

Atomic structure

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What you need to know

By the end of this topic you should be able to:

1. **Picture the atom.** It is mostly empty space. At the centre is a tiny, dense nucleus of protons and neutrons; electrons are arranged in shells in the space around it.
2. **Give the relative charge and relative mass** of a proton, a neutron and an electron.
3. **Use the terms** proton number (atomic number) and nucleon number (mass number).
4. **Describe where the mass and the charge are:** almost all the mass, and all the positive charge, is in the nucleus; the negative charge is spread through the space around it.
5. **Describe how beams of protons, neutrons and electrons behave in an electric field** when they move at the same speed.
6. **Count the protons, neutrons and electrons** in any atom or ion from its proton number, nucleon number and charge.
7. **State and explain how atomic and ionic radius change** across a period and down a group.
8. **Define isotope** in terms of protons and neutrons, and read the notation ${}^A_Z\text{X}$ (nucleon number on top, proton number below).
9. **Explain why isotopes of an element have the same chemical properties** but different physical properties (mass and density only).
10. **Use the terms** shell, sub-shell, orbital, principal quantum number n and ground state.
11. **Know how many orbitals, and how many electrons, fit in s, p and d sub-shells**, and the order of their energies from 1s up to 4p.
12. **Write electronic configurations** of atoms and ions from H to Kr (in full or shorthand, such as $[\text{Ar}] 3d^6 4s^2$), say how many electrons are in each shell, sub-shell and orbital, draw them as **electrons in boxes**, and explain the arrangement using the energy of the electrons and the repulsion between them.
13. **Describe and sketch the shapes of s and p orbitals.**
14. **Describe a free radical:** a species with one or more unpaired electrons.
15. **Define first ionisation energy** and write equations for first, second and later ionisation energies.
16. **Understand that ionisation energies come from the attraction between the nucleus and the outer electron.** Use this to **explain the trends in first ionisation energy** across a period and down a group, and the pattern in successive ionisation energies of one element, using four factors: nuclear charge, distance

of the electron from the nucleus (atomic or ionic radius), shielding by inner shells and sub-shells, and spin-pair repulsion.

17. **Use successive ionisation energies** to deduce an element's electronic configuration and its position in the Periodic Table.

Only ground-state atoms and ions from hydrogen to krypton are tested. d orbital shapes, energy-level diagrams for d orbitals and quantum numbers other than n are **not** part of this topic.

Understand it

Inside the atom

An atom has a **nucleus** at its centre, about 10^{-15} m across, inside an atom about 10^{-10} m across. So the atom is almost all empty space. The nucleus holds the **protons** and **neutrons** (together called **nucleons**). The **electrons** move in the space around the nucleus, arranged in shells.

Particle	Where it is	Relative charge	Relative mass
proton	nucleus	+1	1
neutron	nucleus	0	1
electron	shells around the nucleus	-1	$\frac{1}{1836}$ (almost 0)

So nearly all the **mass** of an atom, and all its positive charge, is packed into the tiny nucleus. The electrons carry the negative charge, spread out in the large space around it, but add almost nothing to the mass.

- **Proton number (atomic number), Z** : the number of protons in the nucleus. It decides which element the atom is.
- **Nucleon number (mass number), A** : the number of protons **plus** neutrons in the nucleus.

In the notation A_ZX the nucleon number is on top and the proton number below: ${}^{37}_{17}\text{Cl}$ has 17 protons and $37 - 17 = 20$ neutrons.

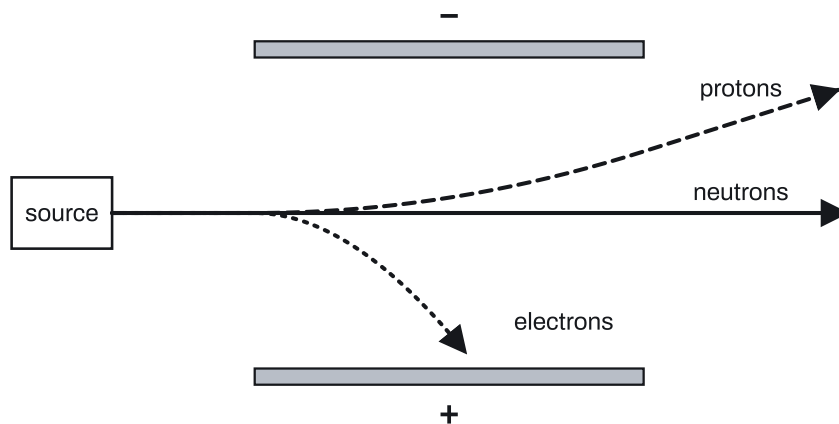
Counting particles in atoms and ions

- **protons** = Z (from the Periodic Table if not given);
- **neutrons** = $A - Z$;
- **electrons** = $Z - \text{charge}$. A $2+$ ion has lost 2 electrons; a $2-$ ion has gained 2.

For ${}^{23}\text{Na}^+$: 11 protons, 12 neutrons, 10 electrons. For ${}^{16}\text{O}^{2-}$: 8 protons, 8 neutrons, 10 electrons. For a group of atoms, add up every atom first: NH_4^+ has $7 + 4 = 11$ protons and $11 - 1 = 10$ electrons.

Beams in an electric field

Fire beams of protons, neutrons and electrons, all at the same speed, between two charged plates.



- **Neutrons** have no charge, so they go straight through.
- **Protons** are positive, so they are pulled towards the **negative** plate.
- **Electrons** are negative, so they bend towards the **positive** plate, the opposite way. They bend **much more** than protons, because the force on each is the same size (equal and opposite charges) but an electron is about 1836 times lighter.

The size of the bend depends on **charge ÷ mass**. A He^+ ion (charge +1, mass 4) has a quarter of a proton's charge-to-mass ratio, so it bends the same way as a proton but less.

Isotopes

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons. Chlorine-35 and chlorine-37 both have 17 protons; one has 18 neutrons and the other 20.

- **Same chemical properties:** chemistry is about electrons. Isotopes have the same number of electrons in the same arrangement (the same electronic configuration), so they react in the same way.
- **Different physical properties:** the extra neutrons make the atoms heavier, so the isotopes have different **masses** and **densities**.

Shells, sub-shells and orbitals

Electrons sit in **shells** (energy levels) numbered by the **principal quantum number** $n = 1, 2, 3, 4, \dots$ Larger n means the electron is, on average, further from the nucleus and higher in energy.

Each shell is split into **sub-shells**, labelled s, p and d. Each sub-shell is made of **orbitals**. An **orbital** is a region of space where there is a high chance of finding an electron, and it holds **at most two electrons, with opposite spins**.

Sub-shell	Orbitals	Maximum electrons
s	1	2
p	3	6
d	5	10

Shell $n = 1$ has only 1s; $n = 2$ has 2s and 2p; $n = 3$ has 3s, 3p and 3d; $n = 4$ starts with 4s and 4p. So the shells hold 2, 8, 18, . . . electrons.

Order of energy. The sub-shells rise in energy in this order:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$$

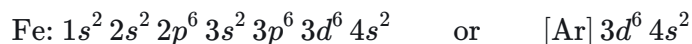
Notice that **4s is lower in energy than 3d**, so 4s fills before 3d.

Electronic configuration

The **ground state** is the arrangement with the lowest energy. To build it:

1. **Fill the sub-shells in order of energy**, lowest first.
2. **Each orbital takes two electrons at most**, with opposite spins.
3. **Within a sub-shell, electrons go into separate orbitals first, with parallel spins, and only pair up when every orbital has one.** Electrons repel each other, so two electrons in one orbital (close together) have more energy than two in separate orbitals.

Iron, $Z = 26$:



[Ar] is shorthand for argon's configuration, $1s^2 2s^2 2p^6 3s^2 3p^6$. Write 3d before 4s (grouping by shell), even though 4s fills first.

Two exceptions to learn: chromium and copper move one 4s electron into 3d, because a half-full or full d sub-shell is more stable.

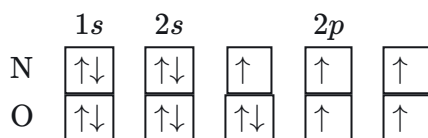


Ions.

- **Negative ions:** add electrons into the next spaces. Cl^- is $1s^2 2s^2 2p^6 3s^2 3p^6$, the same as argon.
- **Positive ions:** remove electrons from the **outermost shell** first. For Period 1–3 elements that is the last sub-shell filled. For the elements Sc to Zn, **the 4s electrons go before the 3d**, because once the 3d is filled 4s is the outer shell. So Fe^{2+} is $[\text{Ar}] 3d^6$ and Fe^{3+} is $[\text{Ar}] 3d^5$; Cu^+ is $[\text{Ar}] 3d^{10}$.

Electrons in boxes

Each box is one orbital; each arrow is one electron, pointing up or down for its spin. Nitrogen ($1s^2 2s^2 2p^3$) and oxygen ($1s^2 2s^2 2p^4$):

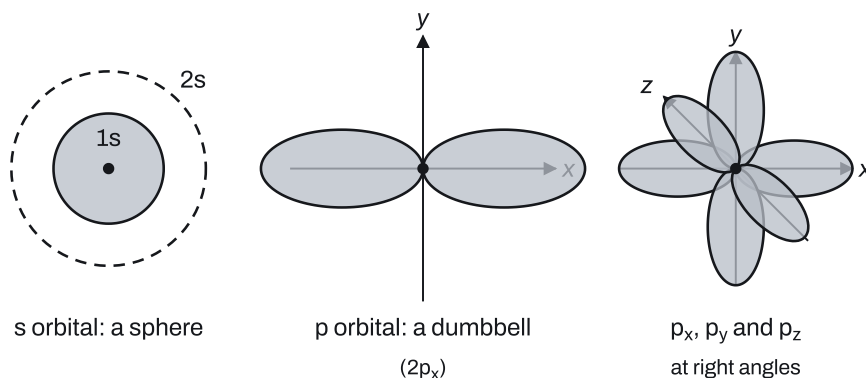


Nitrogen's three 2p electrons are each on their own, with parallel spins (3 unpaired electrons). Oxygen's fourth 2p electron has to pair up, so oxygen has one **paired** p orbital and 2 unpaired electrons. That paired orbital matters for ionisation energy.

To count **unpaired electrons**, draw the boxes for the last sub-shell: Co ($3d^7$) has two paired and three single d orbitals, so 3 unpaired; V ($3d^3 4s^2$) has 3; Ni ($3d^8 4s^2$) has 2.

A **free radical** is a species with one or more unpaired electrons: for example a Cl atom ($3p^5$, one unpaired) or an O atom (two unpaired). A Cl^- ion (all paired) is not.

Shapes of s and p orbitals



- An **s orbital** is a **sphere** centred on the nucleus. A 2s orbital is a bigger sphere than 1s, and 3s is bigger again.
- A **p orbital** is a **dumbbell**: two lobes on opposite sides of the nucleus. The three p orbitals of a sub-shell point along the x , y and z axes, at right angles to each other.

Atomic and ionic radius

Across a period, atoms get smaller. Each element has one more proton, so the nuclear charge increases. The new electron goes into the same shell, and shielding by the inner shells stays about the same, so the outer electrons are pulled in more strongly.

Down a group, atoms get bigger. Each element has one more occupied shell, so the outer electrons are further from the nucleus and more shielded. This outweighs the extra nuclear charge.

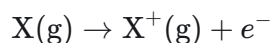
Ions:

- A **positive ion is smaller than its atom**. Na (2, 8, 1) loses its only third-shell electron, so Na^+ (2, 8) has one shell fewer, and the same nuclear charge now pulls on fewer electrons.
- A **negative ion is larger than its atom**. The extra electrons repel each other and the nuclear charge is unchanged, so the electron cloud spreads out.
- **Isoelectronic species** have the same number of electrons. Among them, **more protons means a smaller particle**, because the same electrons feel a bigger nuclear charge. S^{2-} , Cl^- , Ar, K^+ and Ca^{2+} all have 18 electrons, with 16 to 20 protons, so they get smaller in that order.

- Down a group, ions with the same charge get bigger (more shells).

Ionisation energy

The **first ionisation energy** (IE_1) is the energy needed to remove one electron from each atom in one mole of gaseous atoms, forming one mole of gaseous $1+$ ions:

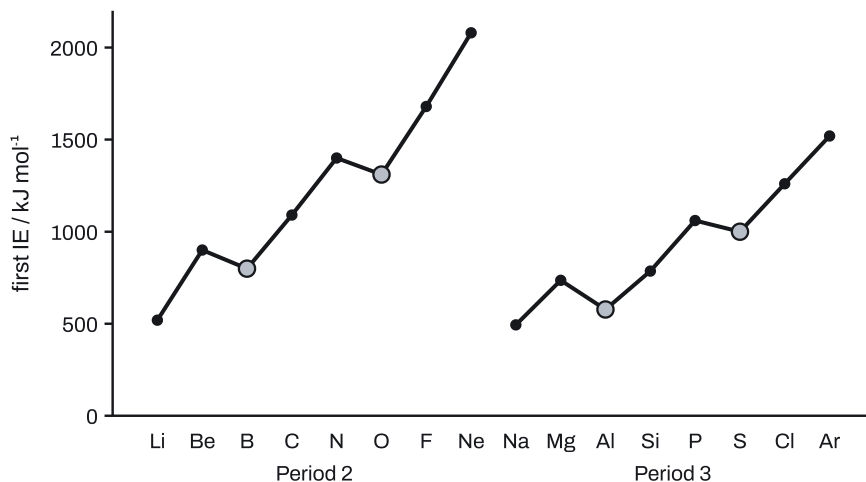


The **second** ionisation energy removes an electron from each $1+$ ion: $X^+(g) \rightarrow X^{2+}(g) + e^-$, and so on; the n th is $X^{(n-1)+}(g) \rightarrow X^{n+}(g) + e^-$. Ionisation energies are always **positive** (endothermic): energy is needed to pull an electron away from the nucleus that attracts it.

How big it is depends on **how strongly the nucleus attracts the electron being removed**. Four factors decide that:

- **Nuclear charge**: more protons, stronger attraction, higher IE.
- **Distance from the nucleus** (atomic or ionic radius): an electron further out is attracted less, lower IE.
- **Shielding** by inner shells and sub-shells: inner electrons repel the outer one and cancel part of the nuclear charge, lower IE.
- **Spin-pair repulsion**: an electron sharing an orbital with another is repelled by it, so it is easier to remove, lower IE.

First ionisation energy across a period and down a group



The general rise across a period. Nuclear charge increases, the electrons are removed from the same shell, and shielding is about the same, so the attraction for the outer electron increases (and the atom gets smaller).

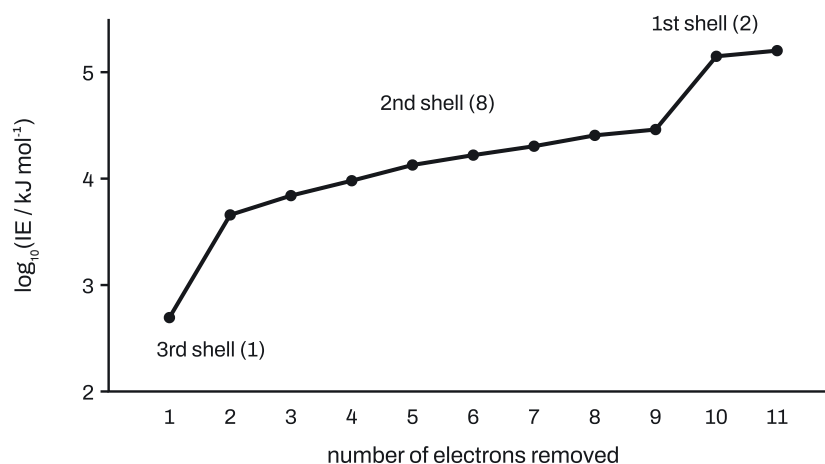
The dip at Group 13 (Be → B, Mg → Al). Aluminium's outer electron is in 3p; magnesium's is in 3s. The **3p sub-shell is higher in energy** and further from the nucleus, and the 3p electron is partly **shielded by the 3s electrons**, so it is easier to remove despite the extra proton.

The dip at Group 16 (N → O, P → S). Sulfur's fourth 3p electron has to go into an orbital that already has an electron in it. The **two electrons in that 3p orbital repel each other (spin-pair repulsion)**, and this outweighs the extra nuclear charge, so one of them is easier to remove. Phosphorus has three unpaired 3p electrons, one per orbital, with no such repulsion.

Down a group, IE_1 falls. The outer electron is in a shell further from the nucleus and more shielded by inner shells. That outweighs the greater nuclear charge.

Helium versus hydrogen. Both lose a 1s electron, but helium has two protons, so its electron is held more tightly (the distance and shielding barely change).

Successive ionisation energies



Taking electrons off one by one, each ionisation energy is **larger** than the one before. The same number of protons now attracts fewer electrons, there is less repulsion between electrons, and the ion is smaller, so the next electron is held more tightly.

There is a **big jump** each time the next electron has to come from a **new shell nearer the nucleus**: it is closer, less shielded and attracted much more strongly. Sodium (2, 8, 1) has a big jump between the 1st and 2nd IE (after its single outer electron) and another between the 9th and 10th (after the eight second-shell electrons).

So the number of electrons removed before the first big jump is the number of outer-shell electrons, which is the group number for Groups 1 and 2, and the group number minus 10 for Groups 13 to 18. Smaller jumps can show a change of sub-shell within a shell (for example 3p to 3s).

Key facts and formulas

"Given" means it is printed in the exam (the Periodic Table in the paper). "Learn it" means you must remember it.

Formula	Status
Proton number and relative atomic mass of each element	Given

Formula	Status
Relative charges +1, 0, −1 and relative masses 1, 1, $\frac{1}{1836}$ of p, n, e	Learn it
neutrons = $A - Z$; electrons = $Z - \text{charge}$	Learn it
Orbitals per sub-shell: s 1, p 3, d 5; electrons: s 2, p 6, d 10; 2 per orbital, opposite spins	Learn it
Energy order: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$	Learn it
Cr [Ar] $3d^5 4s^1$ and Cu [Ar] $3d^{10} 4s^1$	Learn it
Transition-element ions lose 4s electrons before 3d	Learn it
n th IE: $X^{(n-1)+}(g) \rightarrow X^{n+}(g) + e^{-}$	Learn it
Factors: nuclear charge, distance (radius), shielding, spin-pair repulsion	Learn it
$A_r = \frac{\sum(\text{relative isotopic mass} \times \text{abundance})}{\sum \text{abundance}}$	Learn it

How to do it

In the questions we have tagged for this topic (107 questions, 2021–2025, Papers 1 and 2), the types came up this often. 30 of the 107 questions mix more than one type, so the counts add up to 192.

Question type	Questions
Count protons, neutrons and electrons	32
Write or choose an electronic configuration	26
Use successive ionisation energies	25
Unpaired electrons, electron pairs and orbital facts	24
Explain first ionisation energy trends and anomalies	24
Write an equation for (or define) an ionisation energy	16
Explain or order atomic and ionic radii	15
Relative atomic mass from isotope data	12
Define isotopes; explain their properties	8
Sketch or recognise s and p orbitals	6
Particle beams in an electric field; where the mass is	2
Identify free radicals	2

1. Count protons, neutrons and electrons (32 questions)

- Protons** = Z : from the symbol's lower number or the Periodic Table.
- Neutrons** = $A - Z$, from the nucleon number (top number, or "iron-56").

3. **Electrons** = Z – charge: subtract for a positive ion, add for a negative one.

Variations you will meet:

- **Polyatomic ions** (NH_4^+ , OH^- , SO_4^{2-} , N_3^-): add the protons of every atom, then adjust for the charge to get the electrons. Neutrons: add each atom's $A - Z$.
- **An unusual isotope inside a formula**, such as "a water molecule with relative mass 20": normal H_2O has mass 18, so the extra 2 units are extra neutrons (two deuterium atoms, ^2H , each with one neutron). It has 10 protons, 10 neutrons and 10 electrons.
- **Work backwards**: "a $3+$ ion has 10 electrons and 14 neutrons" means $Z = 10 + 3 = 13$ and $A = 13 + 14 = 27$.
- **"Same number of electrons as"**: work out each electron count; species with equal counts are isoelectronic.
- **Tables to complete** in Paper 2 may also ask for the nucleon number or the number of occupied shells (from the configuration).

2. Write or choose an electronic configuration (26 questions)

1. Find the number of **electrons** (not protons) with type 1.
2. Fill in energy order: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$, stopping when the electrons run out.
3. Apply the exceptions: Cr and Cu take $4s^1$. For a positive ion of Sc to Zn, remove **4s first**, then 3d.
4. Write it in the form asked: **full** ($1s^2 2s^2 \dots$, never [Ar]) or **shorthand** ([Ar] $3d^6 4s^2$).

Variations you will meet:

- **Electrons in boxes**: one box per orbital (three for p, five for d), single arrows in separate boxes before any pairing, paired arrows opposite.
- **"Identify what [Ar] means"**: write $1s^2 2s^2 2p^6 3s^2 3p^6$.
- **An ion with the same configuration as a noble gas**: for example X^+ like Ar has 18 electrons, so X has 19, which is potassium.
- **Configuration to position**: the outer shell n is the period; outer s and p electrons give the group ($4s^2 4p^3$ is Group 15, Period 4).

3. Use successive ionisation energies (25 questions)

1. Work out the **ratio (or difference) between each IE and the next** and find the **biggest jump**.
2. The number of electrons removed **before** the jump = the number of **outer-shell electrons**.
3. Turn that into the group (1, 2, then 13–18 for 3–8 outer electrons) and, with the period, the element and its configuration.
4. If asked to explain: after the jump the electron comes from a shell **closer to the nucleus, less shielded**, so much more strongly attracted.

Variations you will meet:

- **Data that start part-way** (for example the 5th to 10th IEs): the jump after the k th value in the list still tells you the total removed before it. A jump between the 6th and 7th means 6 outer electrons, Group 16.
- **Log graphs**: count the points before the first steep rise. For a Period 3 element, the later jumps split the remaining electrons into 8 (second shell) and then 2 (first shell).
- **Sketching the graph**: always rising, gentle within a shell, a big jump between shells, never a dip.
- **Which ionisation has the largest energy**: the one that breaks into a new, inner shell (for example $\text{Cl}^{7+} \rightarrow \text{Cl}^{8+}$, which takes an electron from $n = 2$).
- **Two elements in the same group**: the one with smaller values after the jump is lower in the group (larger atom).

4. Unpaired electrons, electron pairs and orbital facts (24 questions)

1. Write the configuration.
2. Draw boxes for each sub-shell that is not full (full ones are all paired).
3. Count single arrows (unpaired) or double arrows (pairs), in the shell or sub-shell asked for.

For example, pairs of electrons in the $n = 3$ shell of Co ($3s^2 3p^6 3d^7$): $1 + 3 + 2 = 6$ pairs.

Variations you will meet:

- **Number of electrons in a sub-shell type**: "electrons in spherical orbitals" means all s electrons (K: $2 + 2 + 2 + 1 = 7$); "electrons in p orbitals" means all p electrons.
- **Unpaired electrons across a period** rise and fall with the p sub-shell: across Period 2, Li 1, Be 0, B 1, C 2, N 3, O 2, F 1, Ne 0.
- **Statements about an orbital**: any orbital holds at most 2 electrons (a p **sub-shell** holds 6); a 3p orbital is a dumbbell; 3d is higher in energy than 3p in the same shell.
- **Identify an element from orbital occupancy**: "in Period 3, two 3p orbitals full and one half-full" means $3p^5$, chlorine.
- **Define ground state**: the lowest-energy arrangement of the electrons.

5. Explain first ionisation energy trends and anomalies (24 questions)

1. Name the **electron being removed** from each species (which shell and sub-shell, paired or not).
2. Decide which **factor** makes the difference: nuclear charge, distance, shielding, or spin-pair repulsion.
3. Say which factor **outweighs** which, and link it to the **attraction between the nucleus and the outer electron**.

Model answers:

- **Across a period (general rise)**: nuclear charge increases; electron removed from the same shell with similar shielding; so the attraction increases.
- **S lower than P** (or O lower than N): in S the electron is removed from a 3p orbital that holds a pair; spin-pair repulsion makes it easier to remove, which outweighs the increase in nuclear charge.

- **Al lower than Mg** (or B lower than Be): the Al electron is in 3p, higher in energy than 3s (further away and shielded by 3s), which outweighs the extra proton.
- **Down a group**: more shells, so the outer electron is further away and more shielded; this outweighs the greater nuclear charge.

Variations you will meet:

- **Choose a missing value** (for example IE_1 of O, given N and F): it is below N (spin-pair repulsion) and below F (F has more protons in the same shell).
- **Second IEs**: the 2nd IE of K is far larger than that of Ca because K's second electron comes from 3p, in a shell closer to the nucleus and less shielded, while Ca's comes from 4s.
- **Comparing different species**: IE_2 of Mg and IE_1 of Na both remove a 3s electron, but Mg^+ has 12 protons against Na's 11 with the same shielding, so $IE_2(Mg)$ is larger.
- **Isotopes** (^{63}Cu and ^{65}Cu) have the **same** IE: same nuclear charge, same configuration and shielding.
- **Which factor matters most** when the data jump (for example IE_2 of Na much higher than Mg's): the distance from the nucleus / which shell the electron is in.

6. Write an equation for (or define) an ionisation energy (16 questions)

1. Start with **one** gaseous atom or ion of the element, never X_2 : $Cl(g)$, not $Cl_2(g)$.
2. Put **(g)** on both the atom or ion and the product ion.
3. Show **one electron** leaving: $Ca(g) \rightarrow Ca^+(g) + e^-$.
4. For the n th IE the charges go from $(n - 1) +$ to $n +$: third IE of Al is $Al^{2+}(g) \rightarrow Al^{3+}(g) + e^-$.

Definition of first IE: the energy needed to remove one electron **from each atom in one mole of gaseous atoms** to form one mole of gaseous $1+$ ions.

Variation: combining equations, for example $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$ is the sum of the first and second IEs.

7. Explain or order atomic and ionic radii (15 questions)

1. **Count the electrons and shells** in each species.
2. **Different numbers of shells**: more shells, bigger.
3. **Same electrons (isoelectronic)**: more protons, greater nuclear attraction for the same electrons, smaller.
4. State the order and give the reason in those words.

For Cl^- , Ar, K^+ : all have 18 electrons with 17, 18, 19 protons, so the order of increasing radius is $K^+ < Ar < Cl^-$.

Variations you will meet:

- **Atom to positive ion** (Co to Co^{2+}): smaller, because electrons are lost (sometimes a whole shell) and the remaining electrons are attracted more strongly.

- **Anion versus cation with the same electrons:** S^{2-} and K^+ both have 18 electrons, but K^+ has 19 protons against 16, so K^+ is smaller and S^{2-} larger.
- **Down Group 2:** the $2+$ ions get bigger because each has one more shell.
- **Across Period 3:** atoms get smaller (nuclear charge up, same shell).

8. Relative atomic mass from isotope data (12 questions)

This belongs mainly to the topic "Atoms, molecules and stoichiometry" but appears in the same Paper 2 questions as isotopes.

1. **Define A_r :** the weighted average mass of the atoms of an element compared with $\frac{1}{12}$ of the mass of one atom of carbon-12. **Relative isotopic mass:** the mass of one atom of an isotope compared with $\frac{1}{12}$ of the mass of a carbon-12 atom.
2. **Calculate A_r :** multiply each relative isotopic mass by its abundance, add, and divide by the total abundance (100 for percentages). Give the answer to the decimal places asked.
3. **A missing mass or abundance:** write the same equation with the unknown, and solve. If one abundance is missing, it is 100 minus the others.
4. **"What is needed to calculate A_r ":** the relative isotopic mass **and** the relative abundance of each isotope.

Variation: if A_r is below the mass of the most common isotopes, the missing isotope must be lighter than A_r .

9. Define isotopes; explain their properties (8 questions)

1. **Definition:** atoms of the same element with the same number of protons and different numbers of neutrons.
2. **Same chemical properties:** same number and arrangement of electrons.
3. **Different physical properties:** different mass (and so density), because of the different numbers of neutrons.
4. **How two isotopes differ:** say the numbers, for example "copper-65 has two more neutrons than copper-63".

10. Sketch or recognise s and p orbitals (6 questions)

1. Identify the orbital first if the question describes it: "the highest-energy orbital containing electrons in Mg" is 3s; "the lowest-energy orbital in $n = 3$ " is 3s.
2. **s:** draw a circle (a sphere) with the nucleus at the centre.
3. **p:** draw a dumbbell, two lobes either side of the nucleus along one axis. Don't draw a figure-of-eight with the lobes apart from the nucleus.

11. Particle beams in an electric field; where the mass is (2 questions)

1. **Neutrons:** a straight line.
2. **Protons:** curve towards the negative plate.

3. **Electrons**: curve towards the positive plate, **clearly more** than the protons. Label each beam.

4. **Other ions**: compare charge ÷ mass with a proton's; a smaller ratio bends less.

Distribution of mass: almost all the mass is in the small, dense nucleus (protons and neutrons, each of relative mass 1); the electrons, around the nucleus, have almost no mass.

12. Identify free radicals (2 questions)

1. Count the electrons in each species (with the charge).
2. Write the configuration and look for an unpaired electron. Any species with an odd number of electrons is a radical; some with even numbers (the O atom, with two unpaired 2p electrons) are too.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

Chromium has several isotopes, including chromium-50 and chromium-53.

- (a) Give the numbers of protons, neutrons and electrons in one $^{53}\text{Cr}^{3+}$ ion. **[2]**
- (b) Give the full electronic configuration of a Cr^{3+} ion and state how many unpaired electrons it has. **[2]**
- (c) State how an atom of chromium-50 differs from an atom of chromium-53, and explain why the two isotopes have the same chemical properties. **[2]**
- (d) Deduce the number of pairs of electrons in the $n = 3$ shell of a chromium atom. **[1]**
- (a)** Protons 24; neutrons $53 - 24 = 29$; electrons $24 - 3 = 21$. **B2** (all three for 2 marks, two for 1)
- (b)** Cr is $[\text{Ar}] 3d^5 4s^1$. Remove the 4s electron first, then two 3d: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$. **B1** Three d electrons in separate orbitals: **3** unpaired. **B1**
- (c)** Chromium-50 has 26 neutrons, three fewer than chromium-53. **B1** Both have 24 electrons in the same arrangement, and chemical reactions depend on the electrons. **B1**
- (d)** $3s^2$: 1 pair; $3p^6$: 3 pairs; $3d^5$: all five single. Total **4** pairs. **B1**

Example 2

- (a) Define first ionisation energy. **[2]**
- (b) Write an equation, with state symbols, for the second ionisation energy of oxygen. **[1]**
- (c) Explain why the first ionisation energy of oxygen is lower than that of nitrogen. **[2]**
- (d) Explain why the first ionisation energy of boron is lower than that of beryllium. **[2]**
- (a)** The energy needed to remove one electron from each atom in one mole of gaseous atoms **M1**, to form one mole of gaseous $1+$ ions **M2**.
- (b)** $\text{O}^+(\text{g}) \rightarrow \text{O}^{2+}(\text{g}) + e^-$ **B1**
- (c)** Oxygen ($2p^4$) has two electrons paired in one 2p orbital; they repel each other (spin-pair repulsion), so one is easier to remove. **M1** This outweighs oxygen's greater nuclear charge. **M2**
- (d)** Boron's outer electron is in 2p, which is higher in energy (further from the nucleus) than beryllium's 2s **M1**, and it is shielded by the 2s electrons, so it is less strongly attracted despite the extra proton. **M2**

Example 3

The first eight ionisation energies of element Q, in Period 3, are shown, in kJ mol^{-1} .

1st	2nd	3rd	4th	5th	6th	7th	8th
786	1580	3230	4360	16 100	19 800	23 800	29 300

- (a) Deduce the group of Q. Explain your answer. **[2]**
- (b) Identify Q and give the full electronic configuration of the Q^{2+} ion. **[2]**
- (c) Explain why each ionisation energy is larger than the one before. **[2]**
- (a)** The biggest jump is between the 4th and 5th ($16\,100 \div 4360 = 3.7$ times), so Q has **4** outer electrons: **Group 14. M1** The 5th electron comes from a shell closer to the nucleus, less shielded, so it is much more strongly attracted. **M2**
- (b)** Group 14, Period 3: **silicon. B1** Si is $1s^2 2s^2 2p^6 3s^2 3p^2$; Si^{2+} loses the two 3p electrons: $1s^2 2s^2 2p^6 3s^2$. **B1**
- (c)** Each electron is removed from an increasingly positive ion: the same number of protons attracts fewer electrons **M1**, so the ion is smaller and the next electron is held more strongly. **M1**

Example 4

Boron has two isotopes: boron-10 (relative isotopic mass 10.01, abundance 19.9%) and boron-11 (relative isotopic mass 11.01, abundance 80.1%).

- (a) Calculate the relative atomic mass of boron to two decimal places. [2]
- (b) A meteorite sample of boron has $A_r = 10.85$. Calculate the percentage of boron-11 in it. [2]
- (c) The species N^{3-} , O^{2-} , Mg^{2+} and Al^{3+} all have the same electronic configuration. Arrange them in order of increasing radius and explain your order. [3]

$$(a) A_r = \frac{19.9 \times 10.01 + 80.1 \times 11.01}{100} \quad \mathbf{M1 = 10.81 \quad A1}$$

(b) Let the fraction of boron-11 be x : $10.01(1 - x) + 11.01x = 10.85$ **M1**, so $10.01 + x = 10.85$ and $x = 0.84$: **84.0% boron-11. A1**

(c) $\text{Al}^{3+} < \text{Mg}^{2+} < \text{O}^{2-} < \text{N}^{3-}$ **B1**. All have 10 electrons ($1s^2 2s^2 2p^6$) **B1**, but the numbers of protons fall from 13 to 7, so the nuclear attraction for the same electrons gets weaker and the particle gets bigger. **B1**

Common mistakes

- **Writing 4s electrons in transition-metal ions.** Cu^{2+} is $[\text{Ar}] 3d^9$, not $[\text{Ar}] 3d^7 4s^2$. Examiners see ions with 4s electrons still in place every year: remove 4s first.
- **Missing the Cr and Cu exceptions.** $3d^4 4s^2$ for Cr and $3d^9 4s^2$ for Cu are the commonest wrong answers. Learn $3d^5 4s^1$ and $3d^{10} 4s^1$.
- **Starting an IE equation from Cl_2 or H_2** , or leaving out state symbols. The equation starts from one gaseous atom: $\text{Cl}(\text{g}) \rightarrow \text{Cl}^+(\text{g}) + e^-$. Some candidates also write ionisation as gaining an electron; it is always a loss.
- **"Removing an electron from one mole of atoms"** in the definition. The electron is removed from **each** atom in one mole of gaseous atoms.
- **Half an anomaly answer.** For S < P, candidates often name spin-pair repulsion but forget to say it **outweighs the increased nuclear charge**. Mention both sides.
- **Using spin-pair repulsion where it doesn't apply**, for example for Al < Mg. That is a sub-shell effect (3p is higher in energy than 3s), not pairing.
- **Mentioning "distance" without saying what it does.** Say the electron is further from the nucleus **so it is less strongly attracted**.
- **Vague words about radius.** Say "shells" and "nuclear charge" or "number of protons"; don't use "orbitals" or "sub-shells" to explain why a Group 2 ion is bigger. "Increase in overall nuclear charge" without saying the electrons are the same is not enough.
- **Seeing the first IE as the big jump.** The jump that counts is the biggest ratio, which may come much later. And never draw a dip in a successive IE graph.
- **Beams without a clear difference.** Draw the electron beam bending more than the proton beam, and label all three.
- **Using whole-number masses** (like 63 and 65) when the question gives relative isotopic masses to decimals: use the data given.
- **Unclear handwriting of sub-shells:** a 4s that looks like a 9s loses the mark. Write superscripts clearly.

How to get the marks

- **Configurations** are marked all-or-nothing: every sub-shell and superscript must be right. "Full" means no [Ar].
- **Particle tables**: one mark usually needs every entry in a row, so check each number twice.
- **Ionisation energy equations**: one mark for any correct ionisation, one for exactly the right species with (g) on both sides.
- **"Explain" a trend in IE**: name the electron (its shell or sub-shell), name the factor, and say how it changes the **attraction of the nucleus for the outer electron**. For an anomaly, add what it **outweighs**.
- **"Deduce the group"**: give the group and point to where the big jump is in the data ("between the 6th and 7th").
- **"Explain your order"** for radii: isoelectronic (or number of shells), number of protons, and attraction: three separate ideas, one mark each.
- **"State"** needs just the fact; **"explain"** needs the reason. **"Suggest"** means use your knowledge in a new situation, such as an unfamiliar ion.
- **Draw** orbitals with the nucleus at the centre; label the s sphere or p dumbbell if unclear.
- **Calculations** (A_r): show the multiplied terms and the division, and give the answer to the decimal places asked (usually 1 or 2).

Quick check

1. Give the numbers of protons, neutrons and electrons in $^{34}\text{S}^{2-}$.
2. Give the full electronic configuration of Fe^{3+} and state the number of unpaired electrons.
3. Which of these are free radicals: H, He, O, Ne, Na^+ ?
4. The first ionisation energies of Na, Mg and Al are 494, 577 and 736 kJ mol^{-1} , in some order. Match them, and explain why Al's is lower than Mg's.
5. A beam of $^2\text{H}^+$ ions and a beam of protons, at the same speed, pass through the same electric field. Compare their paths.

Answers

1. Protons 16; neutrons $34 - 16 = 18$; electrons $16 + 2 = 18$.
2. Fe is $[\text{Ar}] 3d^6 4s^2$; lose both 4s and one 3d: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$. Five d electrons, one per orbital: 5 unpaired.
3. H ($1s^1$) and O ($2p^4$, two unpaired) are free radicals. He, Ne and Na^+ ($1s^2 2s^2 2p^6$) have every electron paired.
4. Na 494, Al 577, Mg 736. Al's outer electron is in 3p, higher in energy and shielded by the 3s electrons, so it is removed more easily despite the extra proton.
5. Both bend towards the negative plate. ${}^2\text{H}^+$ has the same charge but twice the mass, so its charge $\div$ mass is half a proton's and it bends less (about half as much).

Past questions to try

- **9701/22/F/M/22 Q1**: an IE equation, the Period 2 trend and its anomaly, then a configuration from successive IEs.
- **9701/23/O/N/23 Q1**: unpaired electrons, choosing the missing IE_1 of sulfur, and the radii of P^{3-} , S^{2-} and Cl^- (parts (a) and (b)).
- **9701/22/F/M/25 Q2**: successive IEs of beryllium, mass in the atom and beams in an electric field (parts (b) and (c)).
- **9701/12/M/J/24 Q4**: spotting free radicals among atoms and ions.