CHEMISTRY
TOPIC 2

Complete Typed Study Notes

Electronic Configuration • Orbitals • Ionisation Energy
Periodic Trends • Mass Spectrometry • Isotopes

About this document
These notes are transcribed and organised from the 13 pages of handwritten Topic 2 material. The sequence, terminology, examples and calculations have been retained and arranged into clear revision sections.

Source notes dated 10 September 2026

1. Electronic Configuration

Electronic configuration describes how electrons are arranged in the shells, subshells and orbitals of an atom or ion.

Shells, subshells, orbitals and electrons

The arrangement can be thought of as a hierarchy:

Shell → Subshell → Orbital → Electron

A shell contains one or more subshells. Each subshell contains orbitals, and each orbital can hold a maximum of two electrons.

The principal shell number is represented by **n**. The first shell is n = 1, the second is n = 2, and so on.

Subshell types and electron capacities

Subshell	Number of orbitals	Maximum number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

Each individual orbital can contain a maximum of two electrons, and the two electrons must have opposite spins.

Example: sodium

Sodium has atomic number 11, so a neutral sodium atom has 11 electrons.

$$\text{Na: } 1s^2\ 2s^2\ 2p^6\ 3s^1$$

The shell arrangement is 2, 8, 1. The outermost electron is in the 3s subshell.

Order of filling orbitals: Aufbau principle

Electrons occupy orbitals in order of increasing energy, starting with the lowest-energy orbital available.

1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p

5s → 4d → 5p → 6s → 4f → 5d → 6p

7s → 5f → 6d → 7p

> **Filling reminder**
> The diagonal filling order is used to identify which subshell receives the next electron. A subshell can hold s = 2, p = 6, d = 10 and f = 14 electrons.

2. Electron Configurations and the Periodic Table

For a neutral atom, the number of electrons is equal to its atomic number. The electron configuration can be used to identify its position in the periodic table.

Examples from the notes

Element	Atomic number	Electron configuration	Group / classification
K (potassium)	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	Group 1
Sr (strontium)	38	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$	Group 2
Ga (gallium)	31	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$	Group 3 (old group notation)
Br (bromine)	35	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$	Group 7 (old group notation)
Fe (iron)	26	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$	Transition element

Blocks of the periodic table

Block	Groups in the notation used in the notes	Outer subshell being filled
s-block	Groups 1 and 2	s subshell
p-block	Groups 3 to 8	p subshell
d-block	Transition elements	d subshell

The notes use the older IGCSE group numbering (Groups 1 to 8 for the main-group elements).

Ions and electron configurations

When a positive ion forms, electrons are removed. When a negative ion forms, electrons are gained. For transition-metal ions, electrons are removed from the 4s subshell before the 3d subshell.

Species	Electron configuration
Al^{3+}	$1s^2\ 2s^2\ 2p^6$
S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Fe (atom)	$[Ar]\ 3d^6\ 4s^2$
Fe^{3+}	$[Ar]\ 3d^5$

Counting electrons in ions
Cations have lost electrons: electrons = atomic number - positive charge. Anions have gained electrons: electrons = atomic number + magnitude of negative charge.

Ion / species	Electrons in ion	Atomic number of neutral atom
M^{2+}	18	20
M^{3+}	10	13
M^{3-}	10	7
M^{2-}	18	16
M^{+}	18	19

3. Shapes of Orbitals and Box Notation

s and p orbitals

An **s orbital** is spherical. The 1s, 2s and 3s orbitals are all spherical, but they differ in size and energy.

A **p orbital** has a dumbbell shape. There are three p orbitals in a p subshell, oriented along three perpendicular directions: p_x, p_y and p_z.

The d subshell contains five orbitals. The five d orbitals are commonly labelled d_{xy}, d_{xz}, d_{yz}, $d_{x^2-y^2}$ and d_{z^2}.

Orbital box notation

An orbital is represented by a box, and an electron is represented by an arrow. The direction of the arrow represents electron spin. Two electrons in the same orbital must have opposite spins.

One orbital: [↑↓] One unpaired electron: [↑]

Examples of outer configurations:

Atom	Electron configuration	Outer-shell orbital arrangement
Na (11)	$1s^2\ 2s^2\ 2p^6\ 3s^1$	3s: [↑]
Cl (17)	$[Ne]\ 3s^2\ 3p^5$	3s: [↑↓] 3p: [↑↓] [↑↓] [↑]
S (16)	$[Ne]\ 3s^2\ 3p^4$	3s: [↑↓] 3p: [↑↓] [↑] [↑]
P (15)	$[Ne]\ 3s^2\ 3p^3$	3s: [↑↓] 3p: [↑] [↑] [↑]

Orbital capacity
Each orbital can accommodate a maximum of two electrons. A p subshell has three boxes (six electrons maximum); a d subshell has five boxes (ten electrons maximum).

4. Shorthand Electron Configuration and Transition Metals

Shorthand electron configuration is a shorter way of writing an electron configuration. The inner electrons are represented by the symbol of the preceding noble gas in square brackets.

Element	Full configuration	Shorthand configuration
Cl (17)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	[Ne] $3s^2\ 3p^5$
Mg (12)	$1s^2\ 2s^2\ 2p^6\ 3s^2$	[Ne] $3s^2$
Fe (26)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$	[Ar] $4s^2\ 3d^6$

Anomalous electron configurations

Some transition elements have configurations that differ from the simple filling order. A half-filled d subshell (d^5) and a completely filled d subshell (d^{10}) are relatively stable arrangements.

Element	Expected arrangement from simple filling	Observed arrangement
Chromium, Cr (24)	[Ar] $3d^4\ 4s^2$	[Ar] $3d^5\ 4s^1$
Copper, Cu (29)	[Ar] $3d^9\ 4s^2$	[Ar] $3d^{10}\ 4s^1$

Chromium has a half-filled 3d subshell, while copper has a completely filled 3d subshell. These arrangements are associated with extra stability.

Forming transition-metal ions

Although 4s is filled before 3d, electrons are removed from 4s before 3d when transition-metal cations form.

$$\mathbf{Fe: [Ar]\ 3d^6\ 4s^2}$$

$$\mathbf{Fe^{2+}: [Ar]\ 3d^6}$$

$$\mathbf{Fe^{3+}: [Ar]\ 3d^5}$$

5. Ionisation Energy

First ionisation energy: The energy required to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions.

$$\mathbf{X(g) \rightarrow X^+(g) + e^-}$$

The process is endothermic because energy is needed to remove an electron from the attraction of the nucleus.

Successive ionisation energies

Successive ionisation energies are the energies required to remove further electrons, one at a time, from gaseous ions.

$$\mathbf{First:\ \ X(g) \rightarrow X^+(g) + e^-}$$

$$\mathbf{Second:\ \ X^+(g) \rightarrow X^{2+}(g) + e^-}$$

$$\mathbf{Third:\ \ X^{2+}(g) \rightarrow X^{3+}(g) + e^-}$$

Each successive ionisation energy is greater than the previous one because the remaining electrons experience a stronger attraction to the increasingly positive ion.

Factors affecting ionisation energy

Factor	Effect on how strongly the outer electron is held
Nuclear charge / proton number	A greater positive nuclear charge attracts electrons more strongly, increasing ionisation energy.
Number of occupied shells	More shells generally place the outer electron further from the nucleus, decreasing the attraction.
Shielding effect	Inner-shell electrons repel outer electrons and reduce the attraction from the nucleus. Greater shielding lowers ionisation energy.
Spin-pair repulsion	Two electrons paired in the same orbital repel each other. This can make one of the paired electrons easier to remove.

6. Trends in Ionisation Energy

Trend across a period

First ionisation energy generally increases across a period from left to right.

• The proton number and nuclear charge increase.

• Electrons are added to the same main shell, so shielding does not increase greatly.

• The attraction between the nucleus and the outer electron becomes stronger.

• Atomic size generally decreases, so the outer electron is held more strongly and is harder to remove.

Trend down a group

First ionisation energy generally decreases down a group.

• The number of occupied electron shells increases.

• The outer electron is further from the nucleus.

• Shielding by inner electrons increases.

• The attraction between the nucleus and the outer electron becomes weaker, so the electron is easier to remove.

Overall trend
Ionisation energy generally increases across a period and decreases down a group.

Anomalies in the trend across a period

There are two important small decreases in first ionisation energy across a period. These are explained by subshell energy and electron-pair repulsion.

1. Group 2 to Group 3: the p-subshell electron

Example: beryllium and boron.

$$\text{Be: } 1s^2\ 2s^2$$

$$\text{B: } 1s^2\ 2s^2\ 2p^1$$

The electron removed from boron is in a 2p subshell, which is higher in energy than the 2s subshell electron removed from beryllium. It is therefore easier to remove, so the first ionisation energy decreases from Group 2 to Group 3.

2. Group 5 to Group 6: spin-pair repulsion

Example: nitrogen and oxygen.

$$\text{N: } 1s^2\ 2s^2\ 2p^3$$

$$\text{O: } 1s^2\ 2s^2\ 2p^4$$

In nitrogen, the three 2p electrons occupy separate p orbitals. In oxygen, one of the 2p orbitals contains a pair of electrons. Repulsion between paired electrons makes one electron easier to remove from oxygen. Therefore, the first ionisation energy decreases from Group 5 to Group 6.

7. Large Jumps in Successive Ionisation Energies

A large jump occurs when an electron is removed from an inner shell rather than the outer shell. Inner-shell electrons are closer to the nucleus and experience less shielding, so much more energy is required to remove them.

The position of the large jump helps identify the number of electrons in the outer shell and hence the group of a main-group element.

• A large jump after the first electron indicates one outer-shell electron (Group 1).

• A large jump after the second electron indicates two outer-shell electrons (Group 2).

• A large jump after the third electron indicates three outer-shell electrons (Group 3), and so on.

The handwritten notes illustrate successive ionisation-energy values and show the large jump after removal of the outer-shell electrons. The exact numerical values in that illustration are not sufficiently clear to reproduce confidently; the key interpretation is the position of the jump.

8. Mass Spectrometry

Mass spectrometry provides information about isotopes and their relative abundances. The notes identify the following uses:

• Determining how many isotopes an element has.

• Finding the percentage abundance of each isotope.

• Calculating relative atomic mass.

• Determining relative molecular mass.

• Investigating the molecular structure of compounds.

• Studying diatomic molecules such as Cl_2 and Br_2.

Main stages of a mass spectrometer

Stage	What happens
1. Vaporisation	The sample is converted into a gas.
2. Ionisation	The gaseous atoms or molecules are bombarded by electrons and form positive ions.
3. Acceleration	Positive ions are accelerated by an electric field into an ion beam.
4. Deflection	The ion beam passes through a magnetic field and is deflected according to its mass-to-charge ratio.
5. Detection	Ions reach the detector, producing a mass spectrum.

Reading a mass spectrum

• The x-axis shows mass-to-charge ratio, m/z.

• The y-axis shows relative abundance or relative intensity.

• A peak represents ions with a particular m/z value.

• The height of a peak indicates the relative abundance of those ions.

$$\mathbf{m/z = \text{mass of ion (m)} \div \text{charge number (z)}}$$

For ions with the same charge, lighter ions are deflected more than heavier ions. For ions of the same mass, an ion with a greater positive charge is deflected more. The notes summarise this as: deflection is inversely proportional to mass and directly proportional to charge.

Remember
A common mass-spectrometry ion has a +1 charge. In that case, its m/z value is numerically equal to its mass number or molecular ion mass.

Monatomic element example: neon

The notes show three isotope peaks for neon at m/z 20, 21 and 22. The approximate relative abundances shown are 90.48%, 0.27% and 9.25%, respectively.

Neon isotope	m/z	Relative abundance
^{20}Ne	20	90.48%
^{21}Ne	21	0.27%
^{22}Ne	22	9.25%

The relative atomic mass is calculated as a weighted mean: **A_r = [(mass 1 × % abundance 1) + (mass 2 × % abundance 2) + ...] ÷ 100.**

9. Diatomic Molecules in Mass Spectrometry

A molecule is formed when two or more atoms combine chemically. A diatomic molecule contains two atoms. Chlorine and bromine are examples: Cl_2 and Br_2.

Chlorine isotopes

Chlorine has two main isotopes:

Isotope	Percentage abundance
^{35}Cl	75%
^{37}Cl	25%

The abundance ratio of ^{35}Cl : ^{37}Cl is 3 : 1.

Possible chlorine molecular ions

A chlorine molecule contains two chlorine atoms. The possible combinations are:

Combination	Molecular ion mass (m/z)	Relative occurrence
^{35}Cl - ^{35}Cl	70	3 × 3 = 9
^{35}Cl - ^{37}Cl	72	2 × 3 = 6
^{37}Cl - ^{37}Cl	74	1 × 1 = 1

Therefore, the molecular-ion peak ratio is:

$$^{70}Cl_2 : ^{72}Cl_2 : ^{74}Cl_2 = 9 : 6 : 1$$

The relative molecular mass is the weighted mean:

$$M_r(Cl_2) = [(9 \times 70) + (6 \times 72) + (1 \times 74)] \div 16 = 71$$

If both atomic chlorine ions (m/z 35 and 37) and molecular chlorine ions (m/z 70, 72 and 74) are detected, the spectrum can show five peaks in total.

Bromine isotopes and Br2

The notes show two bromine isotopes with approximately equal abundance:

Isotope	Percentage abundance
^{79}Br	50%
^{81}Br	50%

The possible molecular-ion combinations are:

Combination	Molecular ion mass
^{79}Br - ^{79}Br	158
^{79}Br - ^{81}Br	160
^{81}Br - ^{81}Br	162

Because the isotopes have equal abundance, the relative molecular-ion peak ratio is:

Br_2 peaks at 158 : 160 : 162 = 1 : 2 : 1

The corresponding relative intensities are 25 : 50 : 25.

10. Rules for Writing Electron Configurations

1. Aufbau principle

Electrons occupy the lowest-energy orbitals available first. The filling order is the sequence listed in Section 1.

2. Hund's rule

When orbitals of the same energy are available in a subshell, electrons occupy these orbitals singly first, with parallel spins, before any pairing occurs.

Example: nitrogen $2p^3$ [↑] [↑] [↑]

Example: oxygen $2p^4$ [↑↓] [↑] [↑]

3. Pauli exclusion principle

An orbital can hold a maximum of two electrons, and they must have opposite spins.

[↑↓] allowed [↑↑] not allowed

Subshell capacity reminder

Subshell	Orbitals	Maximum electrons
s	1	2
p	3	6
d	5	10
f	7	14

Anomalous behaviour reminder

The two examples highlighted in the notes are chromium and copper:

Cr: [Ar] $3d^5$ $4s^1$

Cu: [Ar] $3d^{10}$ $4s^1$

The d^5 half-filled and d^{10} completely filled arrangements are relatively stable.

11. Quick Revision Summary

Concept	Key point
Orbital	Region of space where an electron is likely to be found; maximum two electrons per orbital.
s, p, d, f capacities	2, 6, 10 and 14 electrons respectively.
Filling order	Aufbau principle: lowest-energy orbitals fill first.
Hund's rule	Orbitals of equal energy fill singly before pairing.
Pauli exclusion principle	Two electrons in the same orbital have opposite spins.
Shorthand configuration	Use the preceding noble-gas symbol in square brackets.
Transition-metal ions	Remove 4s electrons before 3d electrons.
Electron count in ions	Cation: subtract its charge from atomic number; anion: add the charge magnitude.
First ionisation energy	Energy to remove one electron from each atom in one mole of gaseous atoms.
Across a period	First ionisation energy generally increases.
Down a group	First ionisation energy generally decreases.
IE anomalies	Group 2 to 3: higher-energy p electron; Group 5 to 6: paired-electron repulsion.
Large IE jump	Indicates removal from an inner shell and helps identify outer-shell electron count.
Mass spectrum axes	x-axis: m/z; y-axis: relative abundance.
Chlorine molecular-ion peaks	70, 72, 74 in the ratio 9 : 6 : 1.
Bromine molecular-ion peaks	158, 160, 162 in the ratio 1 : 2 : 1.