The Complete A Level Periodic Table Guide

for H2 Chemistry

Every trend, group and exam tip explained — free JC Chemistry revision guide

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Introduction

Ask any JC2 student which topic quietly decides their H2 Chemistry grade, and most won't say Organic Chemistry or Equilibria. They'll say the **Periodic Table** — because it's the one topic that shows up *inside* almost every other topic.

Bonding questions ask you to explain electronegativity differences. Energetics questions ask you to compare ionisation energies. Transition metal questions build directly on group trends. Even organic mechanism questions quietly rely on your understanding of electronegativity and atomic size. If your grasp of Periodic Table trends is shaky, the cracks show up across the entire H2 Chemistry syllabus — not just in one topic's questions.

That's also exactly why memorisation alone doesn't work here. Students who try to memorise "ionisation energy increases across a period" as an isolated fact tend to freeze the moment Cambridge asks a twist — why nitrogen's first ionisation energy is *higher* than oxygen's, for instance, or why the ionic radius of a transition metal ion doesn't follow the same pattern as its atomic radius. Students who understand *why* each trend happens — in terms of nuclear charge, shielding, and electron configuration — can reason their way through any variation the exam throws at them.

This guide is built to get you to that second stage: real understanding, not surface memorisation. It maps directly onto **Cambridge International A Level Chemistry (syllabus 9701)**, specifically Topic 1 (Atomic Structure) and Topic 9 (The Periodic Table: Chemical Periodicity), plus the related Group 2, Group 17 and Transition Elements topics. We'll walk through the structure of the Modern Periodic Table, explain every trend Cambridge tests at H2 level, go group-by-group through the elements you're examinable on, dive deep into Transition Metals, and connect the Periodic Table to the rest of H2 Chemistry — Bonding, Redox, Organic Chemistry, Electrochemistry, Energetics, and Equilibria. Along the way, you'll find revision tables, exam tips, common student mistakes, and a full FAQ section.

If you're a JC1 or JC2 student preparing for H2 Chemistry, or a parent researching H2 Chemistry tuition — part of Pamela's Place's wider Chemistry tuition programmes in Singapore — this is designed to be the most complete Periodic Table resource you'll find online.

> **Teacher's Advice:** Don't read this guide passively. Every time you hit a trend, pause and try to explain *why* it happens in your own words before reading our explanation. That single habit is what separates students who can only recall trends from students who can apply them under exam conditions.

> **Note: Did You Know?** Mendeleev's original 1869 Periodic Table left deliberate gaps for elements that hadn't been discovered yet — and correctly predicted their properties years before germanium, gallium and scandium were actually found. The "anomalies" H2 students now memorise (like chromium's unexpected electron configuration) are part of that same story: the Periodic Table has always been refined by exceptions, not just its neat overall pattern.

Understanding the Modern Periodic Table

Before diving into trends, you need a rock-solid mental map of how the Periodic Table is organised. Cambridge examiners frequently test this structural understanding directly, especially in MCQ and data-based questions.

Atomic Structure, Isotopes and Mass Spectrometry

Every trend in this guide is really just a story about protons, electrons, and the shells they sit in — so it's worth locking down the vocabulary before going further.

- **Atomic number (Z):** The number of protons in an atom's nucleus. This is what defines an element and determines its position on the Periodic Table.
- **Mass number (A):** The total number of protons and neutrons in an atom's nucleus.
- **Isotopes:** Atoms of the same element (same atomic number) with different numbers of neutrons, and therefore different mass numbers. Isotopes of the same element have identical chemical properties, because chemical behaviour depends on electron configuration, not neutron number.
- **Relative atomic mass (A■):** The weighted average mass of all the isotopes of an element, relative to 1/12th the mass of a carbon-12 atom, taking natural isotopic abundance into account.

Mass spectrometry is how relative atomic mass is actually determined experimentally. A sample is ionised, accelerated, deflected according to mass-to-charge ratio, and detected — producing a spectrum of peaks at each isotope's mass, with peak height proportional to relative abundance. Cambridge frequently asks you to calculate relative atomic mass from a given set of isotope masses and abundances, or to identify an element from its mass spectrum pattern — a classic combined-topic question that sits right alongside Periodic Table content in most exam papers.

> **Exam Tip:** When calculating relative atomic mass from mass spectrometry data, always weight each isotope's mass by its percentage abundance and divide by 100 — a common error is simply averaging the isotope masses without accounting for how common each one actually is.

Groups

Groups are the **vertical columns** of the Periodic Table. Elements in the same group have the same number of valence (outermost) electrons, which is why they share similar chemical properties. For H2 Chemistry, the examinable groups are:

- **Group 1** — Alkali metals (though H2 Chemistry focuses more on Group 2 and Group 17 for detailed reactions)
- **Group 2** — Alkaline earth metals (heavily examined)
- **Group 17** — Halogens (heavily examined)
- **Group 18** — Noble gases
- **Transition metals** — the d-block elements between Group 2 and Group 13

Periods

Periods are the **horizontal rows**. Moving across a period, electrons are added to the same outer shell while protons are also added to the nucleus — this single fact explains almost every periodic trend you'll learn in this guide. H2 Chemistry focuses primarily on **Period 3** (Sodium to Argon) for detailed trend analysis, since it offers a clean, complete example of a period across metals, metalloids, and non-metals.

Blocks

The Periodic Table is divided into s-block, p-block, and d-block (and f-block, not examinable at H2) based on which subshell the outermost electron occupies:

- **s-block** — Groups 1 and 2, where the outermost electron enters an s-orbital
- **p-block** — Groups 13 to 18, where the outermost electron enters a p-orbital
- **d-block** — Transition metals, where the outermost electron enters a d-orbital

Knowing which block an element belongs to helps you predict its electron configuration almost instantly, which is a huge time-saver in exams.

Metals, Non-Metals and Metalloids

- **Metals** (left and centre of the table) tend to lose electrons to form positive ions, are good conductors, and form basic oxides.
- **Non-metals** (upper right) tend to gain electrons to form negative ions, are typically poor conductors, and form acidic oxides.
- **Metalloids** (elements like silicon, along the "staircase" line) show intermediate properties — this is exactly why silicon's oxide, SiO_2, is amphoteric-adjacent in behaviour and often appears in Period 3 oxide questions.

If you're an IP or Secondary student building toward this A-Level content, our Sec 4 Chemistry Tuition guide covers the foundational Periodic Table concepts that this article builds on.

Electron Configuration in Full

Electron configuration is the map that everything else in this guide is drawn from — which block an element sits in, its reactivity, and (as you'll see in Section 5) its transition metal chemistry all trace back to how electrons fill orbitals.

The filling rules:

- **Aufbau principle:** Electrons fill the lowest-energy orbitals first (e.g. 1s before 2s, 2s before 2p).
- **Hund's Rule:** Within a subshell (e.g. the three 2p orbitals), electrons occupy separate orbitals singly, with parallel spins, before any pairing occurs — this minimises electron-electron repulsion.
- **Notation:** Configurations are written as, for example, sodium: $1s^2\ 2s^2\ 2p^6\ 3s^1$ — the superscript shows how many electrons occupy each subshell.

The two exceptions Cambridge always tests — chromium and copper:

You'd expect chromium to be [Ar] $3d^4\ 4s^2$ and copper to be [Ar] $3d^9\ 4s^2$, following the standard filling order. Instead:

- **Chromium:** [Ar] $3d^5\ 4s^1$ — a half-filled d-subshell (5 unpaired electrons) is more stable than the "expected" configuration, so one 4s electron shifts into the 3d subshell.
- **Copper:** [Ar] $3d^{10}\ 4s^1$ — a completely filled d-subshell is similarly more stable, so again one 4s electron shifts into 3d.

> **Common Mistake:** Students often write chromium and copper's configurations by simply following the standard filling order without applying these exceptions — this is one of the most reliably tested "explain the anomaly" questions in the entire Transition Metals topic, so it's worth memorising both exceptions by name rather than trying to derive them from scratch under time pressure.

> **Remember This:** Every trend in this guide traces back to just three factors — **nuclear charge, shielding, and atomic radius**. If you can reason from these three, you can rebuild any trend from scratch, even under exam pressure.

Complete Explanation of Every Periodic Trend

This is the core of H2 Chemistry Periodic Table content — and where most marks are lost or won. For each trend, we cover the pattern across a period, the pattern down a group, the underlying reason, common exceptions, and how Cambridge tests it.

Key Definitions

Before the trends themselves, get these five terms locked down — almost every trend explanation in this guide reduces to some combination of the first two.

Term	Definition
Effective nuclear charge	The net positive charge experienced by an outer electron, after accounting for shielding by inner electrons
Shielding effect	The repulsion between inner-shell electrons and outer-shell electrons, which reduces the nuclear pull felt by outer electrons
First ionisation energy	Energy needed to remove one electron from each atom in one mole of gaseous atoms
Electron affinity	Energy change when one electron is added to each atom in one mole of gaseous atoms
Electronegativity	An atom's relative ability to attract the bonding electron pair in a covalent bond

> **Remember This:** Effective nuclear charge and shielding pull in opposite directions. Across a period, effective nuclear charge rises faster than shielding — atoms hold their electrons tighter. Down a group, shielding rises faster — atoms hold their electrons more loosely. Every trend below is just this one tug-of-war playing out differently.

3.1 Atomic Radius

Across a period (left to right): Atomic radius decreases. **Down a group:** Atomic radius increases.

Why it happens: Across a period, nuclear charge increases while shielding stays roughly constant (electrons are added to the same shell), so the increased nuclear pull draws electrons closer to the nucleus. Down a group, an extra electron shell is added each time, and this increase in shielding outweighs the increase in nuclear charge, so atoms get bigger.

Common exceptions: None significant within Period 3 — this is one of the more "clean" trends, which is why Cambridge often uses it as the entry-level trend question before layering in complexity elsewhere.

How Cambridge tests it: Usually as a supporting explanation for other properties (e.g. "explain why the first ionisation energy of X is higher than Y" often requires you to reference atomic radius as part of your reasoning), or in data-based questions asking you to plot/interpret atomic radius against proton number.

3.2 Ionic Radius

Pattern: Cations (positive ions) are smaller than their parent atom; anions (negative ions) are larger than their parent atom. Across an isoelectronic series (ions with the same electron configuration, e.g. Na^+, Mg^{2+}, Al^{3+}), ionic radius decreases as nuclear charge increases.

Why it happens: Removing electrons to form a cation reduces electron-electron repulsion and, for elements that lose their entire outer shell, removes a whole shell — so the ion shrinks. Gaining electrons to form an anion increases repulsion between electrons, causing the electron cloud to expand.

Common exceptions: Students often assume ionic radius follows the same pattern as atomic radius — it doesn't always, especially once a full shell is lost. This is a classic point of confusion Cambridge deliberately probes.

How Cambridge tests it: Comparative questions like "explain why the ionic radius of Na^+ is smaller than the atomic radius of Na" or ranking ions of an isoelectronic series by size.

3.3 First Ionisation Energy

Across a period: Generally increases, with two well-known dips (Group 13 and Group 16 relative to their neighbours). **Down a group:** Decreases.

Why it happens: Across a period, increasing nuclear charge with roughly constant shielding pulls the outermost electron more tightly, requiring more energy to remove it. Down a group, the outermost electron is further from the nucleus and experiences greater shielding, so less energy is needed to remove it, despite the larger nuclear charge.

Common exceptions (the two most heavily examined anomalies in H2 Chemistry):

- **Group 13 dip** (e.g. boron lower than beryllium; aluminium lower than magnesium): the outermost electron occupies a p-orbital, which is higher in energy and less penetrating than the filled s-orbital of the previous group, making it easier to remove.
- **Group 16 dip** (e.g. oxygen lower than nitrogen; sulfur lower than phosphorus): the fourth p-electron must pair up in an already-occupied p-orbital, and the resulting electron-electron repulsion makes that electron easier to remove.

How Cambridge tests it: This is one of the most frequently examined explain-the-anomaly questions at H2 level. Expect structured questions asking you to explain why nitrogen's first ionisation energy exceeds oxygen's, or why aluminium's is lower than magnesium's — always in terms of electron configuration and subshell arrangement, never "because it's an exception."

3.4 Successive Ionisation Energy

Pattern: Each successive ionisation energy for a given atom is higher than the last, with large jumps occurring when a new (inner, more tightly held) electron shell is broken into.

Why it happens: Removing an electron reduces electron-electron repulsion for the remaining electrons, pulling them closer to the nucleus and making the next electron harder to remove. A big jump signals that the next electron comes from a shell closer to the nucleus, with less shielding and higher effective nuclear charge.

How Cambridge tests it: Successive ionisation energy data is a classic way to *deduce* the group number of an unknown element — look for where the big jump occurs, and count backwards to identify how many electrons were in the outermost shell. This is one of the most reliable "free marks" in H2 Chemistry if you know the method.

> **Exam Tip:** When given a table of successive ionisation energies, don't try to memorise "typical" jump sizes — instead, calculate the ratio between consecutive values. A jump of roughly 2–3× is normal within a shell; a jump of 5× or more usually signals a new shell.

3.5 Electron Affinity

Definition: The energy change when one mole of gaseous atoms gains one electron each to form one mole of gaseous 1- ions.

Pattern: Generally becomes more negative (more energy released) across a period, though this trend is less consistently examined at H2 level than ionisation energy.

Why it happens: Increasing nuclear charge across a period attracts the incoming electron more strongly, releasing more energy. However, elements with a stable half-filled or filled subshell (like Group 15 or Group 18) resist gaining an extra electron, disrupting the trend.

How Cambridge tests it: Usually in the context of comparing halogens (Group 17), where electron affinity becomes less negative down the group as atomic radius increases and shielding reduces the attraction on the incoming electron.

3.6 Electronegativity

Definition: A measure of an atom's ability to attract the bonding pair of electrons in a covalent bond.

Across a period: Increases. **Down a group:** Decreases.

Why it happens: Same logic as ionisation energy — increasing nuclear charge with constant shielding across a period pulls bonding electrons more strongly; increasing atomic radius and shielding down a group weakens that pull.

How Cambridge tests it: Electronegativity is less often tested as a standalone trend and far more often used as the *reasoning tool* for other questions — explaining bond polarity, predicting the direction of dipoles, explaining why HF has an unusually high boiling point (hydrogen bonding), or explaining oxide/chloride behaviour across Period 3. This is the trend most directly connected to H2 Chemistry Bonding topic — see Section 6.

3.7 Metallic Character

Pattern: Decreases across a period, increases down a group — essentially the mirror image of electronegativity and ionisation energy trends.

Why it happens: Metallic character depends on how readily an element loses electrons. As ionisation energy rises across a period, electrons are held more tightly, so metallic character falls. Down a group, ionisation energy falls, so metallic character rises.

3.8 Reactivity

Reactivity trends differ by group, which is exactly why Cambridge tests them group-by-group rather than as a single universal rule:

- **Group 2 metals** become *more reactive* down the group (easier to lose electrons as ionisation energy falls).
- **Group 17 halogens** become *less reactive* down the group (harder to gain electrons as atomic radius increases and electron affinity becomes less negative).

We cover the reactions behind these trends in detail in Section 4.

Periodic Trends Summary Table

Trend	Across a Period	Down a Group	Key Reason
Atomic Radius	Decreases	Increases	Nuclear charge ↑ (period) / extra shell (group)
Ionic Radius (isoelectronic)	Decreases	Increases	Nuclear charge changes electron pull
First Ionisation Energy	Increases (with Gp13 & Gp16 dips)	Decreases	Nuclear charge vs. shielding/distance
Electron Affinity	More negative (general)	Less negative	Nuclear charge vs. shielding
Electronegativity	Increases	Decreases	Nuclear charge vs. shielding/distance
Metallic Character	Decreases	Increases	Ease of losing electrons
Melting Point (Period 3)	Rises then falls sharply	—	Bonding type changes: metallic → giant covalent → simple molecular

Group-by-Group Guide

Group 1 — Alkali Metals (background knowledge)

Group 1 metals (lithium, sodium, potassium) are soft, low-density metals that react vigorously with water and oxygen. While Group 1 detailed reactions are more of a lower-secondary topic, H2 Chemistry expects you to recognise Group 1 as the benchmark for "very high reactivity, very low ionisation energy" when comparing across groups or explaining periodic trends.

Group 2 — Alkaline Earth Metals (heavily examined at H2)

Properties: Group 2 elements (beryllium, magnesium, calcium, strontium, barium) have two valence electrons and form 2+ ions. Reactivity increases down the group.

Key reactions:

- **With water:** Reactivity increases down the group. Mg reacts very slowly with cold water but readily with steam; Ca, Sr and Ba react increasingly vigorously with cold water, producing hydrogen gas and a metal hydroxide.
- **With oxygen:** All Group 2 metals burn to form the metal oxide (e.g. $2Mg + O_2 \rightarrow 2MgO$).
- **Solubility trend of hydroxides:** *Increases* down the group ($Mg(OH)_2$ is only sparingly soluble; $Ba(OH)_2$ is fairly soluble) — this is a frequently tested anomaly students confuse with the sulfate trend below.
- **Solubility trend of sulfates:** *Decreases* down the group ($MgSO_4$ is soluble; $BaSO_4$ is virtually insoluble) — this is the basis of the classic "identify the unknown Group 2 sulfate" qualitative analysis question.

Exam facts: Thermal stability of Group 2 carbonates and nitrates *increases* down the group, because larger cations polarise the carbonate/nitrate anion less, making the compound more resistant to heat-induced decomposition.

Frequently tested question: "Explain why the solubility of Group 2 hydroxides increases down the group while the solubility of their sulfates decreases." This requires you to compare lattice energy and hydration energy trends — a favourite Cambridge structured question.

Group 17 — Halogens (heavily examined at H2)

Properties: Halogens (fluorine, chlorine, bromine, iodine) exist as diatomic molecules and are strong oxidising agents, with oxidising power decreasing down the group.

Key reactions:

- **Displacement reactions:** A more reactive halogen displaces a less reactive halide from solution (e.g. chlorine displaces bromide to form bromine). This is the standard practical/observation question — expect to describe colour changes (e.g. colourless to orange when Cl_2 displaces Br^-).
- **With hydrogen:** All halogens react with hydrogen to form hydrogen halides, with reactivity (and reaction violence) decreasing down the group.
- **Disproportionation:** Chlorine famously disproportionates in cold, dilute NaOH to form NaCl and NaOCl (used in bleach) — a classic redox/Periodic Table crossover question (see Section 6).

Exam facts: Oxidising power decreases down the group because atomic radius increases and shielding increases, so the incoming electron is less strongly attracted — directly tying back to the electron affinity trend in Section 3.

Frequently tested question: "Describe and explain the trend in oxidising power of the halogens down the group," or "explain the observations when chlorine water is added to potassium bromide solution."

Group 18 — Noble Gases

Noble gases (helium, neon, argon) have full outer shells, giving them very high ionisation energies and extremely low reactivity. At H2 level, they're mostly examined as the reference point for "maximum" ionisation energy within a period, and occasionally in the context of London dispersion forces and boiling point trends.

Period 3 Oxides: The Acid-Base Trend

This is one of the most heavily tested standalone topics tied to periodicity, and it's where "across a period" reasoning gets applied to a completely different property: acid-base character.

Oxide	Bonding	Reaction with Water	Resulting pH
Na_2O	Ionic	Reacts readily to form NaOH	Strongly basic (pH ~13–14)
MgO	Ionic	Reacts slowly, only slightly soluble	Weakly basic (pH ~9–10)
Al_2O_3	Ionic/covalent (intermediate)	Insoluble — does not react with water	Amphoteric (reacts with both acids and bases)
SiO_2	Giant covalent	Insoluble — does not react with water	Neutral (no reaction)
P_4O_{10}	Simple covalent	Reacts vigorously to form H_3PO_4	Strongly acidic (pH ~1–2)
SO_2 / SO_3	Simple covalent	React readily to form H_2SO_3 / H_2SO_4	Strongly acidic (pH ~1–3)

Why it happens: As you move across Period 3, bonding shifts from ionic (Na, Mg) through intermediate (Al) to giant covalent (Si) and simple molecular covalent (P, S). Ionic oxides contain a discrete O^{2-} ion that readily accepts protons from water, producing a basic solution. Covalent oxides instead undergo hydrolysis reactions that release H^+ ions, producing an acidic solution. Al_2O_3 sits at the boundary and reacts with both acids and bases, making it amphoteric.

Exam facts: The oxide most commonly asked to be identified as amphoteric is Al_2O_3 — know both its reaction with acid (forming Al^{3+} salts) and with alkali (forming aluminate ions) to answer this fully.

Period 3 Chlorides: Reactions with Water

Chloride	Bonding	Reaction with Water	Resulting pH
NaCl	Ionic	Dissolves, no hydrolysis	Neutral (pH ~7)
$MgCl_2$	Ionic	Dissolves, slight hydrolysis	Weakly acidic (pH ~6)
$AlCl_3$	Covalent (polarised)	Hydrolyses vigorously, often with steamy fumes of HCl	Acidic (pH ~3)
$SiCl_4$	Simple covalent	Hydrolyses violently	Strongly acidic (pH ~2)
PCl_3 / PCl_5	Simple covalent	Hydrolyse readily, releasing HCl fumes	Strongly acidic (pH ~1–2)

Why it happens: Ionic chlorides (Na, Mg) simply dissolve, with only Mg^{2+}'s higher charge density causing mild hydrolysis. From Al onward, the chlorides are covalent and polarised enough that water molecules attack the electron-deficient central atom directly, breaking the Cl bonds and releasing HCl — which is exactly why $AlCl_3$, $SiCl_4$, PCl_3 and PCl_5 all "fume" in moist air.

> **Exam Tip:** Cambridge frequently pairs the oxides and chlorides trends in the same structured question, asking you to explain *both* the acid-base pattern and the hydrolysis pattern using the same underlying idea — the shift from ionic to covalent bonding across the period. Answer both halves with that single unifying reason rather than treating them as two separate facts to memorise.

Transition Metals

Covered in full depth in Section 5, but as a group-level summary: transition metals sit between Group 2 and Group 13, are characterised by an incomplete d-subshell (in at least one oxidation state), and show variable oxidation states, coloured compounds, and catalytic activity — none of which apply to Group 1, 2, or 17 elements.

Group Comparison Table

Group	Valence Electrons	Reactivity Trend Down Group	Typical Ion Formed	Key H2 Reaction Type
Group 1	1	Increases	1+	Reaction with water
Group 2	2	Increases	2+	Reaction with water; solubility trends
Group 17	7	Decreases	1–	Displacement; disproportionation
Group 18	8 (full shell)	Negligible	None (inert)	Reference point for IE trends

Group	Valence Electrons	Reactivity Trend Down Group	Typical Ion Formed	Key H2 Reaction Type
Transition Metals	Variable (d-subshell)	N/A	Variable (e.g. 2+, 3+)	Complex formation, redox, catalysis

Transition Metals Deep Dive

Transition metals are one of the most content-rich sub-topics tied to the Periodic Table at H2 level, and Cambridge examines them in real depth.

d-Block vs. Transition Metal: The Formal Definition

These two terms are not interchangeable, and Cambridge specifically tests the distinction. A **d-block element** is simply any element whose highest-energy electron occupies a d-orbital — this includes scandium and zinc. A **transition metal** is defined more strictly: an element that forms at least one stable ion with a partially filled d-subshell.

This is exactly why **scandium and zinc are d-block elements but not transition metals**:

- Scandium only forms Sc^{3+}, which has an *empty* d-subshell ([Ar] configuration).
- Zinc only forms Zn^{2+}, which has a *completely full* d-subshell ([Ar] $3d^{10}$).

Neither ion has a partially filled d-subshell, so neither qualifies as a transition metal by the formal definition — which is also precisely why their compounds are colourless (see "Colours" below).

> **Exam Tip:** If a question asks you to "explain why scandium is not classified as a transition metal," the formal definition above *is* the answer — don't just say "it doesn't have a partially filled d-subshell" without stating that this is the actual defining criterion for the term "transition metal."

Variable Oxidation States

Because transition metals have both 4s and 3d electrons available for bonding (with similar energy levels), they can lose different numbers of electrons depending on the reaction conditions, giving rise to multiple stable oxidation states. Iron, for example, commonly exists as Fe^{2+} and Fe^{3+}; manganese can range from +2 all the way to +7 (as in MnO_4^-).

> **Common Mistake:** When transition metals form ions, the 4s electrons are always lost *before* any 3d electrons, regardless of the order they were filled in. For example, iron's atomic configuration is [Ar] $3d^6$ $4s^2$, but Fe^{2+} is [Ar] $3d^6$ (not $3d^4$ $4s^2$) — students who forget this rule frequently write the wrong configuration for transition metal ions in exams.

Complex Ions

A complex ion forms when a central transition metal ion is surrounded by ligands bonded through **dative (coordinate) covalent bonds**. Common examples examined at H2 include $[Cu(H_2O)_6]^{2+}$, $[Cu(NH_3)_4(H_2O)_2]^{2+}$, and $[CuCl_4]^{2-}$ — the classic ligand-substitution series used to test colour change and shape reasoning.

Complex ion geometry depends on the ligand's size and the coordination number (the number of dative bonds to the central ion):

Coordination Number	Geometry	Bond Angle	Example
6	Octahedral	90°	$[Cu(H_2O)_6]^{2+}$
4 (small ligands)	Tetrahedral	109.5°	$[CuCl_4]^{2-}$
4 (larger/flat ligands)	Square planar	90°	$[Ni(CN)_4]^{2-}$

Six small ligands (like H_2O or NH_3) typically pack around the central ion octahedrally. Four ligands usually give a tetrahedral arrangement, *except* when the ligands are large — like Cl^- — where steric considerations can favour square planar geometry in specific cases. This is why the classic ligand-substitution series $[Cu(H_2O)_6]^{2+} \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} \rightarrow [CuCl_4]^{2-}$ involves both a colour change *and* a shape change (octahedral to tetrahedral), which Cambridge often asks you to describe together.

Colours

The characteristic colours of transition metal compounds arise because partially filled d-orbitals split into two energy levels in the presence of ligands. Electrons can absorb visible light to jump between these split d-orbitals, and the colour we see is the complementary colour of the light absorbed. This is why Sc^{3+} and Zn^{2+} compounds are colourless — with an empty or completely full d-subshell respectively, there's no d-d electron transition possible (and, as covered above, this is also precisely why neither is classified as a transition metal).

Catalysts

Transition metals make effective catalysts because of their variable oxidation states (allowing them to form intermediate compounds that lower activation energy) and their ability to adsorb reactant molecules onto their surface via available d-orbitals. Classic H2 examples include iron in the Haber process and vanadium(V) oxide in the Contact process — both of which link Transition Metals directly to Chemical Energetics and industrial chemistry questions.

Heterogeneous catalysis occurs when the catalyst is in a different physical state from the reactants — most commonly a solid catalyst speeding up a gas-phase reaction, such as iron in the Haber process or vanadium(V) oxide in the Contact process. The reaction proceeds via adsorption of reactants onto the catalyst's surface, weakening bonds and lowering activation energy, before the products desorb.

Homogeneous catalysis occurs when the catalyst is in the *same* physical state as the reactants, typically via an intermediate species. A classic H2 example is Fe^{2+}/Fe^{3+} catalysing the reaction between iodide and peroxodisulfate ions in aqueous solution — the catalyst is oxidised then reduced back to its original state over the course of the reaction, so it doesn't appear in the overall equation despite being essential to the mechanism.

Autocatalysis is a special case where one of the *products* of a reaction acts as a catalyst for that same reaction. The textbook H2 example is the reaction between manganate(VII) ions (MnO_4^-) and ethanedioate ions ($C_2O_4^{2-}$): the reaction starts slowly, but as Mn^{2+} product accumulates, it begins catalysing the reaction itself, causing the rate to speed up mid-reaction before eventually slowing again as the reactants are used up. This produces a distinctive S-shaped rate curve that Cambridge frequently asks you to sketch and explain.

> **Exam Tip:** If asked to explain an unusually shaped rate-time graph where the reaction *speeds up* partway through (rather than the usual steadily-decreasing rate), autocatalysis is almost always the expected answer — check whether one of the named products could plausibly act as a catalyst for the reaction.

Ligands

A ligand is a molecule or ion that donates a lone pair of electrons to form a dative bond with a central metal ion. H2 Chemistry expects you to distinguish monodentate ligands (e.g. H_2O, NH_3, Cl^-) from bidentate ligands, and to understand ligand substitution reactions and the accompanying colour and shape changes.

Industrial Uses

Beyond catalysis, transition metals and their compounds appear across H2 syllabus applications — from iron and its oxidation states in redox titrations, to manganate(VII) as a common oxidising agent in volumetric analysis questions.

> **Common Mistake:** Students often say a compound is coloured "because it's a transition metal," without explaining the actual mechanism (partially filled d-orbitals + ligand-induced splitting + d-d electron transitions). Cambridge specifically penalises answers that skip this reasoning chain.

H2 Chemistry Applications

The Periodic Table isn't a standalone topic — it's the connective tissue across H2 Chemistry. Here's exactly how it links to other major topics, which is also where most "combined topic" exam questions come from.

Chemical Bonding: Electronegativity differences (Section 3.6) directly determine whether a bond is non-polar covalent, polar covalent, or ionic, and explain molecular polarity and intermolecular forces like hydrogen bonding.

Redox: Oxidising and reducing power trends across Group 17 and Transition Metals (Sections 4 and 5) are the foundation of redox half-equations, disproportionation reactions, and electrode potential comparisons.

Organic Chemistry: Electronegativity differences between carbon and attached atoms (like halogens or oxygen) determine bond polarity, which in turn explains nucleophilic and electrophilic attack in reaction mechanisms — this is why a shaky grasp of electronegativity often causes downstream struggles in Organic mechanisms. For a deeper walkthrough, see our guide on Understanding Organic Chemistry for A Levels.

Electrochemistry: Standard electrode potentials are essentially a quantitative expression of a species' tendency to be reduced — directly connected to ionisation energy and electronegativity trends across the Periodic Table.

Energetics: Lattice energy trends (used to explain Group 2 solubility patterns in Section 4) and bond energy trends (linked to electronegativity and atomic radius) both draw directly on Periodic Table reasoning.

Equilibria: Less directly connected, but the acid-base character of Period 3 oxides (basic → amphoteric → acidic across the period, covered in full in "Period 3 Oxides: The Acid-Base Trend" in Section 4) is frequently tested alongside acid-base equilibria concepts.

Frequently Tested Exam Questions

Cambridge H2 Chemistry Periodic Table questions tend to fall into five recognisable formats:

1. Predicting trends: "Predict and explain the trend in [property] across Period 3 / down Group 2." Always structure your answer as: state the trend → explain in terms of nuclear charge/shielding/radius → name the exception if relevant.

2. Comparing elements: "Explain why the first ionisation energy of magnesium is higher than that of aluminium." Requires precise electron configuration reasoning ($3s^2$ vs. $3s^2 3p^1$), not a general trend statement.

3. Explaining anomalies: "Explain why sulfur has a lower first ionisation energy than phosphorus." Always reference electron pairing within the same subshell and the resulting repulsion.

4. Data-based questions: You'll be given a table or graph of ionisation energies, radii, or melting points and asked to identify an element, deduce its group, or explain a specific data point. Practise reading successive ionisation energy jumps (Section 3.4) — this question type appears almost every year.

5. Structured questions combining topics: E.g. "Using your knowledge of Group 2 chemistry and energetics, explain why the solubility of Group 2 sulfates decreases down the group." These require you to draw on two topics simultaneously, which is exactly why understanding the connections in Section 6 matters more than memorising isolated facts.

> **Exam Tip:** Whenever a question asks you to "explain," Cambridge is checking for a causal chain, not a fact. The strongest answers always follow: **observation** → **underlying factor (nuclear charge/shielding/radius)** → **consequence.** Skipping the middle step is the single most common reason for lost marks on Periodic Table questions.

For more on structuring high-scoring answers across all five question types above, see our guides on Top 10 Chemistry Exam Techniques and Understanding the Chemistry Marking Scheme, which explain exactly how Cambridge allocates marks for causal-chain reasoning like this.

Worked Exam Answers

Knowing the question formats is only half the battle — here's exactly how a full-mark answer is built for three of the most commonly tested Periodic Table questions, with the mark allocation shown.

Worked Example 1: Comparing Elements (3 marks)

Question: Explain why the first ionisation energy of magnesium is higher than that of aluminium.

Model answer:

> **Note:** Magnesium's outermost electron is in the 3s subshell, while aluminium's outermost electron is in the 3p subshell *(1 mark — correct subshell identification)*. The 3p subshell is at a higher energy level than the 3s subshell and is less penetrating, meaning the 3p electron is, on average, further from the nucleus and experiences less effective nuclear charge *(1 mark — reasoning)*. As a result, less energy is needed to remove aluminium's outermost electron, giving it a lower first ionisation energy than magnesium *(1 mark — conclusion linked back to the question)*.

Why this scores full marks: It names the exact subshells (not just "shielding"), explains the energy/penetration difference, and closes by explicitly answering the "why is Mg higher" question rather than trailing off after the mechanism.

Worked Example 2: Explaining an Anomaly (3 marks)

Question: Explain why sulfur has a lower first ionisation energy than phosphorus.

Model answer:

> **Note:** Phosphorus has the electron configuration [Ne] $3s^2$ $3p^3$, with each of the three 3p orbitals singly occupied *(1 mark — configuration)*. Sulfur has the configuration [Ne] $3s^2$ $3p^4$, meaning its fourth 3p electron must pair up with an existing electron in the same orbital *(1 mark — pairing identified)*. The repulsion between these two paired electrons in the same orbital makes it easier to remove one of them, so sulfur's first ionisation energy is lower than phosphorus's, despite sulfur having a higher nuclear charge *(1 mark — conclusion, explicitly acknowledging the "despite higher nuclear charge" twist)*.

Why this scores full marks: It states both configurations explicitly rather than describing them vaguely, and the final sentence directly addresses why the "expected" trend (higher nuclear charge → higher IE) doesn't hold here — this is exactly the kind of anomaly reasoning Cambridge rewards.

Worked Example 3: Structured, Combined-Topic Question (4 marks)

Question: Using your knowledge of Group 2 chemistry and energetics, explain why the solubility of Group 2 sulfates decreases down the group.

Model answer:

> **Note:** Going down Group 2, the ionic radius of the cation increases *(1 mark)*. Both lattice energy and hydration energy become less negative (less exothermic) as ionic radius increases, but lattice energy decreases in magnitude more slowly than hydration energy, because the sulfate anion is large and its contribution to lattice energy is less sensitive to the cation's size *(1 mark — comparative reasoning)*. This means the enthalpy change of solution becomes increasingly endothermic down the group *(1 mark — link to energetics)*, making dissolving less energetically favourable and explaining the decrease in solubility down the group *(1 mark — conclusion tied back to the question)*.

Why this scores full marks: It draws on two topics (Group 2 chemistry *and* energetics, as the question explicitly asks), and every sentence builds toward the final conclusion rather than listing facts in isolation.

> **Teacher's Advice:** Notice that all three model answers follow the same shape: state what changes → explain the underlying reason → draw the explicit conclusion. If you're ever unsure how to structure an explain question under time pressure, default to this three-step shape — it maps directly onto how Cambridge allocates marks.

Common Student Mistakes

Based on patterns we see across O-Level, IP and A-Level Pure Chemistry and Biology students, here are the mistakes that cost the most marks on Periodic Table questions:

1. Stating a trend without explaining the *reason* behind it.
2. Confusing "shielding" with "nuclear charge" when explaining ionisation energy changes.
3. Forgetting the Group 13 and Group 16 dips in the first ionisation energy trend across a period.
4. Assuming ionic radius follows the exact same pattern as atomic radius.
5. Mixing up the Group 2 hydroxide solubility trend (increases down group) with the sulfate solubility trend (decreases down group).
6. Explaining transition metal colour without mentioning d-orbital splitting.
7. Assuming all transition metal ions are coloured (forgetting Sc^{3+} and Zn^{2+} are colourless).
8. Confusing electronegativity with electron affinity — they are related but not interchangeable terms.
9. Misreading successive ionisation energy graphs and miscounting the group number.
10. Forgetting that reactivity trends run in *opposite* directions for Group 2 (increases down) versus Group 17 (decreases down).
11. Writing "because it's more reactive" as an explanation, instead of explaining *why* it's more reactive.
12. Failing to link electronegativity to bond polarity when answering Organic Chemistry mechanism questions.
13. Not using precise electron configuration notation (e.g. $3s^2 3p^1$) when comparing specific elements.
14. Treating "metallic character" and "reactivity" as always meaning the same thing across every group.
15. Rushing data-based questions without first identifying whether the pattern is "across a period" or "down a group" — the direction changes the entire explanation.

These sit alongside a broader set of pitfalls we cover in Common Chemistry Mistakes Students Make, worth reviewing before your next exam.

Study Strategy

What to memorise: Electron configurations for Period 3 elements, the exact wording of definitions (first ionisation energy, electronegativity, electron affinity), and the two key anomalies in the ionisation energy trend (Group 13 and Group 16 dips).

What to understand (not memorise): The causal reasoning behind every trend — nuclear charge, shielding, and atomic radius. If you understand these three factors deeply, you can *derive* any trend on the spot rather than recalling it from memory, which is far more reliable under exam pressure.

Revision techniques:

- Practise writing full causal-chain explanations (observation → factor → consequence) for every trend until it becomes automatic.
- Use past-year data-based questions to practise reading ionisation energy graphs quickly.
- Create your own comparison tables between similar-looking trends (e.g. Group 2 vs. Group 17 reactivity) to force yourself to spot the differences.
- Test yourself by predicting a trend for an *unfamiliar* element or hypothetical group — this checks true understanding rather than memorised patterns.

Last-minute revision tips: In the final days before your exam, focus your limited time on the two most heavily tested areas — first ionisation energy anomalies (Section 3.3) and Group 2/Group 17 reaction trends (Section 4) — since these appear in some form almost every year.

For a broader revision framework that applies this same approach across the whole H2 syllabus, see our H2 A-Level Chemistry Revision Guide and Chemistry Revision Strategy article.

Printable Revision Summary

A single-page, print-ready summary of every trend, group, exception and colour covered in this guide is also available as Pamela's Place's free "A Level Periodic Table Quick Revision Card" — perfect for your desk or wall during revision.

The same information is also available as web tables below, for quick on-screen reference.

Trends Summary

Trend	Across Period	Down Group
Atomic Radius	↓	↑
Ionisation Energy	↑ (Gp13, Gp16 dips)	↓
Electronegativity	↑	↓
Metallic Character	↓	↑

Groups Summary

Group	Reactivity Direction	Notable Behaviour
Group 2	Increases down group	Hydroxide solubility ↑, sulfate solubility ↓
Group 17	Decreases down group	Displacement reactions, disproportionation

Exceptions Summary

Exception	Where It Occurs	Reason
Group 13 IE dip	Between Group 2 and Group 13	p-orbital electron easier to remove than s
Group 16 IE dip	Between Group 15 and Group 16	Paired p-electron repulsion

Transition Metal Colours (illustrative examples)

Ion	Typical Colour	Reason for Colour (or lack of)
$[Cu(H_2O)_6]^{2+}$	Blue	d-d electron transition
$[Cu(NH_3)_4(H_2O)_2]^{2+}$	Deep blue	Ligand substitution changes splitting energy
Sc^{3+}	Colourless	Empty d-subshell — no transition possible
Zn^{2+}	Colourless	Full d-subshell — no transition possible

Reactivity Quick Reference

Group	More Reactive Direction
Group 2	Down the group
Group 17	Up the group

Quick Self-Test

Test yourself before checking the answer key below each question — this is a far more reliable way to gauge your actual exam readiness than just re-reading notes.

1. State and explain the trend in atomic radius across Period 3.

2. Why does aluminium have a lower first ionisation energy than magnesium?

3. Why does sulfur have a lower first ionisation energy than phosphorus?

4. Explain why the solubility of Group 2 hydroxides increases down the group.

5. Explain why the solubility of Group 2 sulfates decreases down the group.

6. Describe the trend in oxidising power of the halogens down Group 17, and explain why it occurs.

7. Write the electron configuration of a chromium atom, and explain why it doesn't follow the standard filling order.

8. Why is Sc^{3+} colourless while $[Cu(H_2O)_6]^{2+}$ is blue?

9. Explain why $AlCl_3$ produces steamy fumes of HCl when exposed to moist air, while NaCl does not.

10. A Group 2 metal, X, has successive ionisation energies (in kJ mol^{-1}) of: 590, 1145, 4912, 6491. Deduce which group X belongs to, and explain your reasoning.

Answer Key

Try each question yourself before checking below.

1. Atomic radius decreases across Period 3 — nuclear charge increases while shielding stays roughly constant (electrons added to the same outer shell), so the increased nuclear pull draws electrons closer to the nucleus.

2. Aluminium's outermost electron is in the 3p subshell, which is higher in energy and less penetrating than magnesium's filled 3s subshell, making it easier to remove.

3. Sulfur's fourth 3p electron must pair with an existing electron in the same orbital; the resulting repulsion makes it easier to remove than phosphorus's unpaired $3p^3$ electrons.

4. As ionic radius increases down the group, hydration energy decreases in magnitude faster than lattice energy for the hydroxide ion, making dissolving more energetically favourable.

5. As ionic radius increases down the group, lattice energy decreases in magnitude more slowly than hydration energy for the larger sulfate ion, making dissolving less energetically favourable.

6. Oxidising power decreases down the group — atomic radius increases and shielding increases, so the incoming electron is less strongly attracted to the nucleus, and electron affinity becomes less negative.

7. [Ar] $3d^5$ $4s^1$ — a half-filled d-subshell is more stable than the "expected" [Ar] $3d^4$ $4s^2$, so one 4s electron shifts into 3d.

8. Sc^{3+} has an empty d-subshell, so no d-d electron transition is possible; $[Cu(H_2O)_6]^{2+}$ has a partially filled d-subshell that splits in the presence of ligands, allowing a d-d transition that absorbs visible light.

9. $AlCl_3$ is covalent and polarised, so water molecules attack the electron-deficient aluminium directly, hydrolysing the bond and releasing HCl; NaCl is purely ionic and simply dissolves without hydrolysis.

10. There's a large jump between the 2nd (1145) and 3rd (4912) ionisation energies, indicating the third electron comes from a new, more tightly held shell. This means X has 2 valence electrons, so X is a Group 2 element.

A Level Periodic Table FAQs

Q: What is the most important Periodic Table trend for H2 Chemistry? First ionisation energy is arguably the most heavily examined trend, because it connects directly to electron configuration, Group 2 and Group 17 chemistry, and Transition Metals reasoning.

Q: Why does ionisation energy dip at Group 13 and Group 16? The Group 13 dip occurs because the outermost electron enters a higher-energy p-orbital that's easier to remove than a filled s-orbital. The Group 16 dip occurs because the fourth p-electron must pair with another electron in the same orbital, and the resulting repulsion makes it easier to remove.

Q: Is the Periodic Table tested in Paper 1, Paper 2, or Paper 3? Periodic Table concepts appear across all papers — as MCQs testing trend recall in Paper 1, as structured explanation questions in Paper 2, and embedded within longer Organic, Energetics, or Redox questions in Paper 3.

Q: Why do Group 2 hydroxides become more soluble down the group while Group 2 sulfates become less soluble? This comes down to how lattice energy and hydration energy change with ionic size, and the two anions (hydroxide vs. sulfate) respond differently — a detailed explanation is covered in Section 4 of this guide.

Q: Why are some transition metal ions colourless? Ions like Sc^{3+} and Zn^{2+} have either a completely empty or completely full d-subshell, so there's no d-d electron transition available to absorb visible light — meaning no colour is produced.

Q: What's the difference between electronegativity and electron affinity? Electron affinity is the energy change when a gaseous atom gains an electron to form a gaseous ion (a measurable, defined energy value). Electronegativity is a relative, dimensionless measure of how strongly an atom attracts a shared pair of electrons within a covalent bond. They're related concepts but are not interchangeable in exam answers.

Q: How do I identify an unknown element from successive ionisation energy data? Look for the biggest jump between consecutive ionisation energies — this signals a new, more tightly held electron shell. Count how many ionisation energies come before that jump; that number equals the number of valence electrons, which tells you the group number.

Q: Why does reactivity increase down Group 2 but decrease down Group 17? Group 2 reactivity depends on how easily electrons are *lost*, which becomes easier down the group as ionisation energy falls. Group 17 reactivity depends on how easily electrons are *gained*, which becomes harder down the group as atomic radius increases and electron affinity becomes less negative.

Q: Does H2 Chemistry examine the whole Periodic Table or just certain groups? H2 Chemistry focuses on Period 3 for general trends, plus Group 2, Group 17, and the Transition Metals for detailed group chemistry. Groups 1 and 18 are mostly background/reference knowledge rather than deeply examined groups.

Q: What's the biggest reason students lose marks on Periodic Table questions? Stating the trend correctly but failing to explain the underlying reason (nuclear charge, shielding, atomic radius) — Cambridge marking schemes consistently award marks for the explanation, not just the observation.

Q: How is the Periodic Table connected to Organic Chemistry in H2 syllabus? Electronegativity differences between carbon and attached atoms determine bond polarity, which directly explains why nucleophiles and electrophiles attack specific positions in a molecule during reaction mechanisms. Our Understanding Organic Chemistry for A Levels guide covers this connection in more depth.

Q: Is memorising the Periodic Table trends enough to score well? No — Cambridge frequently tests unfamiliar variations (unusual elements, hypothetical comparisons, data-based deductions) that require genuine understanding of *why* trends occur, not just recall of the trend itself. See our Top 10 Chemistry Exam Techniques for ways to build that deeper understanding.

Q: Where can I get structured help with H2 Chemistry Periodic Table topics? Pamela's Place offers H2 Chemistry tuition in Singapore, with small Omakase Small Group classes capped at 4–6 students, designed to build exactly this kind of causal, exam-ready understanding rather than rote memorisation.

Conclusion

The A Level Periodic Table isn't a topic you memorise once and move past — it's a foundation that resurfaces in Bonding, Redox, Organic Chemistry, Electrochemistry, Energetics, and Equilibria questions throughout JC1 and JC2. Students who understand the *why* behind each trend — nuclear charge, shielding, and atomic radius — consistently outperform students who only memorise the *what*, especially when Cambridge throws an unfamiliar twist into a question.

If you've read this far, you already have a stronger grasp of Periodic Table reasoning than most students walking into their H2 Chemistry exam. But applying that understanding under timed conditions, across genuinely difficult past-year questions, is a different skill — and it's exactly what structured guidance helps with.

At Pamela's Place, our H2 Chemistry tuition uses small **Omakase Small Group** classes, capped at just 4–6 students, so every student gets the kind of direct, causal-reasoning coaching this guide has walked you through — not passive lecture-style teaching. Our tutor brings 12+ years of experience and has taught 500+ students, with over 80% achieving distinction or B grades — you can read real student results and what our students say about their experience. We're conveniently located at 118 Upper Bukit Timah Rd, The LINQ @ Beauty World, right by Beauty World MRT.

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