

Atomic structure

A-Level Chemistry 9701 · Topic 1 revision notes

Everything from the 6 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Particles in the atom

2. Atomic and ionic radius

3. Isotopes

4. Shells, orbitals and electron configuration

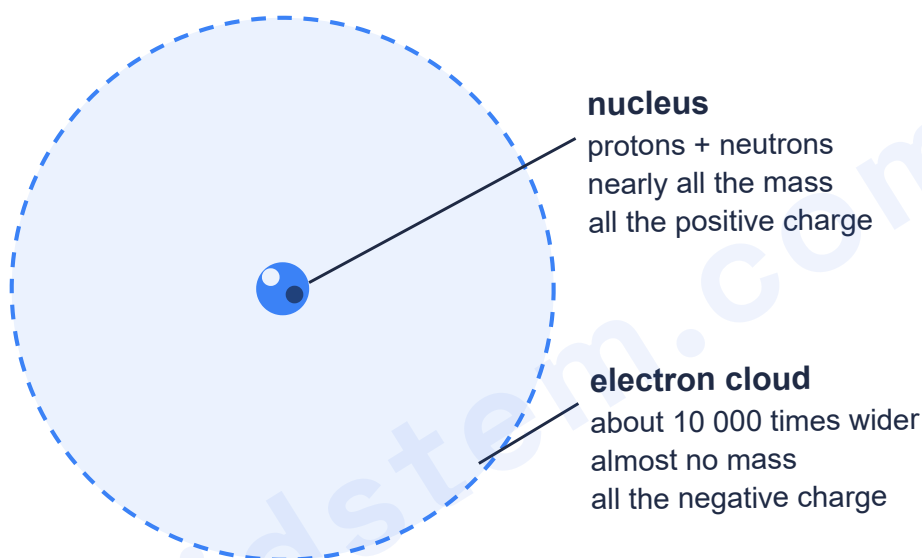
5. Ionisation energy and its trends

6. Successive ionisation energies and electron affinity

1 Particles in the atom

Syllabus 1.1

Protons (relative charge +1, relative mass 1) and neutrons (0, 1) sit in the nucleus, so the nucleus carries nearly all the mass and all the positive charge. Electrons (-1, $1/1836$) fill the space around it. Proton number Z counts protons; nucleon number A counts protons plus neutrons. Neutrons = $A - Z$, and electrons = $Z - (\text{charge on the ion})$.



not to scale: a real nucleus would be far too small to draw

Fig. 1: where the mass and the charge live

Two numbers give everything.

- ✓ **Z** (bottom) = protons = electrons in the neutral atom.
- ✓ **A** (top) = protons + neutrons, so neutrons = $A - Z$.
- ✓ **Ions:** the charge changes electrons only. Positive means electrons lost.
- ✓ **Fields:** deflection follows the sign of the charge and the size of charge/mass.

Watch out!

The charge on an ion changes the **electron** count only. Protons and neutrons are untouched: an ion is still the same element with the same nucleus. And the nucleon number on a symbol is a whole-number count for one atom; it is not the relative atomic mass in the Periodic Table, which is an average over isotopes.

2 Atomic and ionic radius

Syllabus 1.1

Across a period the radius decreases: each step adds one proton and one electron to the same shell, shielding barely changes, so the outer electrons are pulled in harder. Down a group the radius increases: each period adds a whole new shell, further out and more shielded. A cation is smaller than its atom (the outer shell is lost), an anion is larger (extra repulsion), and in an isoelectronic series the ion with the most protons is the smallest.

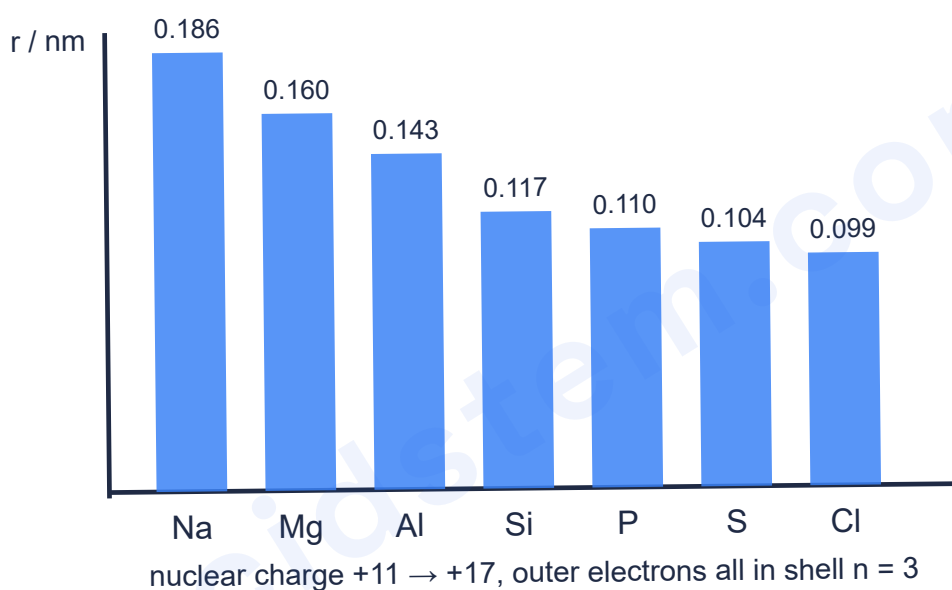


Fig. 1: atomic radius across Period 3 (typical values)

Name the factor every time.

- ✓ **Nuclear charge:** more protons, stronger pull, smaller radius.
- ✓ **Shells:** a new shell means further out and larger.
- ✓ **Shielding:** inner shells cancel part of the pull; electrons in the same shell shield poorly.
- ✓ **Ions:** protons never change; only the electron cloud does.

Watch out!

Ionic radius across Period 3 is not a smooth decrease. It falls from Na^+ to Mg^{2+} to Al^{3+} (all 10 electrons, $n = 2$ outer shell), then **jumps up** at P^{3-} , S^{2-} and Cl^- , which have 18 electrons and an $n = 3$ outer shell, and then falls again along that second isoelectronic set.

3 Isotopes

Syllabus 1.2

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons, so the same proton number Z and a different nucleon number A . They have the same electron configuration and therefore the same chemical properties. Physical properties that depend on mass, such as density and rate of diffusion, differ slightly. Relative atomic mass is the weighted mean of the isotopic masses, which is why it is rarely a whole number.

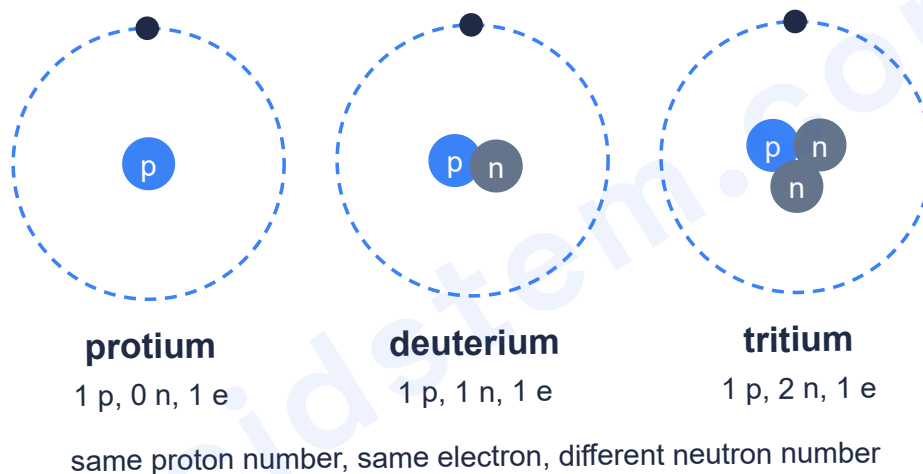


Fig. 1: the three isotopes of hydrogen

**Protons fix the element,
neutrons fix the isotope.**

- ✓ **Definition:** same number of protons, different number of neutrons.
- ✓ **Chemistry:** identical, because the electron configuration is identical.
- ✓ **A_r :** sum of (mass $\times$ fraction) over all isotopes.
- ✓ **Reverse:** let the unknown fraction be x and solve one equation.

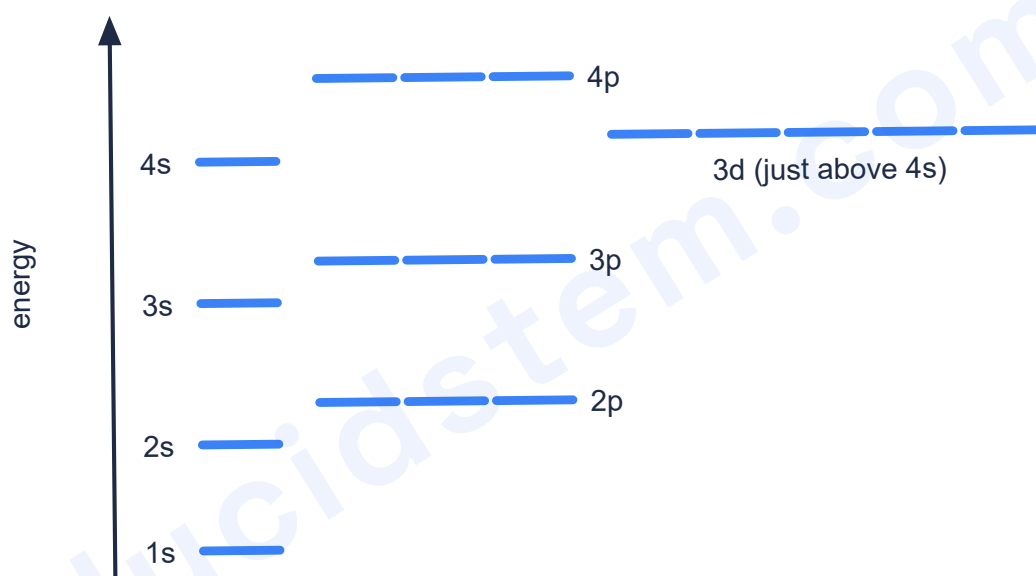
Watch out!

Isotopes have the same number of **protons and electrons**; only the neutron number differs. Saying that isotopes have a different atomic number, or a different number of electrons, loses the mark. And the mass number is a property of one atom; the relative atomic mass is a property of the natural mixture.

4 Shells, orbitals and electron configuration

Syllabus 1.3

Shell n holds s , p , d and f sub-shells up to n of them; an s sub-shell has one orbital, p has three, d has five, and each orbital holds two electrons with opposite spins. Sub-shells fill in order of increasing energy, $1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p$, so $4s$ fills before $3d$. Within a sub-shell, electrons occupy orbitals singly before pairing. Chromium and copper break the pattern to give half-filled and full $3d$ sub-shells.



each dash is one orbital holding up to two electrons; fill from the bottom

Fig. 1: sub-shell energies and the filling order up to $4p$

One order, two exceptions.

- ✓ **Order:** $1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p$. Capacity 2, 6, 10.
- ✓ **Hund:** singly first, then pair.
- ✓ **Ions:** cations lose $4s$ before $3d$; anions fill the next empty orbital.
- ✓ **Exceptions:** $\text{Cr}\ 3d^5 4s^1$, $\text{Cu}\ 3d^{10} 4s^1$.

Watch out!

$4s$ is filled **before** $3d$ and emptied **before** $3d$. Writing Fe^{2+} as $[\text{Ar}]\ 3d^4\ 4s^2$ is the classic error. And chromium is $3d^5\ 4s^1$, not $3d^4\ 4s^2$; copper is $3d^{10}\ 4s^1$, not $3d^9\ 4s^2$.

5 Ionisation energy and its trends

Syllabus 1.3

First ionisation energy is the energy needed to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous $1+$ ions: $X(g) \rightarrow X^+(g) + e^-$. It is larger when the nuclear charge is greater, the outer electron is closer, and shielding is lower. Across Period 3 it generally rises, with dips at aluminium ($3p$ sits above $3s$) and sulfur (spin-pair repulsion in $3p^4$). Down a group it falls because the outer electron is further away and more shielded.

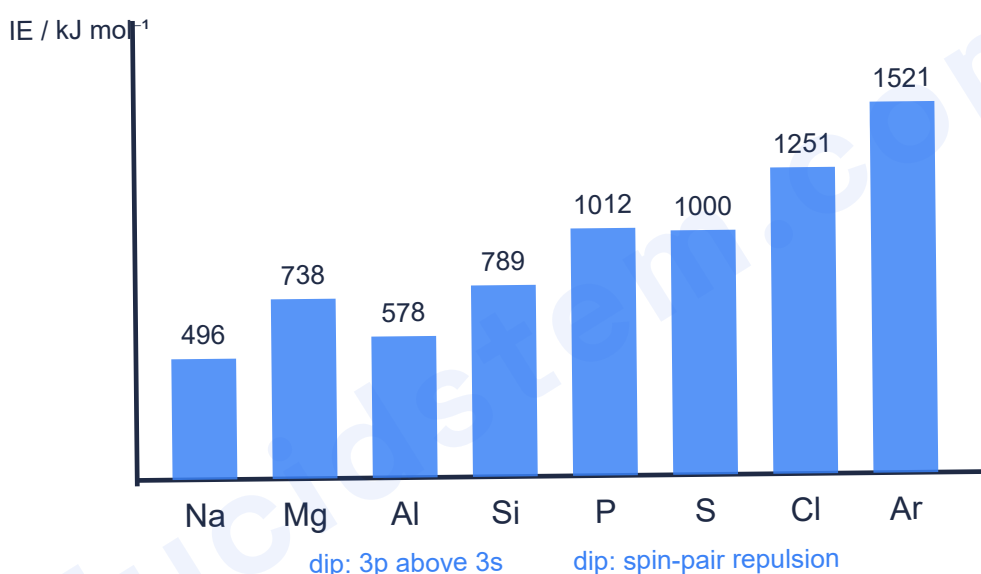


Fig. 1: first ionisation energies across Period 3 (kJ per mole)

Charge, distance, shielding, pairing.

- ✓ **Definition:** one mole, gaseous atoms, gaseous $1+$ ions, one electron each.
- ✓ **Across:** nuclear charge up, same shell, similar shielding: IE rises.
- ✓ **Dips:** Al ($3p$ above $3s$), S (spin-pair repulsion in $3p^4$).
- ✓ **Down:** further out and more shielded: IE falls.

Watch out!

Three things lose marks in a definition: forgetting **gaseous**, forgetting **one mole**, and writing a $2+$ ion for a first ionisation. In explanations, say **shielding** or **outer electron further from the nucleus**; "more electrons" on its own is not an explanation.

6 Successive ionisation energies and electron affinity

Syllabus 1.3

Successive ionisation energies always increase: each electron leaves an ion that is more positive, so the same nuclear charge holds fewer electrons more tightly and the ion shrinks. A large jump appears when the next electron comes from an inner shell, closer to the nucleus and less shielded. The number of electrons removed before the first large jump equals the number of outer-shell electrons, which is the group number. Electron affinity, $X(g) + e^- \rightarrow X^-(g)$, is exothermic for the first electron but endothermic for a second, because a negative ion repels the incoming electron.

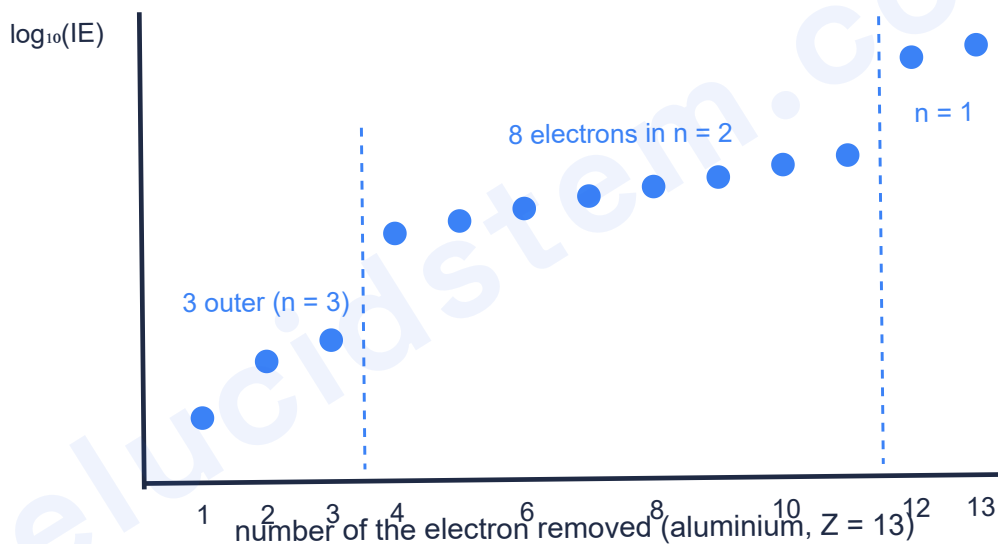


Fig. 1: log of the successive ionisation energies of aluminium

Steps within a shell, jumps between shells.

- ✓ **Always increasing:** more positive ion, same nuclear charge, fewer electrons.
- ✓ **Jump:** next electron comes from an inner shell, closer and less shielded.
- ✓ **Group:** electrons removed before the first jump = outer electrons.
- ✓ **Electron affinity:** first usually exothermic, second always endothermic.

Watch out!

Count the electrons removed **before** the jump, not the position of the jump itself. A large increase between the 2nd and 3rd ionisation energies means two outer electrons, Group 2. And successive ionisation energies are always positive and always increase; a smaller number than the one before is a sign you have misread the table.

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