

Atomic structure and the periodic table

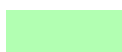
Specification 8462 · AQA

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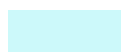
Revision notes for one topic, free to download.

These are the summary notes only. The full edition of this booklet also carries the worked examples, the required practicals and the formula sheets, with every topic in one place.

Colour Codes



Definitions



Diagrams



Notes and warnings



Sub-topic headings

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Define Atom

- the smallest part of an element that can exist.

Define Compound

- two or more elements chemically combined in fixed proportions.

Define Atomic number

- the number of protons in an atom.

Define Isotopes

- atoms of the same element with the same number of protons but different numbers of neutrons.

Define Electronic structure

- the arrangement of electrons in the energy levels (shells) of an atom, eg sodium is 2,8,1.

Define Metal

- an element that reacts to form positive ions.

Define Noble gases (Group 0)

- unreactive non-metals with stable arrangements of electrons.

Define Halogens (Group 7)

- non-metals with seven outer electrons, existing as molecules made of pairs of atoms.

Define Transition elements

- metals between Groups 2 and 3, with properties different from those of Group 1.

Define Element

- a substance made of only one type of atom; there are about 100.

Define Mixture

- two or more elements or compounds not chemically combined; each keeps its own chemical properties.

Define Mass number

- the sum of the protons and neutrons in an atom.

Define Relative atomic mass (A_r)

- an average value that takes account of the abundance of the isotopes of an element.

Define Group

- a column of the periodic table; its elements have the same number of outer electrons and so similar chemical properties.

Define Non-metal

- an element that does not form positive ions.

Define Alkali metals (Group 1)

- metals with a single electron in their outer shell.

Define Displacement reaction

- a reaction in which a more reactive halogen takes the place of a less reactive one in an aqueous solution of its salt.

1.1 A simple model of the atom, symbols, relative atomic mass, electronic charge and isotopes

HIGHER TIER

1.1.1 Atoms, elements and compounds

- All substances are made of **atoms**; an atom is the smallest part of an element that can exist. Atoms of each element are represented by a **chemical symbol**, eg O for oxygen, Na for sodium.
- There are about **100 different elements**, shown in the periodic table.
- **Compounds** are formed from elements by chemical reactions. Chemical reactions always involve the formation of one or more new substances, and often involve a detectable energy change.
- Compounds contain two or more elements chemically combined in **fixed proportions** and are represented by **formulae**. Compounds can only be separated into elements by chemical reactions.
- Reactions can be represented by **word equations** or by equations using symbols and formulae.

You are supplied with a periodic table in the exam and must be able to:

- use the names and symbols of the **first 20 elements**, the elements in **Groups 1 and 7**, and other elements in this specification
- **name compounds** of these elements from given formulae or symbol equations
- write word equations for the reactions in this specification
- write formulae and balanced chemical equations for those reactions
- **HT ONLY** write balanced **half equations** and **ionic equations** where appropriate.

Naming. Metal + non-metal takes the metal first and the non-metal ending becomes *-ide* (sodium chloride); if oxygen is also in the negative ion the ending is *-ate* (copper sulfate).

Balancing. Atoms are not created or destroyed, so each type must be equal on both sides. Change only the big numbers in front of a formula — never a small subscript.

1.1.2 Mixtures

A **mixture** consists of two or more elements or compounds not chemically combined together; the chemical properties of each substance are unchanged. Mixtures can be separated by **physical processes**, which do not involve chemical reactions and make no new substances.

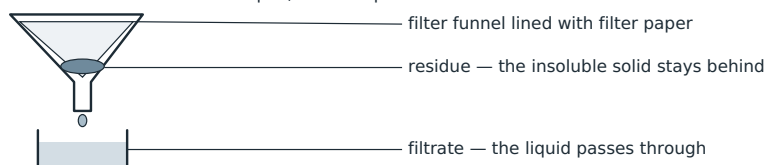
Technique	Separates	How it works
Filtration	An insoluble solid from a liquid	The liquid passes through the filter paper; the larger solid particles stay behind as residue
Crystallisation	A soluble solid from its solution	Solvent evaporates, the solution saturates and crystals of the solute form
Simple distillation	A solvent from a solution	The solvent boils, and the vapour is cooled and condensed; the dissolved solid stays behind
Fractional distillation	Miscible liquids with different boiling points	Vapours rise up a fractionating column; the lowest-boiling liquid reaches the top first
Chromatography	Substances with different solubilities in the solvent	The solvent carries substances up the paper at different rates, separating them

You must be able to describe, explain and give examples of each process, and suggest suitable separation and purification techniques for a mixture when given information about it.

Filtration, crystallisation, simple and fractional distillation, and paper chromatography

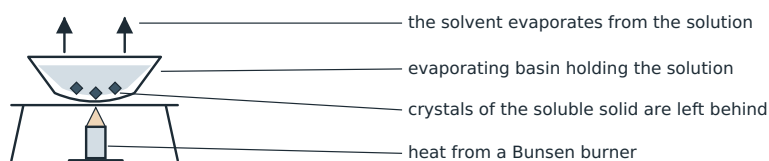
Filtration

Separates an insoluble solid from a liquid, for example sand from water.



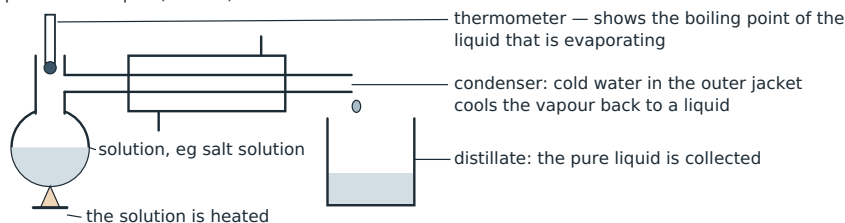
Crystallisation

Separates a soluble solid from its solution, for example salt from salt solution.



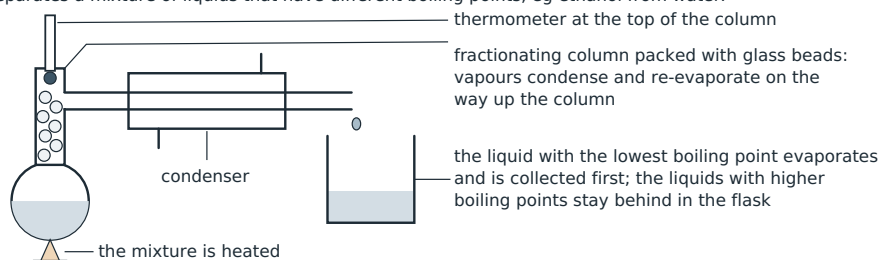
Simple distillation

Separates the liquid (solvent) from a solution — the dissolved solid is left in the flask.



Fractional distillation

Separates a mixture of liquids that have different boiling points, eg ethanol from water.



WATCH OUT

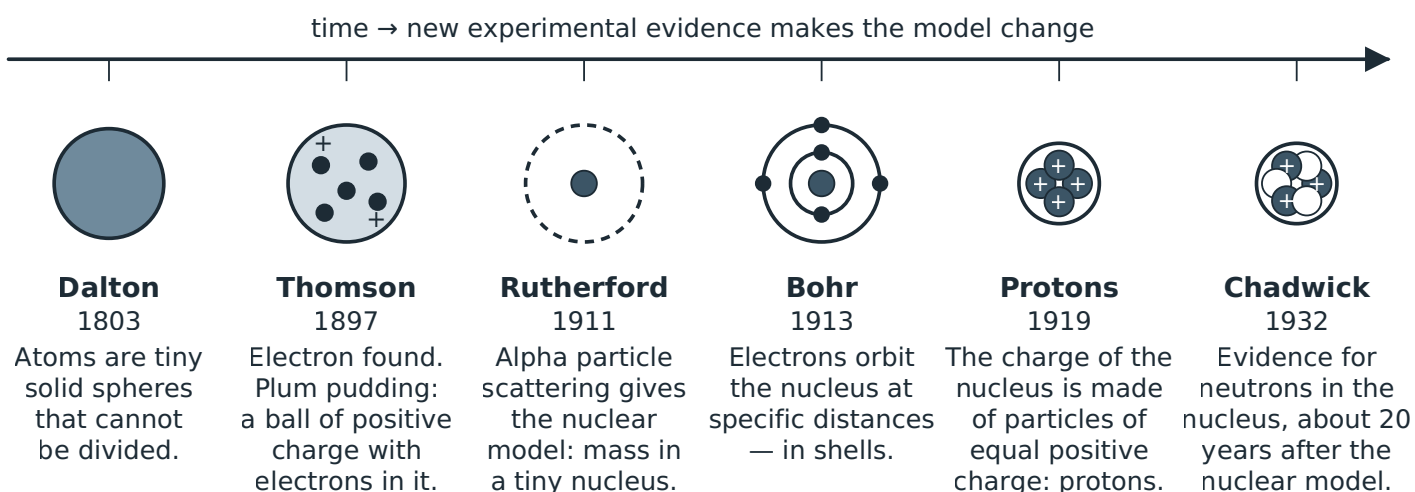
Separating a mixture is a **physical** change — no new substances are made. Only a **chemical reaction** can split a compound into its elements.

1.1.3 The development of the model of the atom

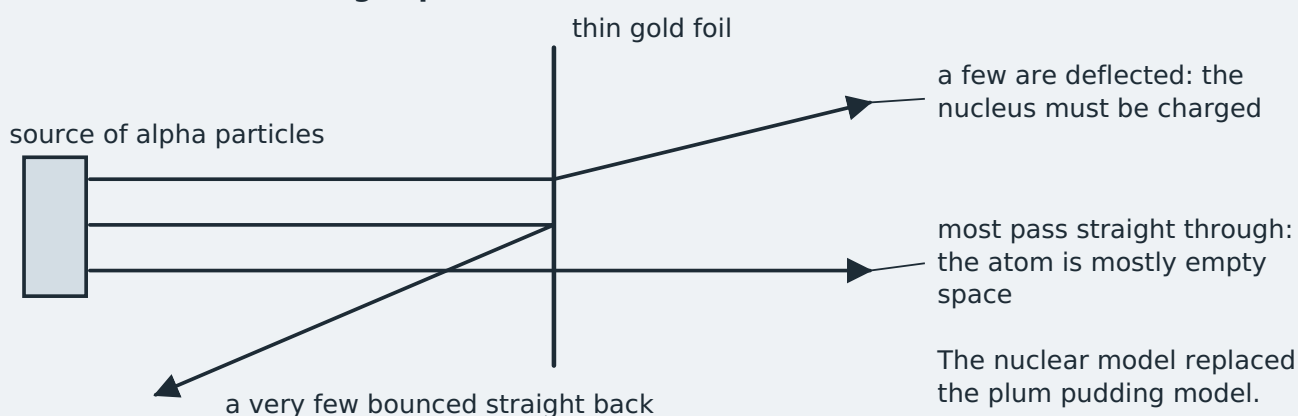
New experimental evidence may lead to a scientific model being changed or replaced. This is a favourite exam question — learn the sequence and the evidence behind each step.

1. **Before the electron was discovered (Dalton):** atoms were thought to be tiny spheres that could not be divided.
2. Discovery of the electron → the plum pudding model: the atom is a ball of positive charge with negative electrons embedded in it.
3. **Alpha particle scattering (Rutherford):** alpha particles fired at thin gold foil — most passed straight through, some were deflected and a very few came back. The results led to the conclusion that the mass of an atom was concentrated at the centre (nucleus) and that the nucleus was charged. This nuclear model replaced the plum pudding model.
4. **Bohr:** adapted the nuclear model by suggesting that electrons orbit the nucleus at specific distances. His theoretical calculations agreed with experimental observations.
5. **Protons:** later experiments led to the idea that the positive charge of any nucleus could be subdivided into a whole number of smaller particles, each with the same amount of positive charge — named protons.
6. **Chadwick:** his experimental work provided the evidence for neutrons within the nucleus, about **20 years** after the nucleus became an accepted scientific idea.

How the model of the atom changed: solid sphere, plum pudding, nuclear, Bohr



Rutherford's scattering experiment



You must be able to describe:

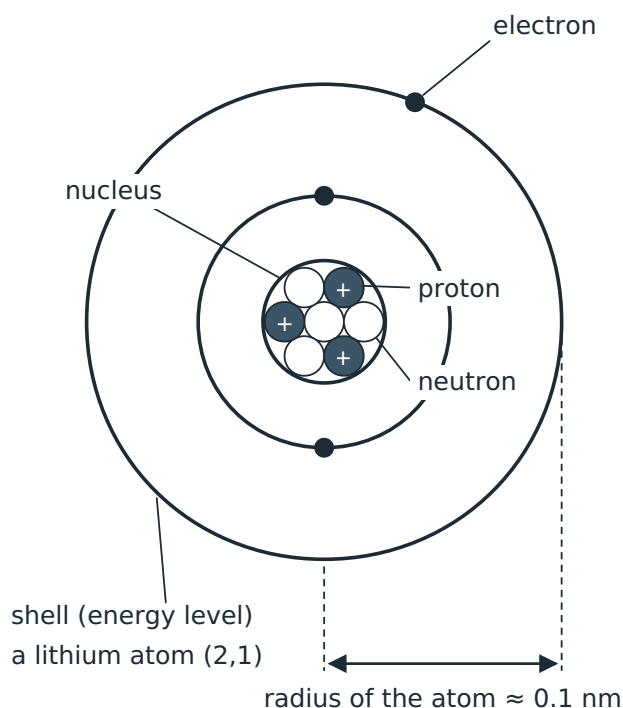
- why the scattering evidence changed the model — the plum pudding model predicted that all alpha particles would pass through with only slight deflection, because charge and mass were spread thinly. As a few were deflected straight back, the positive charge and nearly all the mass had to be concentrated in a tiny, dense, positively charged centre.
- the difference between the plum pudding and nuclear models — plum pudding: positive charge spread through a ball with electrons embedded, no nucleus, no empty space. Nuclear: nearly all the mass and all the positive charge in a tiny central nucleus, electrons outside it, most of the atom empty space.

1.1.4 Relative electrical charges of subatomic particles

Particle	Relative charge	Where it is
Proton	+1	Nucleus
Neutron	0	Nucleus
Electron	-1	Energy levels (shells) around the nucleus

- In an atom the number of electrons equals the number of protons, so atoms have no overall electrical charge.
- The number of protons in an atom of an element is its **atomic number**. All atoms of a particular element have the same number of protons; atoms of different elements have different numbers.
- Use the **nuclear model** to describe atoms: a tiny positive nucleus of protons and neutrons, surrounded by electrons in shells.

Structure of the atom: nucleus of protons and neutrons, electrons in shells



Particle	Relative charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	-1	very small (about 1/1800)

Size and mass of an atom

Radius of an atom ≈ 0.1 nm (1×10^{-10} m).

The radius of the nucleus is less than 1/10 000 of the radius of the atom (about 1×10^{-14} m).

Almost all of the mass of an atom is in the nucleus.

Atomic number = number of protons. Mass number = number of protons + number of neutrons.
In an atom the number of electrons = the number of protons, so an atom has no overall charge.

1.1.5 Size and mass of atoms

- Atoms have a radius of about **0.1 nm** (1×10^{-10} m).
- The radius of a nucleus is **less than 1/10 000** of that of the atom (about 1×10^{-14} m).
- Almost all of the mass of an atom is in the nucleus.

Particle	Relative mass
Proton	1
Neutron	1
Electron	Very small ($\approx 1/1800$)

- The sum of the protons and neutrons is the **mass number**.
- Atoms of the same element can have different numbers of neutrons; these are **isotopes**. Atoms are written with the mass number above and the atomic number below the symbol.

1.1.6 Relative atomic mass

The **relative atomic mass (A_r)** of an element is an average value that takes account of the abundance of the isotopes of the element. This is why many A_r values are not whole numbers.

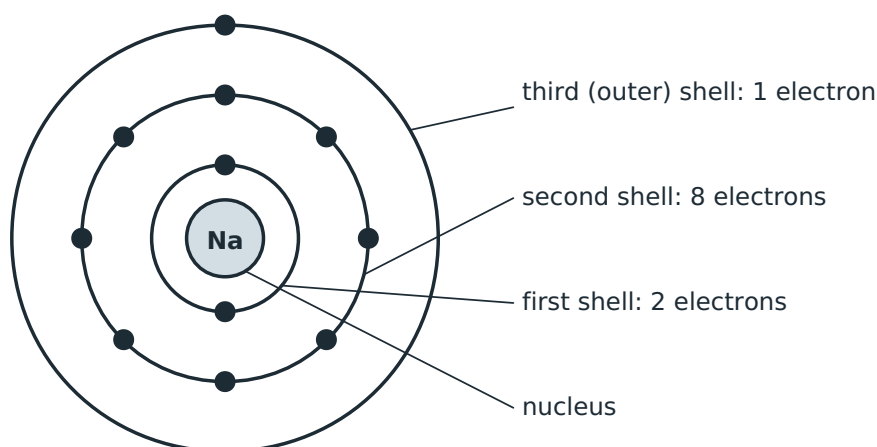
ANALOGY

Isotopes. Two identical cars of the same model — same engine, same badge, so they drive the same — but one carries a heavier load in the boot. Isotopes have the same protons and electrons, so identical **chemical** properties, but different numbers of neutrons and so different masses.

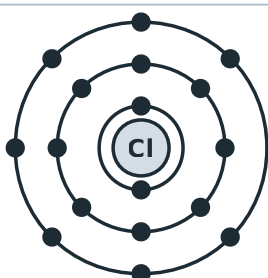
1.1.7 Electronic structure

- Electrons occupy the **lowest available energy levels** (the innermost available shells).
- Electronic structure is written as numbers or drawn as a **diagram**. Sodium is **2,8,1** — two electrons in the lowest energy level, eight in the second, one in the third.
- Shell capacities for the first 20 elements: **2, then 8, then 8**, with the next two electrons in the fourth level (calcium is 2,8,8,2).
- You may answer in terms of either **energy levels** or **shells**, and must be able to represent the structures of the **first twenty elements** in both forms.

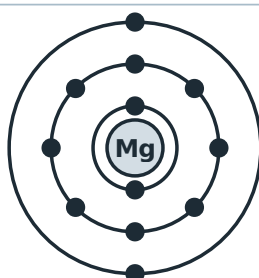
Electronic structures of the first twenty elements



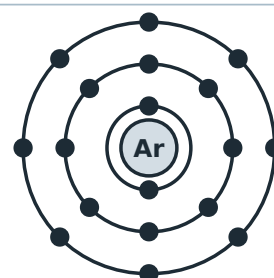
Sodium (Na), 2,8,1
11 protons, so 11 electrons



Chlorine (Cl)
2,8,7
Group 7: 7 outer electrons



Magnesium (Mg)
2,8,2
Group 2: 2 outer electrons



Argon (Ar)
2,8,8
Group 0: full outer shell

- Electrons occupy the lowest available energy levels (the innermost shells) first: 2, then 8, then 8.
- Electrons in the outer shell = the group number. Group 0 atoms have a full outer shell.

1.2 The periodic table

1.2.1 The periodic table

- The elements are arranged in order of **atomic (proton) number**, and so that elements with similar properties are in columns known as **groups**. It is called **periodic** because similar properties occur at **regular intervals**.
- Elements in the same group have the **same number of outer electrons**, which gives them similar chemical properties.
- The **group number** = the number of outer electrons (except Group 0); the **period number** = the number of occupied shells.
- You must be able to explain how an element's position relates to the arrangement of electrons in its atoms and hence to its atomic number, and predict possible reactions and probable reactivity from position.

	Metals	Non-metals
Outer electrons	Few — lost easily	Many — gained or shared
Ions formed	Positive	Negative, or none
Conduction	Good conductors of electricity and thermal energy	Poor conductors
Melting/boiling points	Usually high	Usually low
Appearance	Shiny, malleable, dense	Often dull, brittle when solid

You must be able to explain the differences between metals and non-metals from these characteristic physical and chemical properties, explain how atomic structure relates to position in the table, and explain how the reactions of elements relate to the arrangement of electrons in their atoms and hence to their atomic number.

1.2.5 Group 1

- Group 1 elements are the **alkali metals**, with characteristic properties because of the single electron in their outer shell.
- Reactivity increases going down the group:** the outer shell is further from the nucleus and shielded by more inner shells, so the outer electron is lost more easily.

1.2.4 Group 0

- Group 0 elements are the **noble gases**. They are unreactive and do not easily form molecules because their atoms have **stable arrangements of electrons**.
- They have eight electrons in their outer shell, except **helium**, which has only two.
- The boiling points increase with increasing relative atomic mass (going down the group).
- You must explain how Group 0 properties depend on the outer shell of electrons, and predict properties from given trends down the group.

Reagent	What happens	Equation (sodium)
Oxygen	Tarnish quickly; burn to form a white metal oxide	$4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}$
Chlorine	Burn to form a white metal chloride (an ionic salt)	$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
Water	Float and fizz, giving hydrogen and an alkaline hydroxide solution; Li fizzes steadily, Na melts into a ball, K ignites	$2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$

1.2.6 Group 7

- Group 7 elements are the **halogens**, with similar reactions because they all have seven electrons in their outer shell. They are **non-metals** and consist of molecules made of pairs of atoms (eg Cl_2).
- Going down the group, relative molecular mass, melting point and boiling point increase.
- Going down the group, reactivity decreases — the outer shell is further from the nucleus and more shielded, so an electron is gained less easily.
- With **metals**, the halogens form **ionic** compounds (metal halide salts) in which the halogen is a **1- ion**; with non-metals they form **covalent** molecular compounds with shared pairs of electrons (eg hydrogen chloride).
- A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt.

Property, going down the group	Group 1	Group 7	Group 0
Outer electrons	1	7	8 (helium 2)
Melting and boiling point	Decrease	Increase	Increase
Relative atomic/molecular mass	Increases	Increases	Increases
Reactivity	Increases	Decreases	Unreactive throughout
State at room temperature	All solids	Cl ₂ gas, Br ₂ liquid, I ₂ solid	All gases

WATCH OUT

Group 1 reactivity increases down the group but Group 7 reactivity decreases. The reason is the same in both cases (the outer shell is further from the nucleus and more shielded), but metals must lose an electron and halogens must gain one, so the trend runs the opposite way.

1.3 Properties of transition metals CHEMISTRY ONLY

1.3.1 Comparison with Group 1 elements CHEMISTRY ONLY

CHEMISTRY ONLY The transition elements are metals with similar properties which are different from those of the elements in Group 1. Exemplify these properties by reference to Cr, Mn, Fe, Co, Ni and Cu.

Property	Transition metals	Group 1 metals
Melting point	High	Low, and falls down the group
Density	High	Low — Li, Na and K float on water
Strength and hardness	Strong and hard	Soft — cut with a knife
Reaction with oxygen	Slow; many resist corrosion	Very fast, tarnish immediately
Reaction with water	Slow or none	Vigorous, giving hydrogen and an alkaline solution
Reaction with halogens	Slower	Vigorous

1.3.2 Typical properties CHEMISTRY ONLY

CHEMISTRY ONLY Many transition elements:

- have ions with different charges — eg Fe²⁺ and Fe³⁺, Cu⁺ and Cu²⁺
- form **coloured compounds** — eg copper(II) sulfate blue, iron(II) compounds pale green, iron(III) compounds orange/brown
- are useful as **catalysts** — eg iron in the Haber process, nickel in hydrogenation, manganese(IV) oxide with hydrogen peroxide.

NOTE

NOT REQUIRED BY THE SPECIFICATION

- Details of the experimental work supporting the Bohr model are not required.
- Details of Chadwick's experimental work are not required — only that it gave the evidence for neutrons.
- No knowledge of the structure of the atom beyond the model specified here (a nucleus of protons and neutrons with electrons in energy levels) — no sub-shells or orbitals.
- A periodic table is supplied in the exam, so relative atomic masses and atomic numbers need not be memorised — but the names and symbols of the first 20 elements and of Groups 1 and 7 should be.