot?
8. (Chemistry only, HT) Why is a temperature of 450 °C used in the Haber process, rather than a lower temperature?

Answers: 1. Renewable: new trees can be grown in a reasonable time, as long as they are replanted. 2. Desalination needs a lot of energy, so it is expensive, and the UK has enough fresh water. 3. Sludge is the solid that sinks; it is broken down by anaerobic digestion. Effluent is the liquid; it is treated aerobically with air. 4. Giving values to pollution effects needs judgements (or: companies can choose which stages to include). 5. Zinc is more reactive than iron, so it reacts instead of the iron. 6. It is an alloy, and the different-sized atoms stop the layers sliding; $9/24 = 37.5\%$. 7. Thermosoftening polymers have only weak forces between chains, which are overcome on heating; thermosetting polymers have covalent cross-links, so they do not melt. 8. A lower temperature would give a higher yield but the reaction would be too slow; 450 °C is a compromise.

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