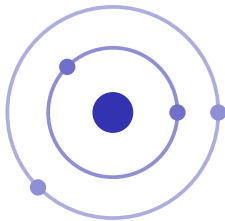


Atomic Structure

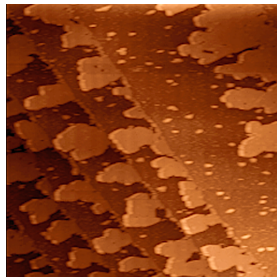
A-Level Chemistry ■ Topic 1

A-Level Chemistry



What we will cover

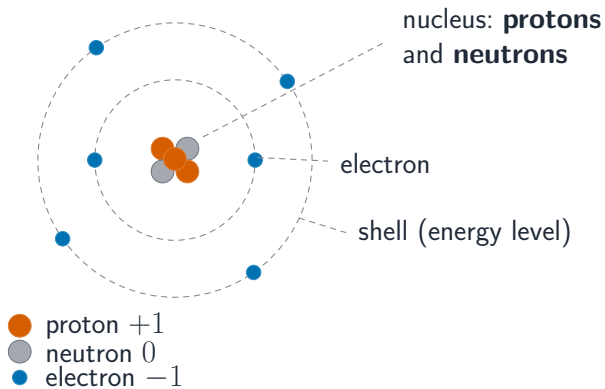
- 1 Inside the atom
- 2 Counting particles & radius
- 3 Isotopes
- 4 Sub-shells & orbitals
- 5 Electronic configuration
- 6 Ionisation energy



real atoms, imaged



Inside the atom



A tiny **nucleus** of **protons** (+1) and **neutrons** (0) holds nearly all the mass; **electrons** (-1, ~0 mass) move in **shells** .

Relative charge and relative mass

We compare the three particles using **relative charge** and **relative mass** .

These are simple numbers, not real units.

Counting particles

proton number Z

nucleon number A

neutrons $= A - Z$

Worked example: ${}_{13}^{27}\text{Al}^{3+}$

13 protons; $27 - 13 = 14$ neutrons; $13 - 3 = 10$ electrons (lost 3 to become $3+$)

A positive **ion** (**cation**) has fewer electrons than protons; a negative ion (**anion**) has more.

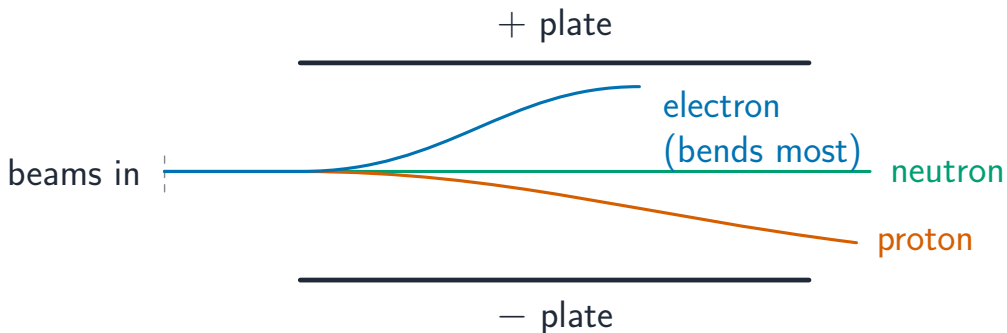
How mass and charge are spread out

Almost all the mass sits in the nucleus, because protons and neutrons are heavy and electrons are very light.

All the positive charge is in the nucleus (the protons).

Beams in an electric field

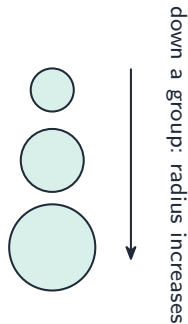
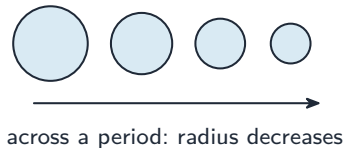
- proton bends to the $-$ plate
- electron bends to the $+$ plate, **far more** (much lighter)
- neutron goes straight (no charge)



Atomic and ionic radius

- across a **period** – **smaller** (stronger **nuclear charge** , same shell)
- down a **group** – **larger** (each step adds a shell)

A cation is smaller than its atom; an anion larger.



Isotopes



17 protons, 18 neutrons

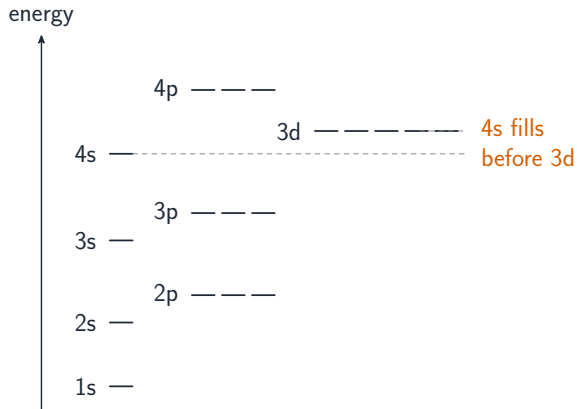


17 protons, 20 neutrons

Isotopes : same protons, different neutrons. Same **chemical** properties (same electrons); different **physical** ones (mass, density).



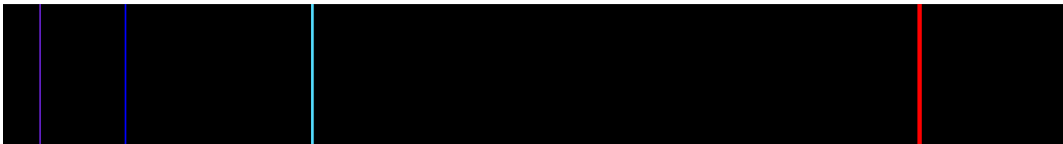
Sub-shells and energy order



Shells split into **sub-shells** s (2), p (6), d (10), made of **orbitals** (2 electrons each).

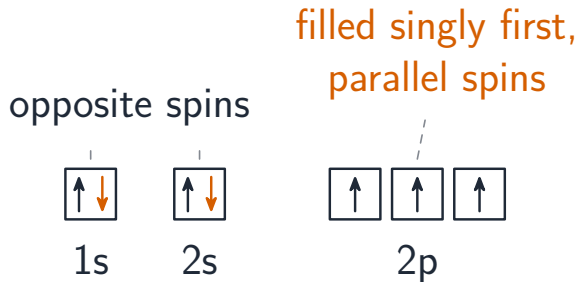
Fill lowest energy first – note
4s fills before 3d.

Hydrogen emits light only at discrete



Hydrogen emits light only at discrete, characteristic wavelengths – its line emission spectrum.

Electronic configuration



Electronic configuration : Fe
 $= [\text{Ar}] 3d^6 4s^2$.

- paired electrons have opposite **spin**
- orbitals fill singly first (less repulsion)

Ions lose 4s before 3d:
 $\text{Fe}^{3+} = [\text{Ar}] 3d^5$.

Filling order, shorthand, and the boxes

the surprise

4s is slightly lower in energy than 3d, so 4s fills first – and empties first when an ion forms

the order

1s 2s 2p 3s 3p 4s 3d 4p ...

the shorthand

use the nearest **noble gas** in square brackets – [Ar]4s² stands for argon's full configuration plus two 4s electrons

in boxes

within a sub-shell, electrons fill **empty orbitals one at a time**, with parallel arrows, before any orbital gets a second

why

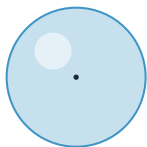
electrons repel, so they stay apart while they can – Hund's rule

the ground state

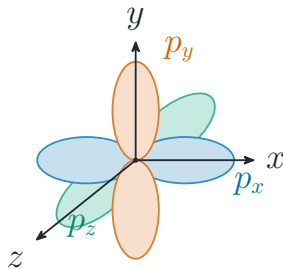
the lowest-energy arrangement; anything else is excited

The 4s-before-3d order is why transition metals lose 4s electrons first when they ionise – the point every configuration question in Topic 28 rests on.

Shapes of orbitals



s orbital
(sphere)



p orbitals
(two lobes each)

- an **s orbital** s is a **sphere**
- a **p orbital** p is a dumbbell;
the three point along x , y , z

A **free radical** has an unpaired electron.

Free radicals

A **free radical** is a species with one or more **unpaired electrons** .



First ionisation energy

First ionisation energy : $X(g) \rightarrow X^+(g) + e^-$.

- rises across a period (more nuclear charge)
- dips at Al ($3p < 3s$) and S (**spin-pair repulsion**)

The dips are evidence for [sub-shells](#).

Defining first ionisation energy exactly

the definition

the energy to remove one electron from each atom in **one mole of gaseous atoms**

why gaseous

so there are **no forces between the particles** to include in the measurement

the equation



the unit

kJ mol^{-1}

what it depends on

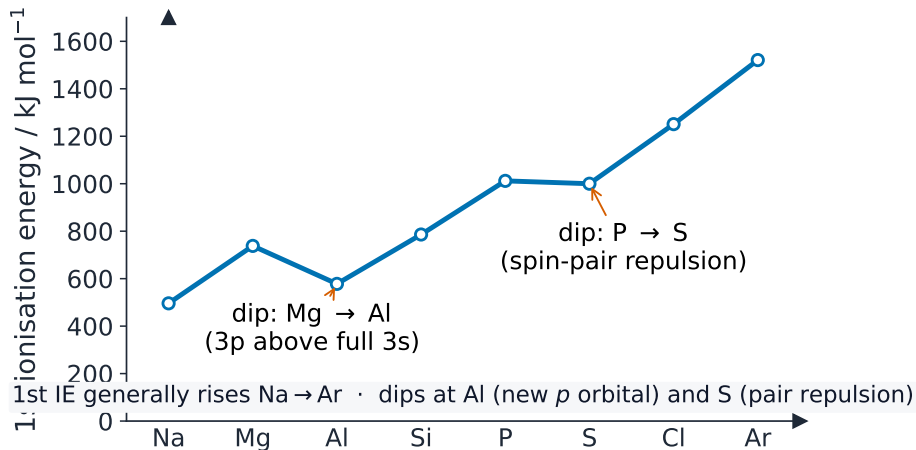
nuclear charge, **atomic radius**, and shielding by inner shells

the trend

increases across a period, decreases down a group – with the small drops that reveal sub-shell structure

The state symbols **(g)** are part of the definition, not decoration. Omitting them loses the mark even when the words are right.

First ionisation energy, in detail



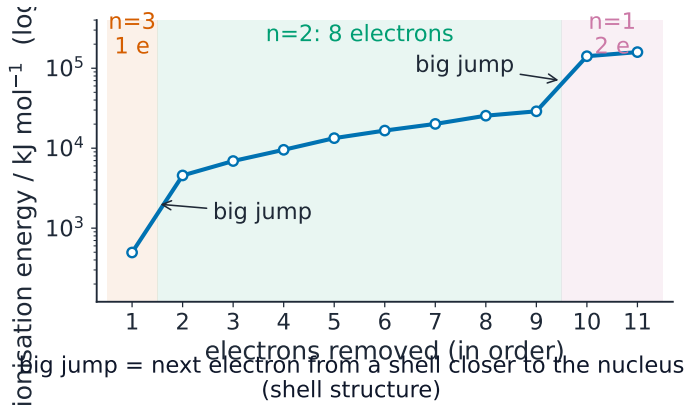
First ionisation energy rises across Period 3 but dips at Al and at S – evidence that sub-shells exist

Successive ionisation energies

Each **successive ionisation energy** is larger (a more positive ion).

Big jumps reveal the shells: sodium's 2, 8, 1 pattern. Electrons removed before the first jump = the group number.

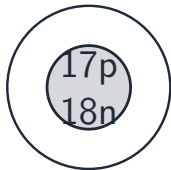
Successive ionisation energies, in detail



Each **successive ionisation energy** is larger (a more positive ion).

Isotopes, continued

same protons, different neutrons



Two chlorine isotopes: same protons, different neutrons

What isotopes change, and what they do not

the atomic number

the proton count – identical between isotopes, so the element is the same

the mass number

protons plus neutrons – this is what differs

chemical properties

identical, because they depend on the electron arrangement, which is unchanged

physical properties

some **depend on mass**, so they differ: a heavier isotope diffuses more slowly and has a higher density and boiling point

neutral atoms

protons equal electrons; an ion has gained or lost electrons, never protons

the atomic radius

unchanged between isotopes, since it is set by the electron shells

Chemistry is electrons; mass effects are physics. That single split answers every isotope question on the paper.

Exam tips

- Define **isotopes** in full: atoms of the same element with the same number of protons but a different number of neutrons – the mark scheme wants both halves.
- 4s fills **before** 3d, but electrons are **removed from 4s first** when forming ions, so Fe^{3+} is $[\text{Ar}]3\text{d}^5$ (not $[\text{Ar}]3\text{d}^34\text{s}^2$).
- In successive ionisation energies, a **big jump** marks the start of a new (inner) shell – use the jumps to place the element in its group.
- Explain every ionisation-energy trend with the same three factors: **nuclear charge, distance and shielding**, plus sub-shell effects for the small dips.
- Learn the exact reason first ionisation energy of oxygen is **below** nitrogen: oxygen's **paired 2p electrons repel**, so one is easier to remove.



Topic 1 – key takeaways

1

The atom

nucleus (p,n) + electrons in shells

2

Isotopes

same protons, different neutrons

3

Configuration

s, p, d sub-shells; 4s fills before 3d

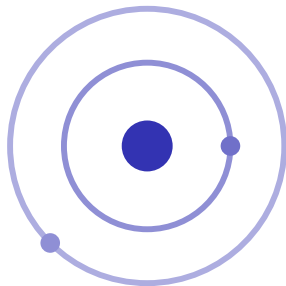
4

Ionisation energy

dips show sub-shells; jumps show shells

Check yourself

- 1 How many neutrons in ${}_{17}^{37}\text{Cl}$?
- 2 Which fills first, 4s or 3d?
- 3 Why does first IE dip from Mg to Al?
- 4 What do big jumps in successive IE reveal?



Answers 1. 20 2. 4s 3. Al's electron leaves a higher-energy 3p sub-shell 4. the shells