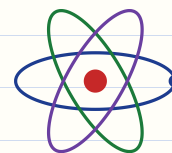


Atomic Structure



Syllabus topic 1: 1.1 Particles in the atom and atomic radius · 1.2 Isotopes · 1.3 Electrons, energy levels and atomic orbitals · 1.4 Ionisation energy

For the Cambridge International AS & A Level Chemistry 9701 syllabus, exams in 2025, 2026 and 2027 (AS Level content)

Syllabus statements covered in this chapter

Syllabus ref.	What you must be able to do	Level
1.1 statements 1–7	The nuclear atom; relative charges and masses of protons, neutrons and electrons; proton (atomic) number and nucleon (mass) number; distribution of mass and charge; beams of particles in an electric field; numbers of protons, neutrons and electrons in atoms and ions; trends in atomic and ionic radius across a period and down a group	AS
1.2 statements 1–4	Definition of isotope; the notation ${}^x_y\text{A}$; why isotopes have the same chemical properties but different mass and density	AS
1.3 statements 1–9	Shells, sub-shells, orbitals, principal quantum number, ground state; s, p and d sub-shells; order of sub-shell energies up to 4p; full and shorthand electronic configurations of atoms and ions (H to Kr); explaining configurations by energy and electron repulsion; electrons-in-boxes; shapes of s and p orbitals; free radicals	AS
1.4 statements 1–8	First ionisation energy and equations for successive IEs; trends across a period and down a group; successive IEs; the factors (nuclear charge, radius, shielding, spin-pair repulsion); deducing configuration and Periodic Table position from IE data	AS

What is inside

1. Concept notes, 8 worked examples
2. Practical skills (Paper 3); common mistakes from examiner reports
3. 20 MCQs (Paper 1 style), explained
4. 3 structured questions (Paper 2 style) with full mark schemes
5. Quick-reference sheet, "I can" checklist

How to read these notes

AS = AS Level content (Papers 1, 2 and 3)

A = A Level only (Papers 4 and 5)

red = key term green = definition

1 = one mark in the mark scheme

★ **Exam tip:** Only the elements hydrogen to krypton are assessed in 1.3 and 1.4, and every atom or ion is in its ground state.

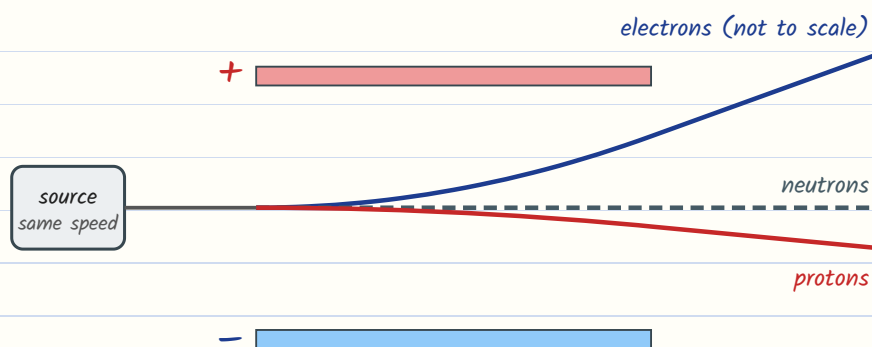
1.1 Particles in the atom and atomic radius

- An atom is **mostly empty space**. At its centre is a very small, very dense **nucleus** containing **protons** and **neutrons** (together called **nucleons**). **Electrons** occupy **shells** in the space around the nucleus. The nucleus is about 10^{-15} m across; the atom is about 10^{-10} m across, roughly 100 000 times larger.

Particle	Relative charge	Relative mass	Where it is
proton	+1	1	nucleus
neutron	0	1	nucleus
electron	-1	$\frac{1}{1836}$ (about 0.0005)	shells around the nucleus

- Proton number** (atomic number) Z = the number of protons in the nucleus of an atom. It identifies the element.
- Nucleon number** (mass number) A = the total number of protons and neutrons in the nucleus. So the number of neutrons = $A - Z$.
- Distribution of mass and charge:** almost all the mass of the atom is in the nucleus (protons and neutrons). All the positive charge is in the nucleus; the negative charge is carried by the electrons spread through the much larger volume around it. A neutral atom has equal numbers of protons and electrons.
- Ions:** a positive ion has **fewer** electrons than protons; a negative ion has **more**. Number of electrons = $Z - (\text{charge})$, e.g. for a $2+$ ion subtract 2, for a $3-$ ion add 3.

Beams of particles in an electric field



electrons bend towards the + plate; protons bend (much less) towards the - plate

Beams entering at the same velocity: the deflection depends on charge $\div$ mass.

- Protons** (positive) are deflected **towards the negative plate**; **electrons** (negative) **towards the positive plate**; **neutrons** (no charge) pass straight through.
- At the same velocity the angle of deflection is proportional to $\frac{\text{charge}}{\text{mass}}$. An electron has the same size of charge as a proton but about $\frac{1}{1836}$ of its mass, so it is deflected **far more**, in the opposite direction.

Worked example 1: protons, neutrons and electrons in atoms and ions

Species	Z	A	protons	neutrons	electrons
${}^{56}_{26}\text{Fe}^{3+}$	26	56	26	$56 - 26 = 30$	$26 - 3 = 23$
${}^{32}_{16}\text{S}^{2-}$	16	32	16	16	$16 + 2 = 18$
${}^{27}_{13}\text{Al}^{3+}$	13	27	13	14	10

The charge changes only the number of electrons; the nucleus (protons and neutrons) is unchanged.

Worked example 2: comparing deflections

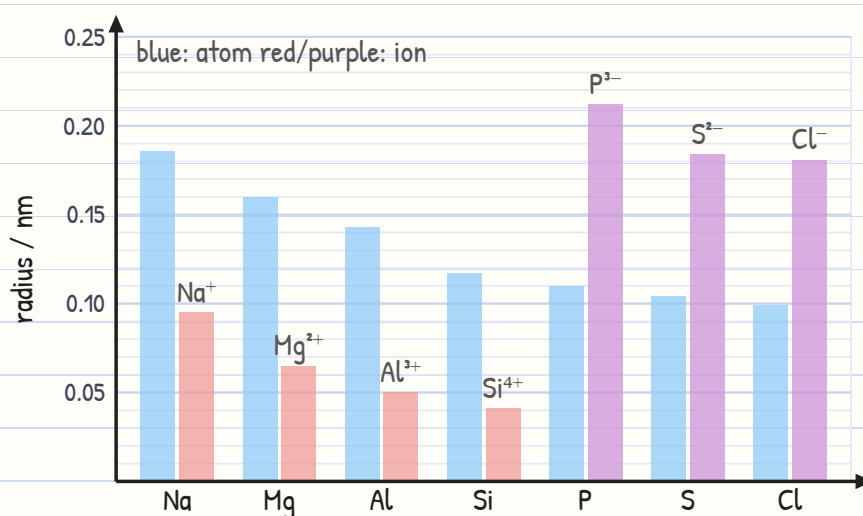
Beams of ${}^1\text{H}^+$, ${}^2\text{H}^+$, ${}^3\text{He}^+$ and ${}^4\text{He}^{2+}$ ions enter the same electric field at the same velocity. Put them in order of decreasing deflection.

charge $\div$ mass: ${}^1\text{H}^+ \frac{1}{1} = 1$; ${}^2\text{H}^+ \frac{1}{2} = 0.50$; ${}^4\text{He}^{2+} \frac{2}{4} = 0.50$; ${}^3\text{He}^+ \frac{1}{3} = 0.33$.

Order: ${}^1\text{H}^+ > {}^3\text{He}^+ = {}^4\text{He}^{2+} > {}^2\text{H}^+$. All are deflected towards the negative plate.

Atomic radius and ionic radius

- **Across a period** (left to right) atomic radius **decreases**: the nuclear charge increases, the outer electrons are in the same shell and shielding by the inner shells stays about the same, so the outer electrons are pulled closer.
- **Down a group** atomic radius **increases**: each element has one more shell, so the outer electrons are further from the nucleus and more shielded; this outweighs the larger nuclear charge.
- **Positive ions** are **smaller** than their atoms: the outer shell is lost and the remaining electrons are pulled in by the same nuclear charge. **Negative ions** are **larger**: extra electrons in the same shell repel each other.



Period 3 radii (Data Booklet values).

- **Across Period 3 ions**: $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$ – all have 10 electrons (2,8) but the nuclear charge rises from 11+ to 14+. $\text{P}^{3-} > \text{S}^{2-} > \text{Cl}^-$ – all have 18 electrons (2,8,8), nuclear charge rises from 15+ to 17+. The anions are much larger than the cations because they have **one more shell**.
- **Down a group** ionic radius increases (e.g. Mg^{2+} 0.065 nm, Ca^{2+} 0.099 nm; Cl^- 0.181 nm, Br^- 0.195 nm) for the same reason as atoms: one more shell each time.

1.2 Isotopes

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

- Notation ${}^x_y\text{A}$: x = nucleon (mass) number, y = proton (atomic) number. E.g. ${}^{35}_{17}\text{Cl}$ (17 p, 18 n) and ${}^{37}_{17}\text{Cl}$ (17 p, 20 n).
- **Same chemical properties:** isotopes have the **same number of electrons and the same electronic configuration**; chemical reactions involve the electrons (especially the outer ones), so the isotopes react in the same way.
- **Different physical properties (mass and density):** different numbers of neutrons give atoms of **different mass**. The atoms are almost the same size, so the substance made from the heavier isotope has a **higher density**.

Worked example 3: density of heavy water

Heavy water, D_2O , contains deuterium, ${}^2_1\text{H}$. Ordinary water has density 1.00 g cm^{-3} . Estimate the density of D_2O , assuming one mole of each liquid occupies the same volume. (A_r : H = 1.0, D = 2.0, O = 16.0)

$M_r(\text{H}_2\text{O}) = 18.0$, $M_r(\text{D}_2\text{O}) = 20.0$. Same volume per mole, so density $\propto$ mass per mole: $1.00 \times \frac{20.0}{18.0} = 1.11 \text{ g cm}^{-3}$. The chemistry is the same: D_2O still reacts with sodium to give hydrogen (as D_2).

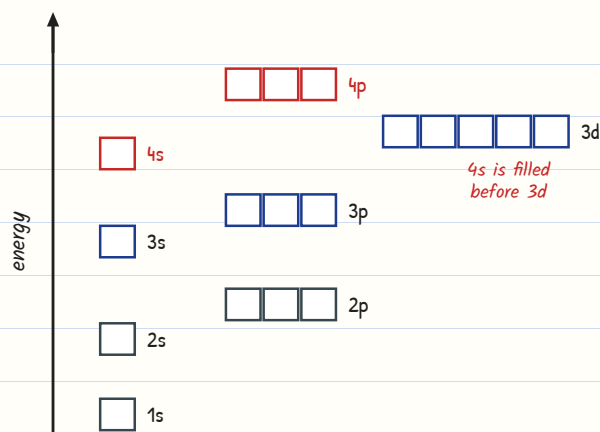
1.3 Electrons, energy levels and atomic orbitals

- **Shells** (energy levels) are numbered by the **principal quantum number**, $n = 1, 2, 3, 4 \dots$. The larger n , the higher the energy and the further the electrons are (on average) from the nucleus.
- Each shell is divided into **sub-shells** (s, p, d). Each sub-shell is made of **orbitals**. An **orbital** is a region of space around the nucleus that can hold up to two electrons, with opposite spins.
- The **ground state** is the lowest-energy arrangement of the electrons (the electronic configuration you write in an exam).

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

Shell	Sub-shells	Max electrons
$n = 1$	1s	2
$n = 2$	2s, 2p	8
$n = 3$	3s, 3p, 3d	18
$n = 4$	4s, 4p, 4d, 4f	32

Order of increasing energy: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$.



Sub-shell energies: 4s lies just below 3d.

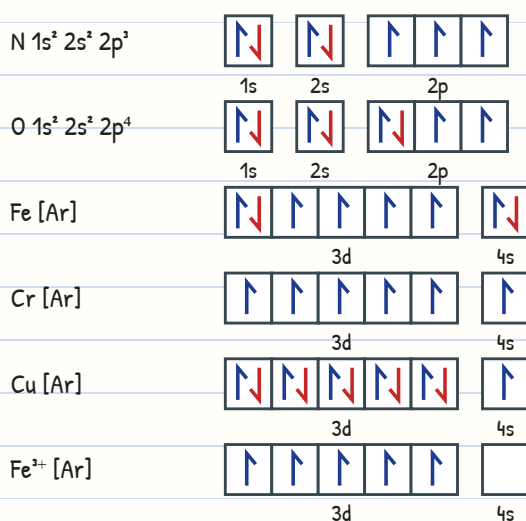
Rules for filling (and why)

- **Lowest energy first:** electrons go into the lowest-energy sub-shell available, so 4s fills before 3d (K and Ca are [Ar] 4s¹ and [Ar] 4s²).
- **Singly before pairing:** within a sub-shell, electrons occupy orbitals **singly with parallel spins** before any orbital takes a second electron, because two electrons in the same orbital repel each other (**inter-electron, spin-pair repulsion**). Nitrogen is 2p: ↑ ↑ ↑, not ↑ ↓ ↑.
- **Two exceptions to learn:** chromium is [Ar] 3d⁵ 4s¹ and copper is [Ar] 3d¹⁰ 4s¹ (a half-filled or full 3d sub-shell with one 4s electron has lower energy).
- **Ions of d-block elements lose their 4s electrons first:** once 3d is occupied, 4s is the outer, higher-energy sub-shell. Fe = [Ar] 3d⁶ 4s², Fe²⁺ = [Ar] 3d⁶, Fe³⁺ = [Ar] 3d⁵.

Two ways to write a configuration (both accepted): full, e.g. Fe 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁶ 4s²; shorthand, using the previous noble gas, e.g. Fe [Ar] 3d⁶ 4s². Write 3d before 4s (shell order), even though 4s fills first.

Electrons-in-boxes notation

Each box is one orbital; arrows show electrons and their spins. Boxes of one sub-shell are drawn joined.

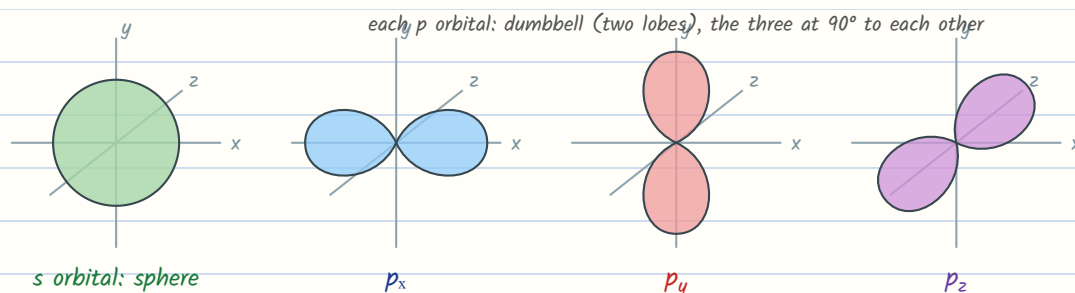


N has three unpaired 2p electrons; O has one pair in a 2p orbital. Fe has 4 unpaired electrons, Fe³⁺ has 5.

Worked example 4: configurations of atoms and ions

Species	Electrons	Full configuration	Shorthand
V ($Z = 23$)	23	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$	$[\text{Ar}] 3d^3 4s^2$
Fe^{2+} ($Z = 26$)	24	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$	$[\text{Ar}] 3d^6$
Cu ($Z = 29$)	29	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$	$[\text{Ar}] 3d^{10} 4s^1$
Br^- ($Z = 35$)	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$	$[\text{Ar}] 3d^{10} 4s^2 4p^6$

Fe^{2+} : remove the two 4s electrons, not two 3d electrons. Br^- has the same configuration as krypton.

Shapes of s and p orbitals

- An **s orbital is spherical** (centred on the nucleus); 2s is larger than 1s.
- A **p orbital is dumbbell-shaped**: two lobes on opposite sides of the nucleus. The three p orbitals of a sub-shell (p_x , p_y , p_z) point along three axes at 90° to each other and have equal energy.

Free radicals

- A **free radical** is a species with one or more unpaired electrons. Examples: a chlorine atom $\text{Cl}\cdot$ ($3p: \uparrow\downarrow \uparrow\downarrow \uparrow$), the methyl radical $\cdot\text{CH}_3$, nitrogen monoxide NO and nitrogen dioxide NO_2 (odd numbers of electrons). Cl^- and Na^+ are not radicals: all their electrons are paired.

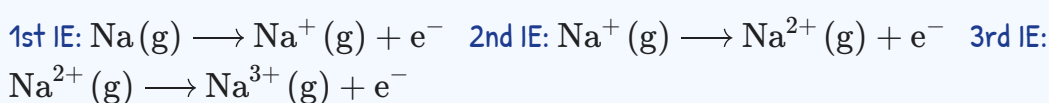
Worked example 5: counting unpaired electrons

How many unpaired electrons are in (a) a phosphorus atom, (b) Cr, (c) Cu^{2+} ? Which are free radicals?

(a) P: $[\text{Ne}] 3s^2 3p^3$; three 3p electrons go singly into three orbitals: **3**. (b) Cr: $[\text{Ar}] 3d^5 4s^1$: **6**. (c) Cu^{2+} : $[\text{Ar}] 3d^9$ (lose 4s¹ then one 3d): four pairs and **1** unpaired. All three have unpaired electrons, so all fit the definition of a free radical (in practice the term is used mostly for reactive atoms and molecular fragments such as $\text{Cl}\cdot$ and $\cdot\text{CH}_3$).

1.4 Ionisation energy

Definition The **first ionisation energy** of an element is the energy needed to remove one electron from each atom in one mole of gaseous atoms of the element, to form one mole of gaseous $1+$ ions.

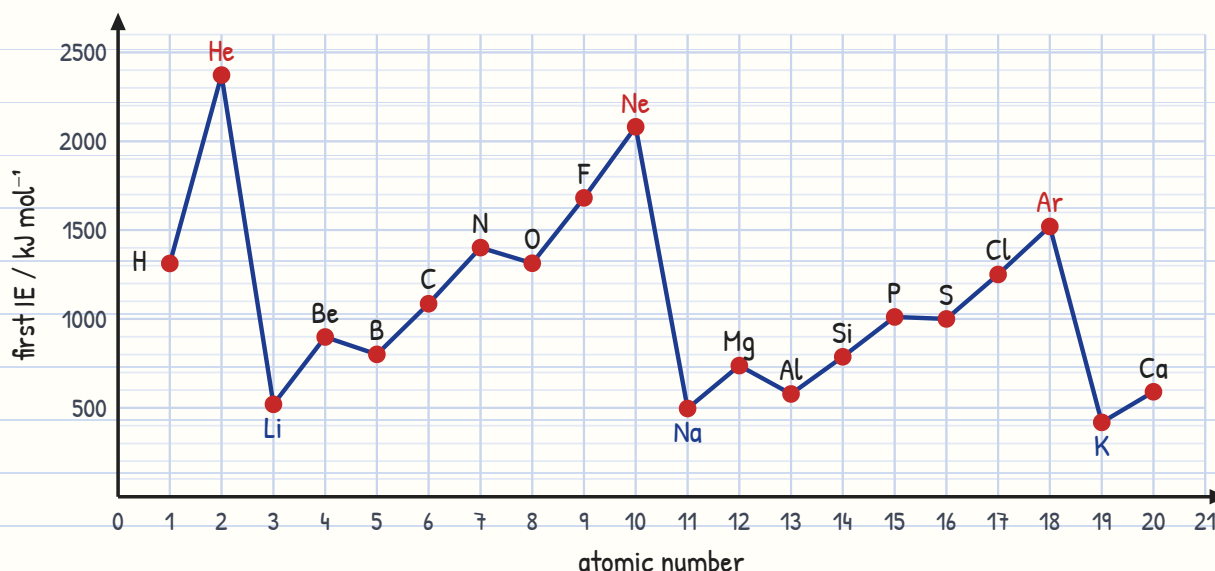


Each **successive ionisation energy** removes one more electron from the gaseous ion. Always one electron per equation, always (g) on every species. IE values are always positive (energy is taken in).

Four factors that decide the size of an ionisation energy

Ionisation energy measures the **attraction between the nucleus and the electron being removed**. The attraction is larger, and the IE higher, when:

1. the **nuclear charge** (number of protons) is larger;
2. the electron is **closer to the nucleus** (smaller atomic/ionic radius, lower shell);
3. there is **less shielding** by electrons in inner shells and inner sub-shells (they repel the outer electron);
4. the electron is **not in a paired orbital**: spin-pair repulsion makes a paired electron easier to remove.



First ionisation energies, hydrogen to calcium: peaks at noble gases (red), troughs at Group 1 (blue).

Trend across a period (e.g. Na to Ar)

- **General increase**: nuclear charge increases; the electron is removed from the same shell; shielding by inner shells is similar; the atomic radius decreases. So the outer electron is held more strongly.
- **Fall from Group 2 to Group 13** (Be → B, Mg → Al): the electron removed from Al is in the **3p** sub-shell, which is higher in energy than 3s, slightly further from the nucleus and partly **shielded by the 3s electrons**.
- **Fall from Group 15 to Group 16** (N → O, P → S): sulfur's electron is removed from a **3p orbital that already contains two electrons**; **spin-pair repulsion** makes it easier to remove. Phosphorus has three unpaired 3p electrons.

Trend down a group (e.g. Group 1 or Group 2)

- **Decrease**: the outer electron is in a shell further from the nucleus, with more inner shells shielding it. These effects **outweigh** the increase in nuclear charge.

Worked example 6: a full-mark explanation

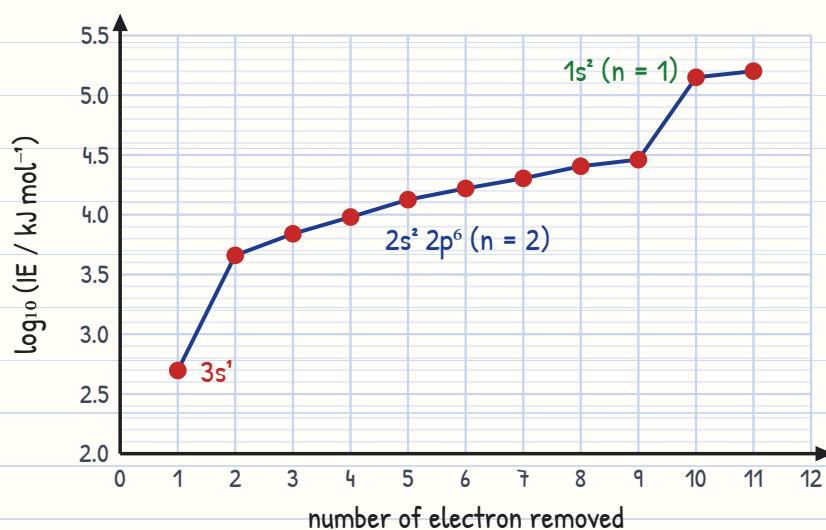
Explain why the first ionisation energy of aluminium (578 kJ mol⁻¹) is lower than that of magnesium (738 kJ mol⁻¹).

Mg: [Ne] 3s²; Al: [Ne] 3s² 3p¹. The electron removed from Al is in the **3p sub-shell**, which is **higher in energy** (further from the nucleus) than 3s, and it is **shielded by the 3s electrons**. This outweighs the greater nuclear charge of Al, so less energy is needed.

Successive ionisation energies

- Successive IEs **always increase**: each electron is removed from an increasingly positive ion, with less electron-electron repulsion and a smaller radius.
- A **big jump** occurs when the next electron comes from a **new shell closer to the nucleus** (much less shielded).

- Count the electrons removed **before the first big jump**: that is the number of outer electrons, so it gives the **group**. The number of jumps tells you how many shells are occupied.
- Smaller jumps inside a shell show sub-shells (e.g. Na 9th $\rightarrow$ 10th is $n = 2 \rightarrow n = 1$; within $n = 2$, 2p before 2s).



Sodium: 1 + 8 + 2 electrons, so Group 1 and three shells.

Worked example 7: group from successive IE data

The first six IEs of element X (in kJ mol^{-1}) are 787, 1577, 3232, 4356, 16 091, 19 805. Deduce the group of X and its outer configuration, given that X is in Period 3.

Ratios of each IE to the one before: 2.00, 2.05, 1.35, **3.69**, 1.23. The big jump is between the 4th and 5th IE, so X has **4 outer electrons**: Group 14. In Period 3 this is **silicon**, $[\text{Ne}] 3s^2 3p^2$. (The 1st $\rightarrow$ 2nd and 2nd $\rightarrow$ 3rd rises are larger than 3rd $\rightarrow$ 4th: the first two electrons come from 3p, the next two from 3s.)

Worked example 8: energy to form a 3+ ion

The first three IEs of aluminium are 578, 1817 and 2745 kJ mol^{-1} . Write the equation for the third IE and calculate the energy needed for $\text{Al}(\text{g}) \rightarrow \text{Al}^{3+}(\text{g}) + 3\text{e}^-$.

3rd IE: $\text{Al}^{2+}(\text{g}) \rightarrow \text{Al}^{3+}(\text{g}) + \text{e}^-$. Total energy = 578 + 1817 + 2745 = **5140 kJ mol^{-1}** (sum of the first three IEs; the third IE alone is not the energy to form Al^{3+} from Al).

Practical skills for Paper 3

Handling and plotting data

- Topic 1 has no laboratory experiment of its own, but Paper 3 marks the same data skills you use here: drawing a results table with each column headed by quantity and unit (e.g. first IE / kJ mol^{-1}), plotting points accurately, and choosing sensible scales.
- Choose scales that use more than half of the grid in each direction, and avoid awkward scales (e.g. one square = 0.45 units): the June 2023 examiners noted that awkward scales almost always cause plotting errors.
- When values cover a huge range (successive IEs run from hundreds to hundreds of thousands of kJ mol^{-1}), plot $\log_{10}$ IE so that every jump can be seen.
- Give calculated values to the number of significant figures justified by the data (usually 3), and do not round early in a multi-step calculation.

- Use the Periodic Table and data section printed in the paper: they give proton numbers and relative atomic masses.

Common mistakes that lose marks

- **Nucleon number is not neutron number:** the nucleon (mass) number counts protons + neutrons. Examiners also report confusion of "nucleon" with "nuclear", and "isotope" with "isomer".
- **Missing state symbols** in ionisation energy equations: every species is (g), including the ion, e.g.
 $\text{Te(g)} \longrightarrow \text{Te}^+(\text{g}) + \text{e}^-$.
- **Wrong orbital labels:** write the full label (1s, not just "s"), and stay within the elements H to Kr (4p is the highest sub-shell needed).
- **Transition-metal ions keeping 4s electrons:** Cu^{2+} is $[\text{Ar}] 3\text{d}^9$, not $[\text{Ar}] 3\text{d}^7 4\text{s}^2$; remove 4s electrons first.
- **Contradictory explanations:** for a decrease in IE down a group, giving "increase in nuclear charge" as the reason works against your own answer; say that the larger distance and extra shielding **outweigh** the extra nuclear charge.
- **Graphs of successive IEs with a dip:** successive IEs always increase; a big jump is a steep rise, never a fall.
- **Forgetting spin-pair repulsion** when explaining the fall from Group 15 to Group 16.
- **Definition of first IE without "gaseous" or "one mole":** both are needed for the mark.

Exam-style multiple-choice questions (Paper 1 style)

20 questions, 1 mark each. Choose one answer, A, B, C or D. Answers with explanations follow the questions.

1. Which statement about the nucleus of an atom is correct?

- A It contains almost all of the mass of the atom but only a tiny fraction of its volume.
- B It contains all the protons, neutrons and electrons of the atom.
- C It occupies about half of the volume of the atom.
- D It carries no overall charge.

2. A species has 16 protons, 18 neutrons and 18 electrons. What is it?

- | | |
|-------------------------------|-------------------------------|
| A ${}^{34}_{18}\text{Ar}$ | B ${}^{36}_{16}\text{S}^{2-}$ |
| C ${}^{34}_{16}\text{S}^{2-}$ | D ${}^{34}_{17}\text{Cl}^-$ |

3. Beams of four ions enter the same electric field at the same velocity. Which ion is deflected least?

- | | |
|--------------------|------------------------|
| A ${}^1\text{H}^+$ | B ${}^3\text{He}^+$ |
| C ${}^2\text{H}^+$ | D ${}^4\text{He}^{2+}$ |

4. Which list shows the atoms in order of increasing atomic radius?

- | | |
|--|--|
| A $\text{Na} < \text{Mg} < \text{S} < \text{Cl}$ | B $\text{Mg} < \text{Na} < \text{Cl} < \text{S}$ |
| C $\text{S} < \text{Cl} < \text{Na} < \text{Mg}$ | D $\text{Cl} < \text{S} < \text{Mg} < \text{Na}$ |

5. Which ion has the largest radius?

- | | |
|-------------------|--------------------|
| A O^{2-} | B N^{3-} |
| C F^- | D Mg^{2+} |

6. Which is the same for an atom of ^{35}Cl and an atom of ^{37}Cl ?
- A the number of electrons in the outer shell
 - B the mass of the atom
 - C the number of neutrons
 - D the nucleon number
7. What is the electronic configuration of the Fe^{3+} ion?
- A $[\text{Ar}] 3\text{d}^3 4\text{s}^2$
 - B $[\text{Ar}] 3\text{d}^6 4\text{s}^2$
 - C $[\text{Ar}] 3\text{d}^4 4\text{s}^1$
 - D $[\text{Ar}] 3\text{d}^5$
8. How many unpaired electrons are there in a phosphorus atom in its ground state?
- A 1
 - B 2
 - C 3
 - D 5
9. Which sub-shell is filled immediately after the 4s sub-shell?
- A 3d
 - B 4p
 - C 4d
 - D 3p
10. What is the maximum number of electrons that the shell with principal quantum number $n = 3$ can hold?
- A 8
 - B 18
 - C 32
 - D 10
11. Which species is a free radical?
- A Cl^-
 - B Na^+
 - C NO_2
 - D H_2O
12. Which equation represents the second ionisation energy of calcium?
- A $\text{Ca}(\text{g}) \longrightarrow \text{Ca}^{2+}(\text{g}) + 2\text{e}^-$
 - B $\text{Ca}(\text{s}) \longrightarrow \text{Ca}^{2+}(\text{s}) + 2\text{e}^-$
 - C $\text{Ca}^+(\text{aq}) \longrightarrow \text{Ca}^{2+}(\text{aq}) + \text{e}^-$
 - D $\text{Ca}^+(\text{g}) \longrightarrow \text{Ca}^{2+}(\text{g}) + \text{e}^-$
13. Which element has the highest first ionisation energy?
- A helium
 - B neon
 - C fluorine
 - D hydrogen
14. Why is the first ionisation energy of aluminium lower than that of magnesium?
- A Aluminium has a smaller nuclear charge.
 - B The electron removed from aluminium is in a 3p sub-shell, of higher energy than 3s and shielded by the 3s electrons.
 - C The electron removed from aluminium is paired.
 - D Aluminium has more shells of electrons.

15. Why is the first ionisation energy of oxygen lower than that of nitrogen?
- A Oxygen has a larger atomic radius.
 - B Oxygen has more shielding by inner shells.
 - C The electron removed from oxygen is in a 2s orbital.
 - D The electron removed from oxygen is in a 2p orbital that contains two electrons, which repel each other.
16. The first five ionisation energies of an element, in kJ mol^{-1} , are 738, 1451, 7733, 10 543 and 13 630. In which group of the Periodic Table is the element?
- A Group 1
 - B Group 13
 - C Group 2
 - D Group 15
17. Why do first ionisation energies decrease down Group 1?
- A The outer electron is further from the nucleus and more shielded, which outweighs the increase in nuclear charge.
 - B The nuclear charge decreases down the group.
 - C Spin-pair repulsion increases down the group.
 - D The outer electron is in a p orbital in the heavier elements.
18. Which atom has the same number of unpaired electrons as an Fe^{3+} ion?
- A Fe
 - B Cr
 - C Cu
 - D Mn
19. Element Y is in Period 3. Its first six ionisation energies, in kJ mol^{-1} , are 1012, 1907, 2914, 4964, 6274 and 21 267. What is Y?
- A sulfur
 - B phosphorus
 - C silicon
 - D chlorine
20. Which statement about atomic orbitals is correct?
- A An s orbital is dumbbell-shaped.
 - B A p sub-shell contains six orbitals.
 - C Each orbital can hold up to two electrons, with opposite spins.
 - D A 2p orbital is spherical.

Answers and explanations

1 A · 2 C · 3 B · 4 D · 5 B · 6 A · 7 D · 8 C · 9 A · 10 B · 11 C · 12 D · 13 A · 14 B · 15 D · 16 C · 17 A · 18 D · 19 B · 20 C

1. **A.** The nucleus is about 10^{-5} of the diameter of the atom; electrons are outside it; it is positively charged.
2. **C.** 16 protons = sulfur; $A = 16 + 18 = 34$; 18 electrons = 2- charge.
3. **B.** charge ÷ mass: 1, 0.33, 0.50, 0.50; the smallest ratio gives the smallest deflection.
4. **D.** Radius decreases across Period 3: Na (0.186) > Mg (0.160) > S (0.104) > Cl (0.099 nm).
5. **B.** N^{3-} , O^{2-} , F^{-} and Mg^{2+} all have 10 electrons; N^{3-} has the smallest nuclear charge (7+).
6. **A.** Isotopes have the same electronic configuration; they differ in neutrons, mass and nucleon number.
7. **D.** Fe is $[\text{Ar}] 3d^6 4s^2$; remove both 4s electrons and one 3d electron.
8. **C.** $3p^3$: one electron in each of the three 3p orbitals.

9. **A.** The order is $4s < 3d < 4p$.
10. **B.** $3s(2) + 3p(6) + 3d(10) = 18$, i.e. $2n^2$.
11. **C.** NO_2 has $7 + 16 = 23$ electrons, an odd number, so one is unpaired.
12. **D.** One electron removed from each gaseous $1+$ ion.
13. **A.** Helium (2372 kJ mol^{-1}): $1s$ electrons, very close to the nucleus, no shielding.
14. **B.** Al's nuclear charge is larger ($13+$), so A is wrong; its $3p$ electron is unpaired.
15. **D.** Spin-pair repulsion; O is smaller than N and the shielding is the same.
16. **C.** The big jump is after the 2nd electron ($1451 \rightarrow 7733$): two outer electrons (this is magnesium).
17. **A.** Nuclear charge increases down the group, but distance and shielding win.
18. **D.** $\text{Fe}^{3+} [\text{Ar}] 3d^5$ has 5; $\text{Mn} [\text{Ar}] 3d^5 4s^1$ has 5; Fe has 4, Cr 6, Cu 1.
19. **B.** The big jump is between the 5th and 6th IE ($6274 \rightarrow 21267$): five outer electrons, Group 15.
20. **C.** s orbitals are spherical and p orbitals dumbbell-shaped; a p sub-shell has three orbitals.

Exam-style structured questions (Paper 2 style)

Answer all parts. Show your working. Total: 25 marks.

Question 1 [8 marks]

Magnesium has three naturally occurring isotopes, ^{24}Mg , ^{25}Mg and ^{26}Mg .

- (a) Define the term *isotope*. [1]
- (b) Complete a table to show the numbers of protons, neutrons and electrons in a ^{25}Mg atom and in a $^{26}\text{Mg}^{2+}$ ion. [2]
- (c) Explain why ^{24}Mg and ^{26}Mg have the same chemical properties. [2]
- (d) A sample of pure ^{26}Mg metal is prepared. Suggest how its density compares with that of a sample of pure ^{24}Mg , and explain your answer. [1]
- (e) Beams of $^{24}\text{Mg}^{2+}$ and $^{26}\text{Mg}^{2+}$ ions travelling at the same velocity pass between two charged plates. State which beam is deflected more, and explain why. [2]

Mark scheme

- (a) atoms of the same element / same number of protons, with different numbers of neutrons [1]
- (b) ^{25}Mg : 12 p, 13 n, 12 e [1]; $^{26}\text{Mg}^{2+}$: 12 p, 14 n, 10 e [1]
- (c) same number of electrons / same electronic configuration [1]; chemical reactions depend on (outer) electrons [1]
- (d) ^{26}Mg denser, because its atoms are heavier (two more neutrons) but about the same size [1]
- (e) $^{24}\text{Mg}^{2+}$ deflected more [1]; same charge but smaller mass, so larger charge/mass ratio ($2/24 > 2/26$) [1]

Question 2 [8 marks]

Titanium, Ti, has proton number 22.

- (a) Give the full electronic configuration of a titanium atom and the shorthand configuration of a Ti^{2+} ion. [2]
- (b) Draw the electrons-in-boxes diagram for the $3d$ and $4s$ sub-shells of a titanium atom, and state the number of unpaired electrons. [2]
- (c) Explain why the two $3d$ electrons in titanium occupy different orbitals. [1]

(d) Sketch the shape of a 4s orbital and of one 3p orbital, each centred on the nucleus. [2]

(e) State what is meant by the *ground state* of an atom. [1]

Mark scheme

- (a) Ti $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$ 1; $Ti^{2+} [Ar] 3d^2$ 1
- (b) 3d: two single electrons in separate boxes with parallel spins, 4s: one pair 1; 2 unpaired electrons 1
- (c) electrons in the same orbital repel (spin-pair repulsion), so singly occupied orbitals give lower energy 1
- (d) sphere for 4s 1; dumbbell (two lobes either side of the nucleus) for 3p 1
- (e) the lowest-energy arrangement of the electrons 1



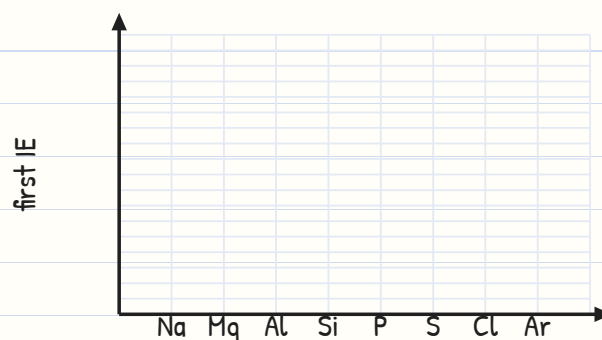
(b) $Ti: 3d^2, 4s^2$.

Question 3 [9 marks]

This question is about the ionisation energies of the elements in Period 3.

(a) Define the term *first ionisation energy*. [2]

(b) On a copy of the axes below, sketch the pattern of first ionisation energies for the elements sodium to argon. [3]

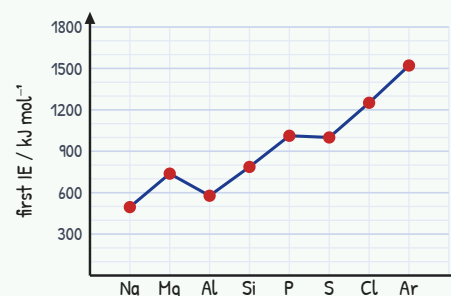


(c) Explain why the first ionisation energy of sulfur is lower than that of phosphorus. [2]

(d) The first eight ionisation energies of element Z, in kJ mol^{-1} , are 1000, 2252, 3357, 4556, 7004, 8496, 27×10^7 , 31×10^7 . Deduce the group of Z and explain your answer. [2]

Mark scheme

- (a) energy to remove one electron from each atom in one mole of 1; gaseous atoms, forming one mole of gaseous $1+$ ions 1
- (b) general increase from Na to Ar, with Ar highest and Na lowest 1; Al lower than Mg 1; S lower than P 1
- (c) in S the electron is removed from a 3p orbital containing a pair of electrons / P has three unpaired 3p electrons 1; spin-pair repulsion makes the electron easier to remove 1
- (d) Group 16 1; big jump between the 6th and 7th IE, so 6 electrons in the outer shell 1



Mark-scheme sketch (b).

Quick-reference sheet

Item	Key facts
particles	p: +1, 1; n: 0, 1; e: -1, 1/1836. Neutrons = $A - Z$; electrons = $Z - \text{ion charge}$
electric field	deflection $\propto$ charge/mass; e ⁻ to + plate (large), p to - plate (small), n undeflected
radius	decreases across a period, increases down a group; cations smaller, anions larger than atoms
isotopes	same protons, different neutrons; same chemistry (same electrons); different mass and density
orbitals	s 1 (2 e), p 3 (6 e), d 5 (10 e); s spherical, p dumbbell; order 1s 2s 2p 3s 3p 4s 3d 4p
configurations	singly before pairing; Cr [Ar] 3d ⁵ 4s ¹ , Cu [Ar] 3d ¹⁰ 4s ¹ ; ions lose 4s first
first IE	$X(g) \longrightarrow X^+(g) + e^-$; factors: nuclear charge, distance, shielding, spin-pair repulsion
IE trends	increase across a period (dips at Group 13 and Group 16); decrease down a group; big jump = new inner shell

"I can" checklist

- ☐ describe the nuclear atom and the charges, masses and positions of the three particles
- ☐ predict how beams of protons, neutrons and electrons behave in an electric field
- ☐ work out protons, neutrons and electrons in any atom or ion
- ☐ explain trends in atomic and ionic radius
- ☐ define isotope and explain their chemical and physical properties
- ☐ write full, shorthand and electrons-in-boxes configurations for H to Kr and their ions
- ☐ sketch s and p orbitals and recognise a free radical
- ☐ define first IE, write IE equations and explain the trends and the successive IE pattern
- ☐ deduce configuration and group from successive ionisation energies