

Structure 1.3.2

Electron Configurations

Sublevels, orbitals, the three rules, ions — plus HL ionization energy

Covers: Structure 1.3.4 and 1.3.5 (SL + HL); Structure 1.3.6 and 1.3.7 (HL only)

Level coding used throughout: **SL + HL** = examinable for everyone | **HL ONLY** = AHL, not required at SL

Prerequisite: Structure 1.3.1 notes (emission spectra, energy levels, the $2n^2$ rule).

0 Syllabus map

IB syllabus point	Level	What you must be able to do	Section
Structure 1.3.4	SL + HL	Recognise that main energy levels divide into s, p, d and f sublevels of successively higher energy. Recognise the shape and orientation of s and p atomic orbitals.	1, 2
Structure 1.3.5	SL + HL	Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to deduce electron configurations for atoms and ions up to $Z = 36$.	3–6
Structure 1.3.6	HL ONLY	Explain the trends and discontinuities in first ionization energy across a period and down a group.	8
Structure 1.3.7	HL ONLY	Deduce the group of an element from its successive ionization energy data.	9

1 From main energy levels to sublevels **SL + HL**

The simple "shell" model of 1.3.1 is refined: each main energy level n splits into **sublevels** labelled **s, p, d and f**, of successively higher energy within that level. Each sublevel is made of a fixed number of **orbitals**, and each orbital holds a maximum of **two** electrons.

Sublevel	Number of orbitals	Max electrons	First appears in level
s	1	2	$n = 1$
p	3	6	$n = 2$
d	5	10	$n = 3$
f	7	14	$n = 4$

MEMORISE: 1, 3, 5, 7 → 2, 6, 10, 14

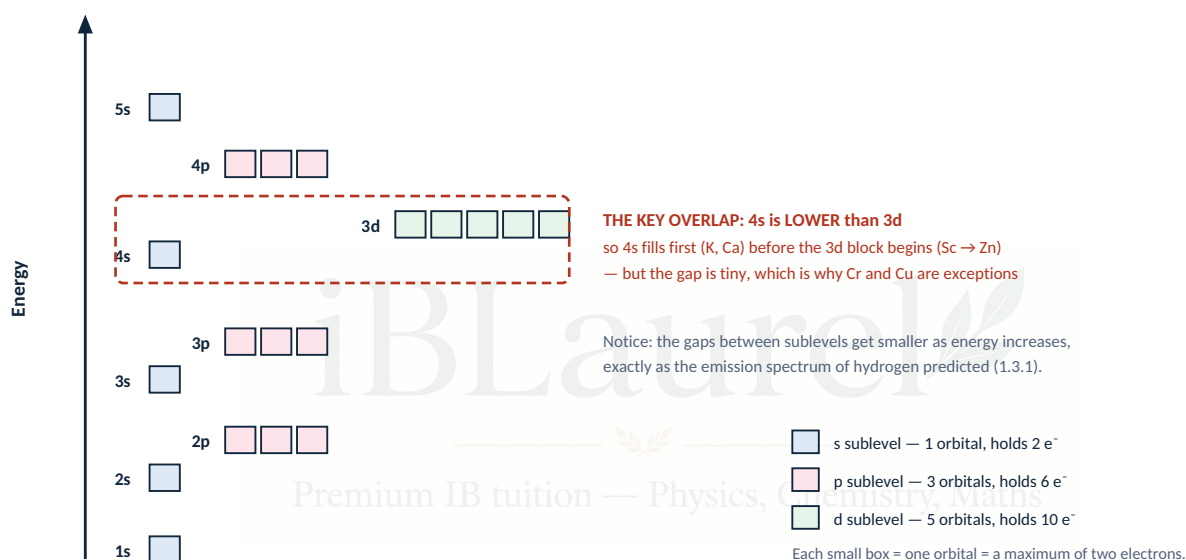
The number of orbitals goes up in **odd numbers** (1, 3, 5, 7); the electron capacity is simply **double** that (2, 6, 10, 14). Every other number in this topic can be rebuilt from those two rows, so learn them and nothing else.

Main level n	Sublevels present	Total orbitals (n^2)	Max electrons ($2n^2$)
1	1s	1	2
2	2s, 2p	4	8
3	3s, 3p, 3d	9	18
4	4s, 4p, 4d, 4f	16	32

1.1 The energy ordering — and the 4s / 3d overlap

Sublevel energies rise in the order $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$. The crucial point is that once we reach the fourth level, the **4s sublevel lies slightly below the 3d sublevel**, so 4s is filled first. This single fact explains why the third period ends at argon and why the d-block begins at scandium.

FIGURE 1 • RELATIVE ENERGIES OF THE SUBLEVELS (TO $Z = 36$)



2 Orbitals and their shapes SL + HL

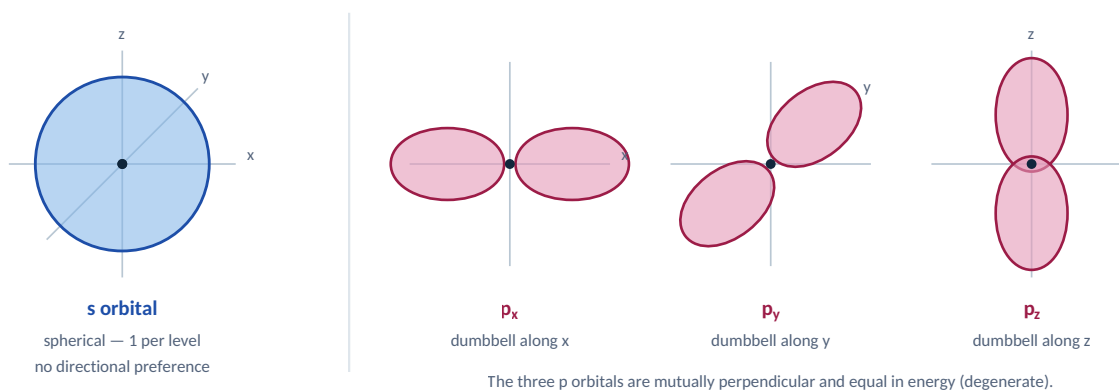
DEFINITION — ATOMIC ORBITAL

An **orbital** is a region of space around the nucleus in which there is a **high probability (conventionally 95%) of finding an electron**. Each orbital can hold a maximum of **two electrons**, and those two must have **opposite spin**.

Note the wording: an orbital is *not* a path or an orbit. Never describe electrons as "circling" the nucleus — that is the discarded Bohr picture and examiners penalise it.

- **s orbitals are spherical**. There is one s orbital per main energy level; a 2s orbital has the same shape as a 1s orbital but is larger and higher in energy.
- **p orbitals are dumbbell-shaped** (two lobes either side of the nucleus). There are three of them, **mutually perpendicular**, aligned along the x, y and z axes: p_x , p_y , p_z . They are **degenerate** — equal in energy.
- You are **not** required to know the shapes of d or f orbitals at SL or HL in this topic.

FIGURE 2 • SHAPES AND ORIENTATIONS OF S AND P ATOMIC ORBITALS



The black dot marks the position of the nucleus. Shading shows the 95% probability boundary, not a physical surface.

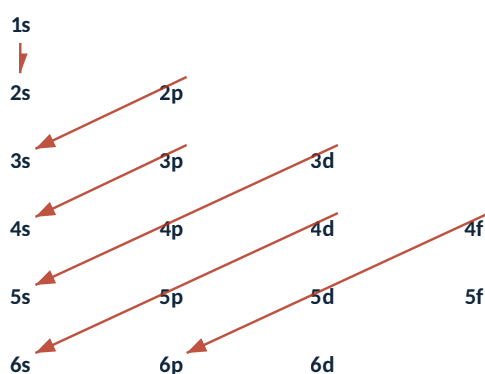
3 The three rules that build every configuration SL + HL

Rule	Statement	In practice
Aufbau principle ("building up")	Electrons occupy the lowest available energy sublevel first .	Fill in the order 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p
Pauli exclusion principle	An orbital holds a maximum of two electrons , and they must have opposite spins .	Every full box is drawn $\uparrow\downarrow$, never $\uparrow\uparrow$
Hund's rule (of maximum multiplicity)	Within a set of degenerate orbitals, electrons occupy them singly with parallel spins before any pairing occurs.	Fill p or d boxes one at a time, all arrows up, then go back and pair

WHY HUND'S RULE WORKS — THE ONE-LINE REASON

Electrons are negatively charged and **repel** one another. Occupying separate orbitals keeps them further apart and minimises repulsion, so the singly-occupied arrangement is lower in energy.

FIGURE 3 • THE AUFBAU FILLING ORDER



Read the diagonals from top-right to bottom-left:

$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p$

For IB (up to $Z = 36$) you only ever need:

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

$2 + 2 + 6 + 2 + 6 + 2 + 10 + 6 = 36$ electrons (krypton)

Alternative memory route: read the periodic table itself left to right, row by row — it is built in exactly this order.

FIGURE 4 • HUND'S RULE AND THE PAULI PRINCIPLE IN ORBITAL DIAGRAMMS

NITROGEN (Z = 7): $1s^2 2s^2 2p^3$ — CORRECT

three 2p electrons occupy separate orbitals,
all spins parallel — Hund's rule satisfied

NITROGEN — WRONG (a very common error)



pairing before every degenerate orbital has one
electron — breaks Hund's rule
($\uparrow \uparrow$ in one box would break Pauli as well)

OXYGEN (Z = 8): $1s^2 2s^2 2p^4$ — the fourth p electron must now pair up

Only after all three p orbitals hold one electron
does pairing begin. This extra pair repulsion is the
reason O has a LOWER first IE than N (see §8).

Convention: an up arrow and a down arrow in the same box represent two electrons of opposite spin (Pauli).

4 Writing electron configurations SL + HL

4.1 Method

- Find the number of electrons (= Z for a neutral atom).
- Fill in Aufbau order: **1s 2s 2p 3s 3p 4s 3d 4p**, using capacities 2, 2, 6, 2, 6, 2, 10, 6.
- Stop when you have used all the electrons.
- Rewrite in order of n**: quote 3d *before* 4s in the final answer.

THE 4S / 3D WRITING CONVENTION — DO NOT LOSE THIS MARK

We **fill** 4s before 3d, but once 3d contains electrons it drops below 4s in energy, so we **write** 3d first.

Iron: fill as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6 \rightarrow$ **write as** $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$.

Both orders are usually accepted by IB, but the 3d-first version is the expected one and is essential for getting ions right.

4.2 Condensed (noble gas core) notation

Replace the inner electrons with the previous noble gas in square brackets. Fe = $[\text{Ar}] 3d^6 4s^2$. This is faster, less error-prone, and always accepted.

Element	Z	Full configuration	Condensed
N	7	$1s^2 2s^2 2p^3$	$[\text{He}] 2s^2 2p^3$
Cl	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[\text{Ne}] 3s^2 3p^5$
Ca	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	$[\text{Ar}] 4s^2$
Sc	21	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$	$[\text{Ar}] 3d^1 4s^2$
Fe	26	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$	$[\text{Ar}] 3d^6 4s^2$
Zn	30	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$	$[\text{Ar}] 3d^{10} 4s^2$
Br	35	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$	$[\text{Ar}] 3d^{10} 4s^2 4p^5$
Kr	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$	$[\text{Kr}]$

WORKED EXAMPLE 1 — BROMINE (Z = 35)

Fill: $1s^2 (2) \rightarrow 2s^2 (4) \rightarrow 2p^6 (10) \rightarrow 3s^2 (12) \rightarrow 3p^6 (18) \rightarrow 4s^2 (20) \rightarrow 3d^{10} (30) \rightarrow 4p^5 (35)$. ✓

Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$ or $[\text{Ar}] 3d^{10} 4s^2 4p^5$

Always check that your superscripts add up to Z. $2+2+6+2+6+10+2+5 = 35$. ✓

5 The two exceptions: chromium and copper SL + HL

The 4s and 3d sublevels are extremely close in energy. For two elements, a small energy gain from a **half-full or completely full 3d sublevel** outweighs the cost of promoting an electron out of 4s.

Element	Predicted (wrong)	Actual	Reason
Cr (24)	$[\text{Ar}] 3d^4 4s^2$	$[\text{Ar}] 3d^5 4s^1$	A half-full 3d sublevel ($3d^5$) plus a half-full 4s gives extra stability
Cu (29)	$[\text{Ar}] 3d^9 4s^2$	$[\text{Ar}] 3d^{10} 4s^1$	A completely full 3d sublevel ($3d^{10}$) gives extra stability

MEMORISE

Only **two** exceptions appear in the IB syllabus: **Cr and Cu**. In both, one electron moves from 4s to 3d, leaving $4s^1$. Chromium goes to d^5 (half-full), copper to d^{10} (full).

6 Electron configurations of ions SL + HL

Negative ions (anions)

Add electrons in the normal Aufbau order. $\text{Cl}^- (18 e^-) = 1s^2 2s^2 2p^6 3s^2 3p^6 = [\text{Ar}]$.

Positive ions (cations)

Remove electrons from the **outermost main energy level first** — that is, from the level with the highest n , not simply the last sublevel you filled.

THE TRANSITION-METAL RULE — THE SINGLE BIGGEST SOURCE OF LOST MARKS IN 1.3

For d-block elements, **4s electrons are lost before 3d electrons**, because 4s is in the outer ($n = 4$) main energy level.

$\text{Fe} = [\text{Ar}] 3d^6 4s^2 \rightarrow \text{Fe}^{2+} = [\text{Ar}] 3d^6$ (not $3d^4 4s^2$) $\rightarrow \text{Fe}^{3+} = [\text{Ar}] 3d^5$

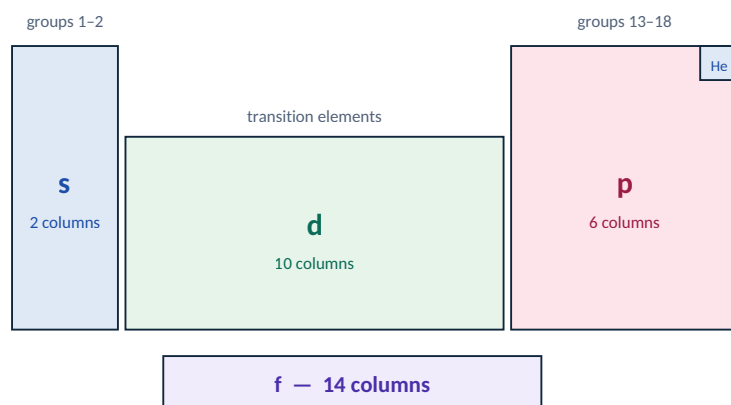
Writing 3d before 4s in the atom makes this automatic: just delete from the right-hand end.

Ion	Electrons	Configuration	Note
Na^+	10	$1s^2 2s^2 2p^6$	Isoelectronic with Ne
O^{2-}	10	$1s^2 2s^2 2p^6$	Isoelectronic with Ne and Na^+
Ca^{2+}	18	$[\text{Ne}] 3s^2 3p^6$	$= [\text{Ar}]$
Fe^{2+}	24	$[\text{Ar}] 3d^6$	Both 4s electrons removed
Fe^{3+}	23	$[\text{Ar}] 3d^5$	Half-full 3d — extra stability, hence Fe^{3+} is common
Cu^+	28	$[\text{Ar}] 3d^{10}$	From $[\text{Ar}] 3d^{10} 4s^1$; full 3d
Cu^{2+}	27	$[\text{Ar}] 3d^9$	4s electron then one 3d electron
Zn^{2+}	28	$[\text{Ar}] 3d^{10}$	Full 3d — why Zn is not classed as a transition element
Cr^{3+}	21	$[\text{Ar}] 3d^3$	From $[\text{Ar}] 3d^5 4s^1$: lose $4s^1$ then two 3d

7 Reading configurations off the periodic table

SL + HL

FIGURE 5 • THE BLOCKS OF THE PERIODIC TABLE MAP DIRECTLY ONTO SUBLEVELS



How to use it

- The **block** tells you the sublevel being filled.
- The **period number** gives n for s and p blocks; for the d block it is (period – 1).
- The **position within the block** gives the superscript.

Example: selenium, period 4, p block, 4th column

→ ends in $4p^4$

→ $[\text{Ar}] 3d^{10} 4s^2 4p^4$ ($Z = 34$ ✓)

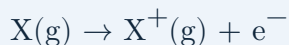
FAST CHECKS

- The number of electrons in the outer main energy level = the **group number** (groups 1, 2) or group number – 10 (groups 13–18).
- The **period number** = the highest value of n that is occupied.
- Your superscripts must always sum to Z (atoms) or $Z - \text{charge}$ (ions).

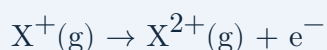
8 First ionization energy: trends and discontinuities HL ONLY

DEFINITIONS

First ionization energy: the energy required to remove one mole of electrons from one mole of gaseous atoms in their ground state.



Second ionization energy: removal of a second mole of electrons from the resulting 1+ ions.



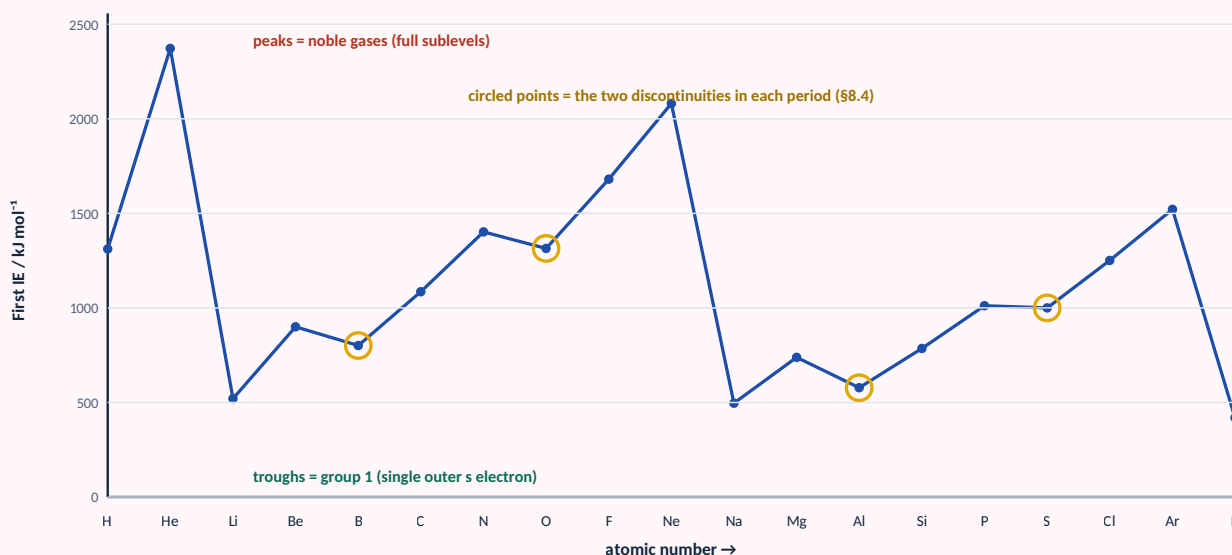
Ionization is always **endothermic**, so every ionization energy is positive.

8.1 The three factors that control ionization energy

Every explanation in this section is built from the same three factors. Name all three that are relevant, and say which one dominates.

Factor	Effect on IE	Comment
Nuclear charge (number of protons)	Greater nuclear charge → stronger attraction → higher IE	Increases across a period
Atomic radius / distance of the outer electron from the nucleus	Greater distance → weaker attraction → lower IE	Increases down a group
Shielding (screening by inner electrons)	More inner shells → more repulsion of the outer electron → lower IE	Roughly constant across a period; increases down a group

FIGURE 6 • FIRST IONIZATION ENERGY AGAINST ATOMIC NUMBER (H TO K)



Each period begins at a group 1 minimum and rises to a noble gas maximum. The two small dips within period 2 (B and O) and period 3 (Al and S) are the discontinuities explained below.

8.2 Across a period: a general increase

Across period 3, from Na to Ar: the nuclear charge increases by one proton per element, while electrons are added to the **same** main energy level, so shielding is essentially unchanged. The effective nuclear attraction on the outer electron therefore increases and the atomic radius decreases, so the **first ionization energy rises**.

8.3 Down a group: a decrease

Down group 1, from Li to K: each element has one more occupied main energy level, so the outer electron is **further from the nucleus** and is **shielded by more inner electrons**. These two effects outweigh the increase in nuclear charge, so the **first ionization energy falls**.

8.4 The two discontinuities — you must be able to explain both

Dip 1: group 2 → group 13 (Be → B, Mg → Al)

- Beryllium's outer electron is removed from a **2s** orbital; boron's is removed from a **2p** orbital.
- The **2p sublevel is higher in energy** than 2s, and a 2p electron is slightly further from the nucleus and better shielded by the 2s electrons.
- So less energy is needed to remove it, and the first IE **falls** despite the extra proton.

Dip 2: group 15 → group 16 (N → O, P → S)

- Nitrogen is $2p^3$ — all three p orbitals are **singly occupied**.
- Oxygen is $2p^4$ — one orbital now contains a **pair** of electrons.
- The two paired electrons in the same orbital **repel each other**, so one of them is easier to remove.
- The first IE therefore **falls** from N to O.

PRECISION MATTERS HERE

Do **not** write "nitrogen's half-full p sublevel is especially stable" on its own — IB mark schemes want the **reason**: in oxygen there is **repulsion between the two electrons paired in the same 2p orbital**. Equally, for Be → B, say that the electron is removed from a **higher-energy p sublevel**, not merely "a different orbital".

WORKED EXAMPLE 2 — EXPLAINING AN ANOMALY

Q. The first ionization energy of sulfur (1000 kJ mol^{-1}) is lower than that of phosphorus (1012 kJ mol^{-1}), although sulfur has a higher nuclear charge. Explain. [2]

A. Phosphorus has the configuration $[\text{Ne}] 3s^2 3p^3$, with one electron in each 3p orbital, whereas sulfur is $[\text{Ne}] 3s^2 3p^4$ and has two electrons paired in one 3p orbital. [1] The repulsion between these paired electrons makes one of them easier to remove, so sulfur's first ionization energy is lower. [1]

9 Successive ionization energies HL ONLY

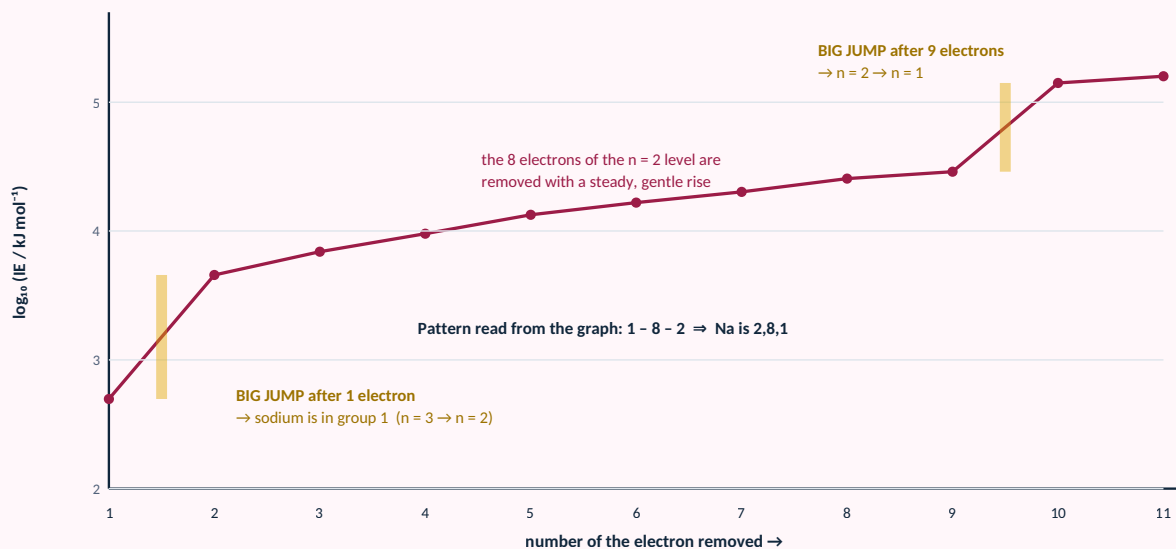
Removing successive electrons from the same atom always requires **more** energy each time, because the remaining electrons are held by the same nuclear charge in an increasingly positive ion. What matters for identification is **where the large jumps occur**.

THE RULE

A **large jump** in successive ionization energies occurs when an electron is removed from a **new, inner main energy level**, which is closer to the nucleus and less shielded.

The number of electrons removed before the first large jump = the group number.

FIGURE 7 • SUCCESSIVE IONIZATION ENERGIES OF SODIUM (LOGARITHMIC SCALE)



A logarithmic scale is used because the values span three orders of magnitude; the jumps remain clearly visible.

WORKED EXAMPLE 3 — DEDUCING THE GROUP

Q. The first six ionization energies of an element, in kJ mol^{-1} , are 578, 1817, 2745, 11 578, 14 831, 18 378. Deduce the group of the element and justify your answer. [2]

A. The first large jump comes between the **third and fourth** ionization energies ($2745 \rightarrow 11\,578$, a rise of more than four-fold). [1] Three electrons are therefore removed relatively easily from the outer main energy level, so the element is in **group 13**. [1]

(The data are those of aluminium.)

HOW TO SPOT "THE BIG JUMP" RELIABLY

Do not judge by the raw difference — judge by the **ratio**. A genuine change of main energy level typically at least **doubles** the value; the steady rise within a level is usually well under a factor of two.

10 Common errors checklist

SL + HL

Error	Correction
Writing Fe^{2+} as $[\text{Ar}] 3d^4 4s^2$	Remove the 4s electrons first: $[\text{Ar}] 3d^6$
Pairing electrons in p orbitals too early	Hund's rule: singly occupy all degenerate orbitals first
Two arrows pointing the same way in one box	Pauli: the two electrons in an orbital must have opposite spin
Cr written as $[\text{Ar}] 3d^4 4s^2$; Cu as $[\text{Ar}] 3d^9 4s^2$	$3d^5 4s^1$ and $3d^{10} 4s^1$
Describing an orbital as an "orbit" or "path"	A region of space with a high probability of finding an electron
HL "N has a stable half-full sublevel" as the whole explanation of the $\text{N} \rightarrow \text{O}$ dip	State the repulsion between the paired electrons in oxygen's 2p orbital
HL Forgetting "gaseous" or "one mole" in the IE definition	All three of one mole / gaseous / ground state are needed

11 Practice questions

Roughly 1.5 minutes per mark. Full answers with mark-scheme wording follow in Section 12.

Part A — Multiple choice SL + HL

- How many orbitals are there in the third main energy level?
A. 3 B. 5 C. 9 D. 18
- Which is the correct ground-state electron configuration of chromium?
A. $[\text{Ar}] 3d^4 4s^2$ B. $[\text{Ar}] 3d^5 4s^1$ C. $[\text{Ar}] 3d^6$ D. $[\text{Ar}] 3d^3 4s^3$
- Which statement describes Hund's rule?
A. An orbital holds at most two electrons of opposite spin.
B. Electrons fill the lowest energy sublevel first.
C. Degenerate orbitals are each singly occupied before pairing begins.
D. The 4s sublevel is filled before the 3d sublevel.
- What is the electron configuration of Cu^{2+} ?
A. $[\text{Ar}] 3d^9$ B. $[\text{Ar}] 3d^{10}$ C. $[\text{Ar}] 3d^7 4s^2$ D. $[\text{Ar}] 3d^8 4s^1$
- How many unpaired electrons are there in a ground-state atom of phosphorus?
A. 0 B. 1 C. 3 D. 5
- HL ONLY Which element has the lowest first ionization energy?
A. Li B. Na C. K D. Mg
- HL ONLY The successive ionization energies of an element show a large increase between the second and third values. The element is in group:
A. 1 B. 2 C. 3 D. 13

Part B — Structured questions SL + HL

- State the maximum number of electrons that can occupy (a) a single orbital and (b) the 3d sublevel, and explain briefly why they differ. [2]
- Draw the full orbital diagram (arrow-in-box) for a ground-state atom of oxygen. State the two rules you used and explain how each is shown in your diagram. [3]
- Write the full electron configurations of (a) Ca, (b) Fe and (c) Br. [3]
- Write the electron configurations of (a) Fe^{3+} , (b) Zn^{2+} and (c) S^{2-} . [3]
- Explain why the ground-state configuration of copper is $[\text{Ar}] 3d^{10} 4s^1$ rather than the predicted $[\text{Ar}] 3d^9 4s^2$. [2]
- Sketch and label the shapes of a 2s orbital and a $2p_x$ orbital. State one similarity and one difference between the three 2p orbitals. [3]
- Explain why the 4s sublevel is filled before the 3d sublevel, and state the consequence of this for the electron configuration of potassium. [2]

Part C — HL extension HL ONLY

- Define the term *first ionization energy* and write an equation, including state symbols, for the first ionization of magnesium. [2]
- Explain why the first ionization energy increases from sodium to argon. [3]
- Explain why the first ionization energy of aluminium (578 kJ mol^{-1}) is lower than that of magnesium (738 kJ mol^{-1}). [3]
- Explain why the first ionization energy decreases from lithium to potassium. [2]

- 19.** The first five ionization energies of an element X, in kJ mol^{-1} , are: [3]
738, 1451, 7733, 10 540, 13 630.
(a) Deduce the group to which X belongs, justifying your answer. [2]
(b) Suggest the identity of X. [1]
- 20.** Explain why the second ionization energy of any element is always greater than its first ionization energy. [2]



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12 Answers & mark-scheme notes

Part A

1. **C** — level 3 contains $3s$ (1) + $3p$ (3) + $3d$ (5) = 9 orbitals, i.e. n^2 , holding $2n^2 = 18$ electrons.
2. **B** — one $4s$ electron is promoted to $3d$ to give a half-full $3d^5$ sublevel.
3. **C** — A is Pauli, B is Aufbau, D is a consequence of Aufbau.
4. **A** — Cu is $[\text{Ar}] 3d^{10} 4s^1$; remove the $4s$ electron, then one $3d$ electron $\rightarrow [\text{Ar}] 3d^9$.
5. **C** — P is $[\text{Ne}] 3s^2 3p^3$; by Hund's rule the three $3p$ electrons occupy separate orbitals, so 3 are unpaired.
6. **C** — potassium's outer electron is in the highest main energy level of the four, so it is furthest from the nucleus and most shielded.
7. **B** — two electrons are removed easily before the jump, so the outer level holds two electrons.

Part B

8. [2]

- (a) **2** electrons (with opposite spins). [1]
 (b) **10** electrons, because the d sublevel contains **five** orbitals, each holding two electrons. [1]

9. [3]

Diagram: $1s \uparrow \downarrow \quad 2s \uparrow \downarrow \quad 2p \uparrow \downarrow \mid \uparrow \mid \uparrow$ [1]

- **Pauli exclusion principle:** each box contains at most two electrons and they are drawn with opposite spins. [1]
- **Hund's rule:** the three $2p$ orbitals are degenerate, so the first three p electrons occupy them singly with parallel spins; only the fourth pairs up. [1]

10. [3]

- (a) Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ [1]
 (b) Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ [1]
 (c) Br: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$ [1]
 $4s$ before $3d$ is also accepted, but be consistent.

11. [3]

- (a) Fe^{3+} : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ ($[\text{Ar}] 3d^5$) [1]
 (b) Zn^{2+} : $[\text{Ar}] 3d^{10}$ [1]
 (c) S^{2-} : $1s^2 2s^2 2p^6 3s^2 3p^6$ ($= [\text{Ar}]$) [1]

Marker's note: in (a) and (b) the $4s$ electrons must be removed first. $[\text{Ar}]3d^3 4s^2$ for Fe^{3+} scores zero.

12. [2]

- The $3d$ and $4s$ sublevels are very close in energy. [1]
- Promoting one $4s$ electron gives a **completely full $3d^{10}$** sublevel, which is a lower-energy (more stable) arrangement overall. [1]

13. [3]

- $2s$ drawn as a **sphere** centred on the nucleus. [1]
- $2p_x$ drawn as two lobes (a **dumbbell**) along the x -axis, with the nucleus between them. [1]
- Similarity: the three $2p$ orbitals have the **same shape and the same energy** (degenerate). Difference: they point along **different, mutually perpendicular axes** (x , y , z). [1]

14. [2]

- The $4s$ sublevel is **lower in energy** than $3d$, and by the Aufbau principle the lowest available sublevel is filled first. [1]
- Potassium ($Z = 19$) therefore ends in **$4s^1$** , i.e. $[\text{Ar}] 4s^1$, rather than $[\text{Ar}] 3d^1$. [1]

Part C **HL ONLY****15. [2]**

- The energy required to remove **one mole of electrons** from **one mole of gaseous atoms** in their **ground state**. [1]
- $\text{Mg(g)} \rightarrow \text{Mg}^+(\text{g}) + \text{e}^-$ (state symbols required) [1]

16. [3]

- Nuclear charge (number of protons) increases from Na to Ar. [1]
- Electrons are added to the **same main energy level**, so shielding remains approximately constant and the atomic radius decreases. [1]
- The attraction between the nucleus and the outer electron therefore increases, so more energy is needed to remove it. [1]

17. [3]

- Mg is $[\text{Ne}] 3s^2$, so its electron is removed from a **3s** orbital; Al is $[\text{Ne}] 3s^2 3p^1$, so its electron is removed from a **3p** orbital. [1]
- The 3p sublevel is **higher in energy** than 3s and is further from the nucleus / better shielded by the 3s electrons. [1]
- Less energy is therefore needed to remove the 3p electron, so Al has the lower first ionization energy. [1]

18. [2]

- Going down the group, the outer electron occupies a **higher main energy level**, so it is further from the nucleus, and shielding by inner electrons increases. [1]
- The attraction between the nucleus and the outer electron is therefore weaker (these effects outweigh the increased nuclear charge), so less energy is required to remove it. [1]

19. [3]

(a) The large jump occurs between the **second and third** ionization energies ($1451 \rightarrow 7733$, more than a five-fold increase). [1] Two electrons are removed from the outer main energy level before an electron must be taken from a lower level closer to the nucleus, so X is in **group 2**. [1]

(b) **Magnesium** (a first ionization energy of 738 kJ mol^{-1} matches Mg). [1]

20. [2]

- The second electron is removed from a **positive ion** rather than a neutral atom. [1]
- The remaining electrons are attracted by the same nuclear charge but there is now one fewer electron, so each is held more strongly (the ion is also smaller), and more energy is required. [1]

End of Structure 1.3.2 · Together with Structure 1.3.1 this completes syllabus topic **Structure 1.3 — Electron configurations**.