Atomic structure and the periodic table

AQA · GCSE · Chemistry

Atomic structure and the periodic table: evidence and predictions

A filter can separate sand from salt solution, but cannot separate sodium from chlorine in a salt crystal. The difference is chemical combination. This topic connects the particles in a substance to separation, atomic models and patterns of reactions.

Use a supplied periodic table to identify elements. Count atoms before balancing an equation; count protons before identifying an element. Higher Tier ionic/half-equations and Chemistry-only transition-metal content are labelled below.

Elements, compounds and equations (4.1.1.1)

An **atom** is the smallest part of an **element** that can exist. An element contains atoms with one proton number. It may contain single atoms, such as helium, or molecules, such as O_2. O_2 contains two atoms but only one element.

A **compound** contains two or more elements chemically combined in fixed proportions. H_2O always has two hydrogen atoms for each oxygen atom. Compounds have properties different from their elements and can be separated into elements only by chemical reactions. A **chemical reaction** makes one or more new substances and often has a detectable energy change.

Learn names and correctly capitalised symbols: Co means cobalt, while CO is carbon monoxide. Use the first twenty elements, Groups 1 and 7 and other named elements in this course. The first twenty are:

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

Formulae give composition: $MgCl_2$ is magnesium chloride; Na_2O is sodium oxide. A **word equation** names reactants and products: magnesium + oxygen → magnesium oxide. A **symbol equation** uses their formulae.

Worked method – balancing: start with $Mg + O_2 \rightarrow MgO$. There are two oxygen atoms on the left and one on the right. Put 2 before MgO, then 2 before Mg: $\mathbf{2Mg + O_2 \rightarrow 2MgO}$. Check Mg: 2 = 2; O: 2 = 2. Change whole-particle coefficients, not the subscripts in a substance's formula. MgO_2 would be a different formula. Three H_2O molecules contain six H atoms and three O atoms.

Higher Tier – charge as well as atoms: zinc forms Zn^{2+} by losing two electrons: $Zn \rightarrow Zn^{2+} + 2e^-$. Both sides have charge zero overall. A copper ion gains two: $Cu^{2+} + 2e^- \rightarrow Cu$. Adding the half-equations cancels electrons: $\mathbf{Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu}$. Sulfate ions present unchanged in solution are spectator ions and are omitted. Check atoms and total charge after cancellation; do not cancel species that change.

Choosing a separation (4.1.1.2)

A **mixture** contains substances not chemically combined. Each retains its chemical properties. Physical separation makes no new substance; choose a method using the components' properties and the product you need.

Method	Property used and example
filtration	Insoluble particles cannot pass through filter paper; sand stays as residue while water and dissolved salt form filtrate.
crystallisation	Dissolved solid forms crystals when a concentrated solution cools; recover salt from solution.
simple distillation	A liquid evaporates and then condenses; recover water from salt solution. Non-volatile salt remains in the flask.
fractional distillation	Miscible liquids have different boiling points; repeated evaporation/condensation in a column improves separation of ethanol and water.
chromatography	Dissolved substances move differently between solvent and stationary material; separate colours in ink.

Worked method – pure, dry salt from sand, salt and water: stir to dissolve any remaining salt. Filter; sand remains on the paper and salt solution passes through. Heat the filtrate gently to evaporate some water, then cool to form crystals. Filter the

crystals and dry them between filter papers. Washing the sand with a small amount of water can recover salt solution trapped with it. Filtering alone cannot collect dissolved salt.

Distillation recovers a liquid by evaporation followed by condensation.

Worked choice – recover the water instead: use simple distillation. Heat the salt solution; water vapour leaves the flask, cools in a condenser and collects as liquid water. Evaporation without collecting the vapour loses the desired product. Choose fractional distillation for two miscible liquids; a filter cannot separate them. For chromatography, keep the pencil start line above the solvent so the sample moves up the stationary material instead of dissolving directly into the solvent reservoir.

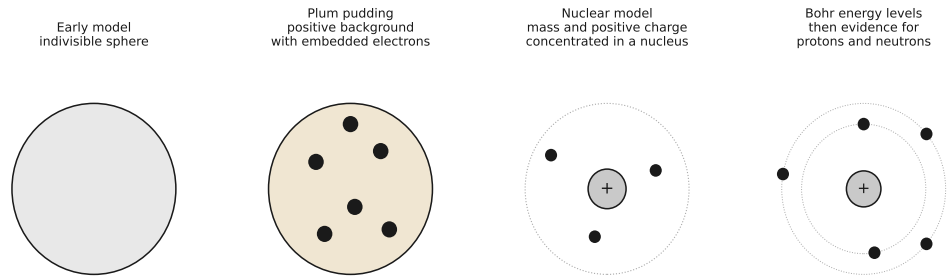
How evidence changed the atomic model (4.1.1.3)

A scientific **model** is a representation that explains evidence. New observations can change or replace it. Early atoms were tiny indivisible spheres. Discovery of the negative electron required internal structure, giving the **plum pudding model** : electrons embedded in a spread-out ball of positive charge.

In alpha scattering, most positive alpha particles passed through a thin metal foil, some were deflected and very few came back. These observations support different conclusions:

Observation	Nuclear-model conclusion
Most passed straight through	Most of the atom is empty space.
Some were deflected	Positive charge is concentrated and repels the positive alpha particles.
Very few returned towards the source	A tiny, dense central region contains most of the mass.

The **nuclear model** replaced spread-out positive charge with a tiny positive nucleus and surrounding electrons. Bohr then suggested electrons occupy specific distances/energy levels, consistent with observations. Later work identified protons as units of nuclear positive charge. Chadwick's evidence established neutrons in the nucleus about twenty years after the nuclear model became accepted. Detailed Bohr/Chadwick experimental apparatus is not required.



Worked explanation: "Some particles bounced back, so the whole atom is solid" contradicts the observation that most passed through. The combined evidence requires a small dense nucleus inside a mostly empty atom. A model must account for the full pattern, not one selected observation.

Counting particles, size and isotopes (4.1.1.4–4.1.1.6)

Particle	Relative mass	Relative charge	Location
proton	1	+1	nucleus
neutron	1	0	nucleus
electron	very small	−1	shells/energy levels

The **atomic number** Z is the proton number and identifies the element. All atoms of that element have the same Z. The **mass number** A is protons + neutrons. A neutral atom has equal proton and electron numbers, so its charges cancel. An ion forms by changing electrons; changing proton number would change the element.

Worked particle count: a chlorine-37 atom has Z = 17. Protons = 17; neutrons = 37 − 17 = 20; electrons = 17. Cl^- has gained one electron, giving 18 electrons while protons/neutrons stay 17/20. For calcium-40, Z = 20: Ca^{2+} has 20 protons, 20 neutrons and 18 electrons.

Isotopes are atoms of the same element with different neutron numbers. Chlorine-35 and chlorine-37 both have 17 protons, but 18 and 20 neutrons respectively. They remain chlorine; different mass number does not mean different element.

Typical atom radius is about $\mathbf{0.1\ nm = 1 \times 10^{-10}\ m}$; 1 nm = 10^{-9} m. Nucleus radius is less than one ten-thousandth of the atom's radius, of order 10^{-14} m. Almost all mass is in the nucleus. These are approximate scales, not exact dimensions of every atom.

Worked scale: if an atom's radius were enlarged to 1 m, a nucleus one ten-thousandth as large would have radius 1/10 000 m = 0.0001 m = 0.1 mm. The actual specified nucleus is smaller than that ratio. Compare radius with radius; an atom's diameter is twice its radius. Atoms of diameter about 2×10^{-10} m would number about 5×10^6

along a 1 mm line, ignoring gaps: $(1 \times 10^{-3})/(2 \times 10^{-10})$.

Relative atomic mass A_r is an abundance-weighted mean, with no unit. It need not be a whole number. Use the percentage abundances, not an unweighted average:

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

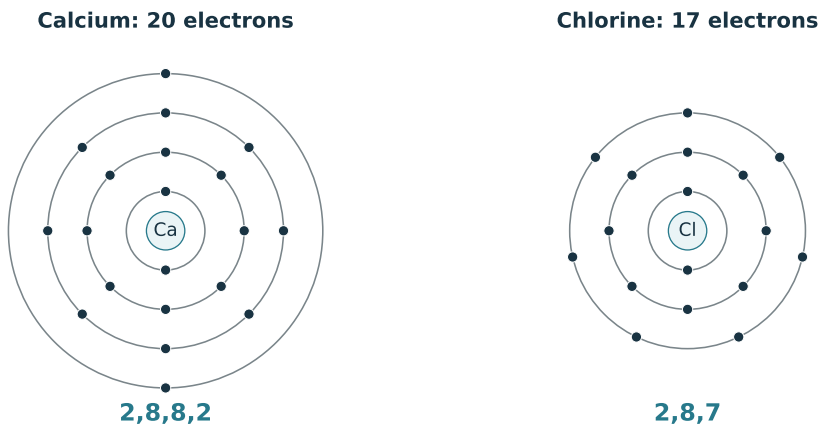
Worked mean: 75% chlorine-35 and 25% chlorine-37 give $(35 \times 75 + 37 \times 25)/100 = (2625 + 925)/100 =$ **35.5**. This is closer to 35 because that isotope is more abundant; it lies between 35 and 37. No single atom has "35.5 neutrons" or a mass number of 35.5.

Electron arrangements and periodic position (4.1.1.7–4.1.2.1)

An **outer shell electron** is in the highest occupied shell of an atom and helps determine its reactions.

Electrons occupy the lowest available **energy levels** first. For the first twenty neutral atoms, use the GCSE shell model: first shell up to 2, second up to 8, then 8 before the fourth begins. This is the required limited model, not a universal rule for all elements.

Worked arrangement: calcium has Z = 20 and therefore 20 electrons. Place 2 in the first shell, 8 in the second, 8 in the third and 2 in the fourth: **2,8,8,2**. A diagram must show those four shells and the same total number of electrons. Chlorine is **2,8,7**, with three occupied shells and seven outer electrons.



The modern **periodic table** orders elements by proton number. A **period** is a row; the number of occupied shells gives the period. A **group** is a column containing elements with related chemical properties. For main Groups 1–7, the group number gives the outer-electron number. Group 0 has a full outer shell: helium has two electrons, while the other noble gases in this model have eight. This exception prevents treating every group member as having identical electron counts.

Worked prediction: an unknown first-twenty element has 2,8,1. It has 11 electrons/protons, so it is sodium, in Period 3 and Group 1. Its single outer electron helps predict formation of a +1 ion and reactions similar to lithium/potassium. An element with 2,8,7 instead tends to gain one electron in ionic reactions. Position supports predictions; it does not replace reaction evidence.

Developing the periodic table; metals and non-metals (4.1.2.2–4.1.2.3)

Early tables used increasing atomic weight before subatomic particles were known. They were incomplete and strict mass ordering put some elements with unlike neighbours. **Mendeleev** left gaps for undiscovered elements and sometimes changed mass order to keep similar properties together. He predicted properties of missing elements; later discoveries matching those predictions supported his arrangement. Isotope knowledge explained why average atomic-weight order did not always match element identity. Modern ordering uses proton number.

Worked evidence judgement: finding an element with the predicted properties in a gap supports the table more strongly than merely finding any new element. A failed prediction would prompt review of the arrangement. Leaving gaps was a testable proposal, not proof by itself.

Metals, mainly towards the left/bottom, form positive ions in reactions by losing electrons. Non-metals, mainly towards the right/top, do not characteristically form positive ions; they may gain electrons or share them in covalent bonds. Most elements are metals. Metals typically conduct heat/electricity, are shiny and can be bent or shaped; solid non-metals are often brittle and poor conductors. These are characteristic patterns with exceptions, such as conducting graphite, not definitions without exceptions.

Group 0, Group 1 and Group 7 (4.1.2.4–4.1.2.6)

Noble gases are unreactive and do not easily form molecules because their atoms have stable full outer shells. They are monatomic. Their boiling points rise down Group 0 as relative atomic mass increases. Use supplied trend data to predict an intermediate value or whether a later member boils higher; do not claim exact values without data.

Alkali metals in Group 1 have one outer electron and form +1 ions. Lithium, sodium and potassium react with oxygen to form oxides, with chlorine to form chlorides, and with water to form a metal hydroxide and hydrogen. In the school model, $4Li + O_2 \rightarrow 2Li_2O$; $2Na + Cl_2 \rightarrow 2NaCl$; $\mathbf{2K + 2H_2O \rightarrow 2KOH + H_2}$. Do not assume every oxygen product under every condition has the same formula; use supplied

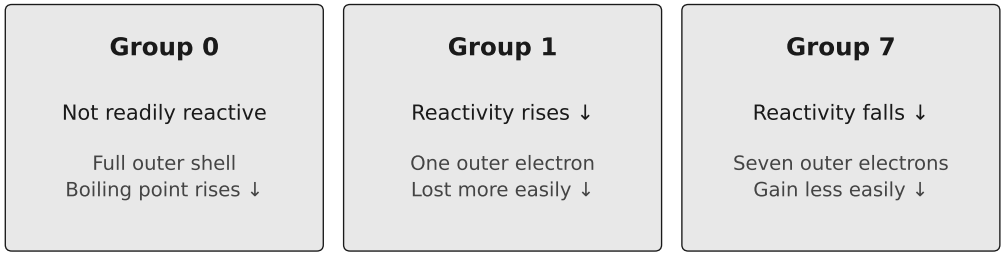
formulas when balancing.

With water, lithium fizzes and moves slowly; sodium reacts faster and often melts into a moving ball; potassium reacts still more vigorously and may ignite with a lilac flame. These are observations of supervised demonstrations. Reactivity rises down the group: more occupied shells place the outer electron farther from the nucleus with more shielding, so it is lost more easily despite greater nuclear charge.

Halogens in Group 7 have seven outer electrons and form diatomic molecules: Cl_2, Br_2 and I_2. At room temperature chlorine is a pale-green gas, bromine a red-brown liquid and iodine a grey-black solid. Relative molecular mass, melting point and boiling point rise down the group. Reactivity falls: the outer region is farther from the nucleus and more shielded, so attraction for an incoming electron is weaker.

Halogens form ionic halides with metals, such as NaCl, and covalent compounds with non-metals, such as HCl. Do not say all halogen compounds contain ionic bonds. **Displacement** occurs when a more reactive halogen replaces a less reactive one in an aqueous halide solution.

Worked displacement: chlorine added to potassium bromide produces bromine, giving an orange/brown solution: $\mathbf{Cl_2 + 2KBr \rightarrow 2KCl + Br_2}$. Chlorine gains electrons from bromide ions. Bromine does not displace chloride ions because it is less reactive. A "no reaction" result can therefore help establish a reactivity order. For Higher Tier: $Cl_2 + 2e^- \rightarrow 2Cl^-$; $2Br^- \rightarrow Br_2 + 2e^-$; net $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$. Atoms and charge are conserved.



Down a group, the outer occupied shell is farther from the nucleus.

Transition metals – Chemistry only (4.1.3.1–4.1.3.2)

Use chromium, manganese, iron, cobalt, nickel and copper as examples of **transition metals**. Compared with Group 1 metals, they typically have higher melting points and densities, greater strength/hardness, and lower reactivity with oxygen, water and halogens. Copper does not react with cold water as sodium does. Compare named examples and conditions rather than assuming all metals behave identically.

Many transition elements form ions with different charges and coloured compounds, and act as **catalysts**. Iron forms Fe^{2+} and Fe^{3+}; aqueous iron(II) salts can be pale green and iron(III) salts yellow/brown. Aqueous copper(II) salts are often blue. Colour depends on the compound/conditions: not every compound of an element has one fixed colour. Iron is the Haber-process catalyst; nickel catalyses hydrogenation. These examples illustrate typical properties, not properties of every possible transition compound.

Worked formula link: sulfate is SO_4^{2-}. Fe^{2+} gives $FeSO_4$; Fe^{3+} requires two iron ions and three sulfate ions, giving $Fe_2(SO_4)_3$. The charges balance to zero in both compounds. Different proportions/iron charges mean different compounds despite the same three element names.

Check yourself

- Choose a separation from component properties and the product required; explain what happens at each stage.
- Link alpha-scattering observations to the nuclear model, rather than memorising a timeline alone.
- Count particles in atoms and ions; calculate a weighted mean; distinguish radius, diameter and scale.
- Represent first-twenty electron arrangements as both numbers and shell diagrams, then predict periodic position.
- Use evidence to explain group trends and displacement; compare named transition-metal properties with Group 1.