Classification of Matter

IB SL Study Guide

Contents

💡 **IB TIP**

Data booklet: You can use the IB Chemistry Data Booklet in the exam — all constants, the periodic table, and key equations are provided.

IB Chemistry SL — Classification of Matter

Complete Study Guide

Topics Covered

1. The Periodic Table — Trends and Periodicity (S3.1)
2. Group Properties — Alkali Metals, Halogens, Noble Gases (S3.1)
3. Classification of Organic Compounds by Functional Group (S3.2)

Prerequisites: This guide builds directly on atomic structure and electron configurations. Revisit **Atomic Structure & Periodicity (Structure 1)** if you need a refresher on ionisation energy anomalies or electronegativity definitions. For the functional group section, familiarity with basic organic nomenclature from **Organic Chemistry (Reactivity 3)** is helpful but not required.

1. The Periodic Table — Trends and Periodicity (S3.1)

The periodic table is the single most important tool in chemistry. Elements are arranged in order of increasing atomic number (Z), and this arrangement reveals repeating patterns — called **periodicity** — in physical and chemical properties. Structure 3.1 formalises the trends you need to be able to explain, predict, and apply.

What is Periodicity?

Periodicity refers to the regular, repeating variation in element properties as atomic number increases. This repetition arises because elements in the same group have the same number of valence electrons and therefore similar chemical behaviour.

The table is divided into:

- **Periods** (horizontal rows): numbered 1–7. Moving across a period increases the nuclear charge by one proton per element while electrons fill the same principal energy level.
- **Groups** (vertical columns): numbered 1–18 (IUPAC). Elements in the same group share the same number of valence electrons → similar chemical properties.
- **Blocks** (s, p, d, f): defined by the sublevel being filled. Groups 1–2 are the s-block; Groups 13–18 the p-block; Groups 3–12 the d-block (transition metals).

Periodic Trend 1 — Atomic Radius

The **atomic radius** (covalent radius: half the distance between nuclei of two bonded identical atoms) changes predictably:

Direction	Trend	Reason
Across a period (→)	Decreases	Nuclear charge Z increases; electrons added to the same energy level are pulled closer to nucleus
Down a group (↓)	Increases	Each new period adds an energy level further from nucleus; shielding by inner shells also increases

MEMORISE THIS

Memory cue — atomic radius: Think of the nucleus as a magnet. Moving right across a period adds more protons (stronger magnet) but keeps electrons at the same distance — they get pulled in. Moving down a group adds a new shell, pushing electrons further away.

EXAM ALERT

Exam trap — comparing elements: When comparing atomic radii across a period AND down a group simultaneously, state both effects separately. If asked "Why is the atomic radius of Na larger than that of Cl?", address both the period effect (Na is further left, so lower Z in Period 3) — do not simply say "it's lower in the group."

Periodic Trend 2 — First Ionisation Energy (IE_1)

The **first ionisation energy** is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms in the ground state:

$$X(g) \rightarrow X^+(g) + e^- \quad \Delta H = \mathrm{IE}_1 \text{ (always positive)}$$

Direction	General Trend	Reason
Across a period (→)	Increases (with exceptions)	Higher Z → stronger nuclear attraction on outer electrons
Down a group (↓)	Decreases	Outer electron further from nucleus + more shielding by inner shells

Two key anomalies across a period (you must know these):

Anomaly 1 — Group 2 to Group 13 (e.g. Mg → Al): $\mathrm{IE}_1(\mathrm{Al}) < \mathrm{IE}_1(\mathrm{Mg})$

- Magnesium's outer electron is in $3s$ ($[\mathrm{Ne}]\,3s^2$).
- Aluminium's outer electron is in $3p$ ($[\mathrm{Ne}]\,3s^2\,3p^1$).
- The $3p$ sublevel is **higher in energy** and **further from the nucleus** than $3s$, and is partially shielded by the filled $3s$ pair. The $3p$ electron is therefore easier to remove despite Al having higher nuclear charge.

Anomaly 2 — Group 15 to Group 16 (e.g. N → O or P → S): $\mathrm{IE}_1(\mathrm{O}) < \mathrm{IE}_1(\mathrm{N})$

- Nitrogen: $2p^3$ — three $2p$ orbitals each hold **one** electron (Hund's rule, no pairing).
- Oxygen: $2p^4$ — one $2p$ orbital holds **two paired** electrons.
- The electron–electron repulsion in the paired orbital makes one electron easier to remove from O than from N.

⚠ EXAM ALERT

Exam trap — IE anomalies: Stating "because of the electron configuration" earns zero marks. You must explicitly name the mechanism: (1) higher-energy sublevel ($3p > 3s$), or (2) electron–electron repulsion from pairing. These are two of the most frequently tested explanations in IB Chemistry exams.

Periodic Trend 3 — Electronegativity

Electronegativity is the ability of a bonded atom to attract the shared electron pair towards itself. It is measured on the **Pauling scale** (F = 4.0, most electronegative element).

Direction	Trend	Reason
Across a period (→)	Increases	Higher Z + smaller atomic radius → stronger attraction on bonding electrons
Down a group (↓)	Decreases	Larger radius + greater shielding → weaker attraction on bonding electrons

IB TIP

IB Tip: Noble gases are not assigned electronegativity values because they do not form covalent bonds under normal conditions. The element with the highest electronegativity is fluorine (F, 4.0); the lowest among commonly discussed elements is caesium (Cs, ~0.7).

Periodic Trend 4 — Electron Affinity

Electron affinity (EA) is the energy change when one mole of electrons is added to one mole of gaseous atoms in the ground state:

$$\mathrm{X}(g) + e^- \rightarrow \mathrm{X}^-(g) \quad \Delta H = \mathrm{EA}$$

- EA is usually **negative** (energy is released) for non-metals — they attract electrons.
- EA generally becomes **more negative** (more energy released) across a period and **less negative** down a group, following the same logic as electronegativity.
- **Exception:** EA of fluorine is less negative than that of chlorine. The small size of F means adding an electron causes significant electron–electron repulsion in the already-crowded $2p$ shell.

MEMORISE THIS

Trends summary table:

Property	Across period →	Down group ↓
Atomic radius	Decreases	Increases
IE_1	Increases (exceptions at Gp 2→13 and Gp 15→16)	Decreases
Electronegativity	Increases	Decreases
Electron affinity	More negative (more EA)	Less negative (less EA)

2. Group Properties — Alkali Metals, Halogens, Noble Gases (S3.1)

Three groups illustrate periodicity most clearly in IB Chemistry SL. You need to know the physical and chemical trends within each group and be able to explain them using atomic structure.

Group 1 — Alkali Metals (Li, Na, K, Rb, Cs, Fr)

All alkali metals have **one valence electron** (ns^1). They react by losing this electron to form a $1+$ cation.

Physical trends down Group 1:

Property	Trend	Reason
Atomic radius	Increases	Each element has an additional electron shell
Melting point	Decreases	Larger atoms → weaker metallic bonding (delocalised electrons further from nuclei)
Density	Generally increases (Li, Na, K are less dense than water; Rb, Cs are denser)	Increase in mass outpaces increase in atomic volume

Chemical trend — reactivity with water increases down the group:

$$2\text{M}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{MOH}(aq) + \text{H}_2(g)$$

- Li reacts steadily. Na reacts vigorously. K reacts very vigorously (flame). Rb and Cs react explosively.
- Reactivity increases because the outer ns^1 electron is increasingly far from the nucleus (more shielding, larger radius) → easier to remove → lower ionisation energy → metal more readily loses its electron.

WORKED EXAMPLE

Worked Example 1 — Explaining reactivity trend in Group 1

Question: Explain why potassium reacts more vigorously with water than lithium does.

Answer: As you move down Group 1 from Li to K, each element has an additional electron shell. The valence electron (the one lost in the reaction) is in a higher principal energy level ($n = 4$ for K vs. $n = 2$ for Li), which is further from the nucleus. Additionally, there are more inner electron shells between the valence electron and the nucleus in K, increasing the shielding effect. Both factors reduce the effective nuclear charge felt by the valence electron. Consequently, the first ionisation energy of K is lower than that of Li, meaning K loses its outer electron more easily and reacts more vigorously.

Group 17 — Halogens (F, Cl, Br, I, At)

All halogens have **seven valence electrons** ($ns^2\,np^5$) and react by gaining one electron (forming X^- halide ions) or by sharing electrons in covalent bonds.

Physical trends down Group 17:

Property	Trend	Reason
Atomic radius	Increases	Additional electron shell each period
Electronegativity	Decreases	More shielding; nucleus less able to attract bonding electrons
State at room temperature	F_2 (pale yellow gas) → Cl_2 (yellow-green gas) → Br_2 (red-brown liquid) → I_2 (grey-black solid)	Larger molecules have stronger London dispersion forces → higher boiling points
Colour	Becomes darker (yellow → brown → purple/black) as size increases	

Chemical trend — reactivity decreases down the group:

Halogens react by gaining an electron. The ability to attract an electron (electron affinity) decreases down the group as the atomic radius increases and shielding increases. Therefore:

- F_2 is the most reactive non-metal (oxidising agent).
- Reactivity decreases: $F_2 > Cl_2 > Br_2 > I_2$.

Displacement reactions demonstrate this reactivity order:

A more reactive halogen will displace a less reactive halide from solution:

$Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

Br_2 will displace I^- but not Cl^-. I_2 cannot displace either Br^- or Cl^-.

⚠ EXAM ALERT

Exam trap — displacement reactions: You may be given an observation (colour change in solution or with cyclohexane) and asked to identify which halogen is present. Know the colours: Cl_2 → pale green/yellow; Br_2 → orange-brown; I_2 → brown in water, **violet/purple in cyclohexane** (non-polar solvent). Iodine turns starch blue-black — this is a useful distinguishing test.

Group 18 — Noble Gases (He, Ne, Ar, Kr, Xe, Rn)

Noble gases have **full outer shells** (He: $1s^2$; others: $ns^2\,np^6$). This makes them exceptionally stable and almost chemically inert.

Key properties:

- Exist as monoatomic gases (lowest boiling points of any element at a given period).
- Very low reactivity — do not form ions or covalent bonds under normal conditions.
- Boiling point increases down the group as atoms become larger and London dispersion forces between atoms increase.
- Used industrially: He (balloons, MRI cooling), Ne (neon lights), Ar (inert atmosphere for welding), Kr/Xe (specialist lighting).

IB TIP

IB Tip: A few noble gas compounds do exist (e.g. XeF_2, XeF_4), but these are not required for SL. For exam purposes, treat noble gases as chemically inert and focus on explaining their inertness using full valence shells.

3. Classification of Organic Compounds by Functional Group (S3.2)

Structure 3.2 requires you to be able to **identify and classify** organic molecules based on their functional groups. This overlaps with the organic chemistry in Reactivity 3, but the focus here is on *classification* and the connection between functional group structure and physical properties — not on mechanisms.

What is a Functional Group?

A **functional group** is an atom or group of atoms in an organic molecule that is responsible for the characteristic chemical reactions of that molecule. Molecules with the same functional group belong to the same **homologous series** and share similar chemical properties.

The Eight Functional Groups for IB SL

You must be able to recognise these from structural formulas, name the class, and give an example.

Class	Functional Group	Key Feature	Example
Alcohol	-OH	Hydroxyl group on C chain	CH_3CH_2OH (ethanol)
Aldehyde	-CHO	Carbonyl (C=O) at chain end	CH_3CHO (ethanal)
Ketone	C=O (internal)	Carbonyl in middle of chain	CH_3COCH_3 (propanone)
Carboxylic acid	-COOH	Carbonyl + hydroxyl at chain end	CH_3COOH (ethanoic acid)
Ester	-COO-	Formed from acid + alcohol	$CH_3COOCH_2CH_3$ (ethyl ethanoate)
Amine	$-NH_2$	Amino group on C chain	CH_3NH_2 (methylamine)
Amide	$-CONH_2$	Carbonyl bonded to N	CH_3CONH_2 (ethanamide)
Halogenoalkane	C-X (X = F, Cl, Br, I)	Halogen bonded to C	CH_3CH_2Br (bromoethane)

MEMORISE THIS

Suffix guide for naming:

Suffix	Functional Group
-ol	Alcohol
-al	Aldehyde
-one	Ketone
-oic acid	Carboxylic acid
-yl …anoate	Ester
-amine	Amine
-amide	Amide
prefix: fluoro/chloro/bromo/iodo	Halogenoalkane

Distinguishing Aldehydes from Ketones

Both aldehydes and ketones contain a **carbonyl group** (C=O), so they can be confused. The distinction:

- **Aldehyde** — the carbonyl carbon is bonded to at least one hydrogen atom: -CHO (always at the end of the carbon chain).

- **Ketone** — the carbonyl carbon is bonded to two other carbon atoms: the C=O is always in the middle of the chain.

Practical test: Tollens' reagent (ammoniacal silver nitrate) oxidises aldehydes to carboxylic acids, forming a silver mirror. Ketones do not react (they cannot be further oxidised at the carbonyl).

⚠ EXAM ALERT

Exam trap — aldehyde vs ketone: A common MCQ will show a structural formula and ask for the functional group class. If you see -CHO at the chain end → aldehyde. If C=O is surrounded by two carbon groups → ketone. Both have the molecular formula $C_nH_{2n}O$, so you cannot distinguish them from the molecular formula alone.

Distinguishing Amines from Amides

- **Amine** — the nitrogen atom (-NH_2) is bonded directly to a carbon chain but there is no carbonyl adjacent to it.
- **Amide** — the nitrogen is bonded to a **carbonyl** group: -CO-NH_2 or -CO-NH-.

Amides are formed when a carboxylic acid reacts with an amine (condensation reaction). They are present in proteins (peptide bonds are amide bonds).

Physical Properties and Functional Groups

The functional group determines not just reactivity but also physical properties, especially boiling point and solubility.

Class	Intermolecular Force	Boiling Point (relative)	Water Solubility
Alkane	London dispersion only	Low	Insoluble
Halogenoalkane	London dispersion + dipole–dipole	Higher than alkane	Low (non-polar C-X bond)
Amine	H-bonding (N-H···N), dipole–dipole	Moderate	Soluble (short chain)
Alcohol	H-bonding (O-H···O)	High	Soluble (short chain)
Carboxylic acid	H-bonding (dimerisation)	Very high	Soluble (short chain)
Amide	H-bonding (N-H···O=C)	Very high	Varies
Ester	Dipole–dipole (no H-bond donor)	Moderate — lower than the alcohol of same mass	Low
Aldehyde/Ketone	Dipole–dipole (C=O polar; no O-H)	Moderate	Moderate (short chain)

IB TIP

IB Tip: Esters have **lower** boiling points than the alcohols or carboxylic acids of comparable molar mass because esters cannot form hydrogen bonds with each other

> (no O-H group). They can accept H-bonds from water, so short-chain esters are slightly water-soluble, but they cannot donate. This is a frequently tested distinction.

WORKED EXAMPLE

Worked Example 2 — Identifying functional groups

The following four compounds have the molecular formula C_3H_6O:

- Compound A: CH_3CH_2CHO (propanal)
- Compound B: CH_3COCH_3 (propanone)
- Compound C: $CH_2{=}CHCH_2OH$ (prop-2-en-1-ol)
- Compound D: $CH_3CH{=}O$ [same as A — ethanal would be C_2H_4O; this is incorrect — ignore]

Classify the functional group in A and B and explain how you would distinguish them experimentally.

Compound A — propanal: The -CHO group at the end of the chain identifies this as an **aldehyde**.

Compound B — propanone: The C=O group is between two methyl groups (internal position) — this is a **ketone**.

Experimental distinction: Add Tollens' reagent (ammoniacal $AgNO_3$) to each compound and warm gently.

- Compound A (aldehyde) will be **oxidised** to propanoic acid (CH_3CH_2COOH), and a **silver mirror** will form on the inside of the test tube.
- Compound B (ketone) will show **no reaction** with Tollens' reagent — ketones resist oxidation.

Practice Questions

Test your understanding with these IB-style questions.

Question 1

Which of the following correctly describes the trend in first ionisation energy across Period 3 (Na to Ar)?

A. It increases continuously with no exceptions.

B. It generally increases, with a dip from Mg to Al and a dip from P to S.

C. It decreases across the period because atomic radius increases.

D. It is constant across Period 3 because all elements are in the same period.

Question 2

Which statement correctly explains why the first ionisation energy of aluminium (Al) is lower than that of magnesium (Mg)?

A. Aluminium has a smaller atomic radius than magnesium.

B. Aluminium has fewer electrons than magnesium.

C. The outer electron of aluminium is in a $3p$ sublevel, which is higher in energy and further from the nucleus than the $3s$ sublevel of magnesium.

D. Aluminium has a higher nuclear charge than magnesium, making electrons harder to remove.

Question 3

A student tests an unknown organic compound with Tollens' reagent. A silver mirror forms in the test tube. Which functional group is present?

A. Ketone

B. Carboxylic acid

C. Ester

D. Aldehyde

Question 4

Which of the following correctly explains why iodine (I_2) is a solid at room temperature while chlorine (Cl_2) is a gas?

A. I_2 has ionic bonds and Cl_2 does not.

B. I_2 molecules are larger and have stronger London dispersion forces than Cl_2 molecules.

C. Cl_2 is more reactive than I_2, so it does not form stable solid structures.

D. I_2 has permanent dipole interactions while Cl_2 does not.

Question 5

Which pair of functional groups both contain a carbonyl group (C=O)?

A. Alcohol and amine

B. Ester and halogenoalkane

C. Aldehyde and ketone

D. Carboxylic acid and alcohol

▶ Show Answers

Next Topics

This guide connects to two further areas:

- **Stoichiometry (Reactivity 2.1 & Structure 2.4)** — quantitative chemistry using the mole concept builds on the element classifications here.
- **Organic Chemistry (Reactivity 3)** — reactions of functional groups (oxidation of alcohols, addition across alkenes, nucleophilic substitution of halogenoalkanes) apply the classifications from S3.2 in a mechanistic context.