MEGA LECTURE

IB DP CHEMISTRY

Structure 3: Classification of Matter

Structure 3.1 The Periodic Table

Classification of Elements

Revision Notes · Standard and Higher Level

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What the syllabus requires

Structure 3.1 asks you to see the periodic table not as a chart to memorise but as a consequence of electron configuration. Use this checklist before the exam.

Understanding	You should be able to...
Layout of the table	Relate an element's period and group to the number of occupied electron shells and the number of valence electrons; identify the s, p, d and f blocks.
Classifying elements	Distinguish metals, non-metals and metalloids and locate them on the table.
Periodic trends	Explain trends in atomic radius, ionic radius, ionisation energy, electron affinity and electronegativity using nuclear charge, shielding and radius.
Chemical trends	Describe the reactivity of Group 1 and Group 17, and the acid−base character of the period 3 oxides.
(HL) Transition elements	Define a transition element and explain variable oxidation states, coloured ions, catalysis, magnetism and complex-ion formation.
(HL) Origin of colour	Explain colour in terms of d-orbital splitting and absorption of visible light, and the factors that change it.

> **Exam note:** Sections flagged **(HL)** are assessed only at Higher Level. Everything else is common to SL and HL, but HL papers push the same ideas into less familiar, multi-step questions.

1. Structure of the periodic table

The modern periodic table arranges the elements in order of increasing **atomic number** Z (number of protons). Elements with similar chemical behaviour fall into the same vertical column because they share the same outer-shell electron arrangement. The table is the visual expression of electron configuration.

Periods, groups and blocks

- A **period** is a horizontal row. The period number equals the highest principal quantum number n (the number of occupied electron shells).
- A **group** is a vertical column. Elements in the same group have the same number of **valence electrons**, which is why they behave alike.
- A **block** is named after the subshell being filled by the last electron: the s-block (Groups 1−2 and He), the p-block (Groups 13−18), the d-block (transition elements) and the f-block (lanthanoids and actinoids).

Position and electron configuration

The link works both ways: given the configuration you can place the element, and given the position you can write the configuration. For a main-group element the group number gives the valence electron count directly.

Element	Configuration	Period (highest n)	Group	Valence e−
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	3	1	1

Element	Configuration	Period (highest n)	Group	Valence e−
Cl	[Ne] $3s^2$ $3p^5$	3	17	7
Ca	[Ar] $4s^2$	4	2	2
As	[Ar] $3d^{10}$ $4s^2$ $4p^3$	4	15	5

> **Group-number rule:** For s- and p-block elements the number of valence electrons is the last digit of the group number for Groups 13−18 (e.g. Group 15 → 5 valence electrons), and equals the group number itself for Groups 1 and 2.

Metals, non-metals and metalloids

A diagonal staircase running from boron to astatine separates metals (to the left, the majority of elements) from non-metals (to the upper right). Elements lying along the staircase – B, Si, Ge, As, Sb, Te – are **metalloids**, showing intermediate properties (e.g. semiconduction).

Property	Metals	Non-metals
Electrical conductivity	Good (delocalised electrons)	Poor (insulators; C as graphite is an exception)
State at room temperature	Mostly solid (Hg is liquid)	Solids, liquids and gases
When they react	Tend to lose electrons → form cations	Tend to gain electrons → form anions
Oxides	Basic (or amphoteric)	Acidic (or neutral)

2. Periodicity of atomic properties

Periodicity means the repeating pattern of properties as Z increases. Almost every trend in this section is explained by the balance between three factors, so learn them as a toolkit and apply them again and again.

Factor	What it does
Nuclear charge	More protons pull the electrons in more strongly, so radius falls and the outer electrons are held more tightly.
Shielding (screening)	Inner-shell electrons repel outer electrons and reduce the effective nuclear charge felt by the valence electrons.
Number of shells / distance	Adding a shell places valence electrons further out; distance weakens the attraction to the nucleus.

Atomic radius

Across a period (left → right): radius decreases. Each step adds one proton and one electron to the same shell. Shielding stays roughly constant, so the effective nuclear charge rises and pulls the shell inwards.

Down a group: radius increases. Each element has one more occupied shell, so the valence electrons are further out and better shielded, outweighing the larger nuclear charge.

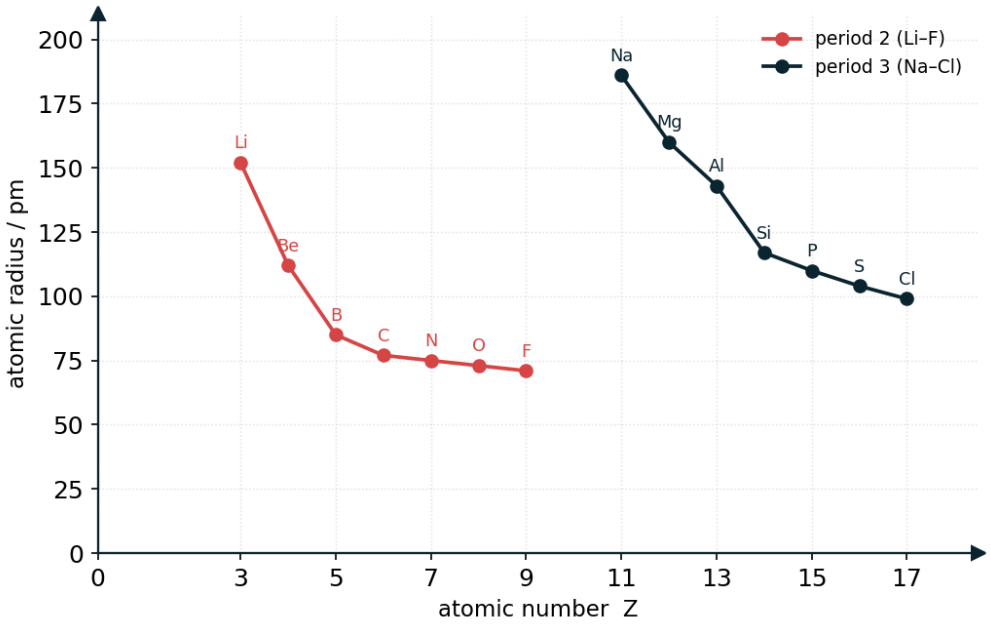

Figure 1. Atomic radius against atomic number. Radius falls across each period (Li–F, Na–Cl) as nuclear charge rises at nearly constant shielding, then jumps up at the start of the next period (Na) where a new shell is added.

Ionic radius

- **Cations are smaller than their parent atoms.** Losing electrons often empties the outer shell, and the remaining electrons feel a greater effective nuclear charge per electron, so they are drawn in. Example: Na is larger than Na^+.
- **Anions are larger than their parent atoms.** Added electrons increase electron–electron repulsion while the nuclear charge is unchanged, so the electron cloud expands. Example: Cl is smaller than Cl^-.
- **Isoelectronic series:** ions with the same number of electrons (e.g. N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+} all have 10 electrons). Within such a series, the radius **decreases** as nuclear charge increases, because the same 10 electrons are pulled by ever more protons.

> **Ranking tip:** To order an isoelectronic series by size, just order by nuclear charge: most protons → smallest ion. So $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$.

First ionisation energy

First ionisation energy (IE_1) is the energy to remove one mole of electrons from one mole of gaseous atoms: $X(g) \rightarrow X^+(g) + e^-$. It measures how tightly the outermost electron is held.

Across a period: IE_1 generally increases (rising nuclear charge, similar shielding, smaller radius). **Down a group: IE_1 decreases** (more shells and more shielding place the outer electron further from the nucleus).

Two dips interrupt the rise across a period, and examiners love them:

- **Group 13 dip** (e.g. Al lower than Mg). The electron removed comes from a 3p orbital, which is higher in energy and slightly better shielded than the filled 3s pair, so it is easier to remove.
- **Group 16 dip** (e.g. S lower than P). In Group 16 the p-subshell reaches its fourth electron, forcing a pair into one p-orbital. Electron–electron repulsion within that pair makes one electron easier to remove than in the stable half-filled p^3 arrangement of Group 15.

Successive ionisation energies: IE always rises for successive removals from the same atom. A large jump signals that the next electron comes from an inner shell − powerful evidence for the number of valence electrons and hence the group.

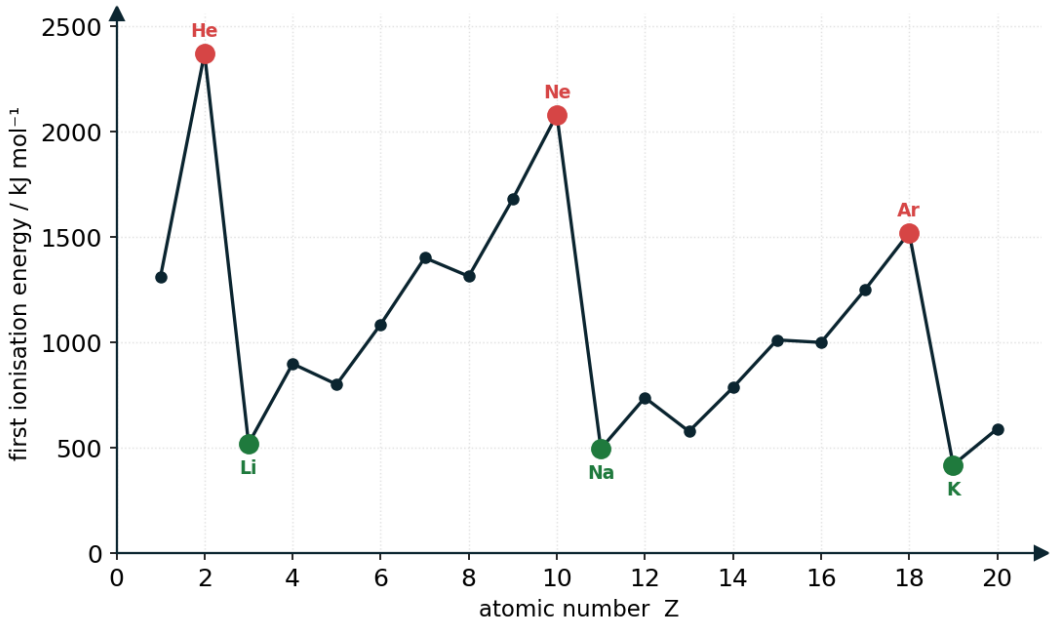

Figure 2. First ionisation energy against atomic number (Z = 1−20). IE_1 peaks at the noble gases He, Ne and Ar (red) and drops to a minimum at each alkali metal Li, Na, K (green). Small dips B below Be and O below N (repeated at Al and S) reveal sub-shell structure.

Electron affinity

First electron affinity is the enthalpy change when one mole of gaseous atoms each gain an electron: X(g) + e^- → X^-(g). For most elements it is **exothermic** (negative) because the added electron is attracted to the nucleus. It becomes more exothermic across a period (stronger nuclear pull) and less exothermic down a group (electron added further out). Non-metals, especially the halogens, release the most energy.

Electronegativity

Electronegativity is the relative ability of an atom to attract a shared (bonding) pair of electrons. On the Pauling scale it **increases across a period** (rising nuclear charge, smaller radius) and **decreases down a group** (larger radius, more shielding). Fluorine is the most electronegative element (4.0); the noble gases are normally excluded.

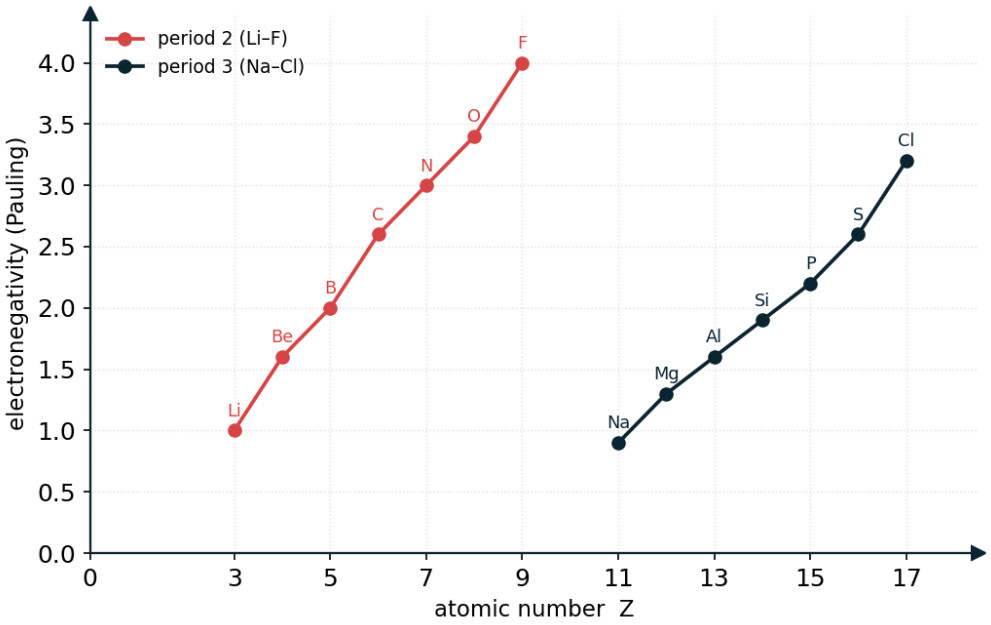

Figure 3. Electronegativity (Pauling) against atomic number. It rises across each period (Li−F, Na−Cl) as the atom contracts and nuclear charge grows; fluorine is the highest at 4.0. Down a group it falls (F 4.0 > Cl 3.2 in Group 17).

Property	Across a period (→)	Down a group (→)	Main reason
Atomic radius	Decreases	Increases	Nuclear charge vs. shells / shielding
Ionisation energy	Increases (2 dips)	Decreases	Attraction on the outer electron
Electron affinity	More exothermic	Less exothermic	Pull on an incoming electron
Electronegativity	Increases	Decreases	Pull on a bonding pair

3. Metallic character and melting points

Metallic character means the tendency to lose electrons and form positive ions. It tracks the opposite of ionisation energy and electronegativity.

- **Across a period metallic character decreases** (elements hold electrons more tightly), so non-metallic character increases.
- **Down a group metallic character increases** (outer electrons are lost more easily). The most metallic elements sit at the bottom left.

Melting points across period 3

The melting point across a period does not follow a single trend; it depends on the type of structure and bonding, which is why period 3 is a classic exam case.

Elements	Structure and bonding	Melting point trend
Na, Mg, Al	Giant metallic lattice	Rises: more valence electrons and higher charge give stronger metallic bonding (Al highest of the three)
Si	Giant covalent (macromolecular)	Highest of period 3: a strong 3-D network of covalent bonds must be broken

Elements	Structure and bonding	Melting point trend
P_4, S_8, Cl_2	Simple molecular	Low: weak London forces between molecules; S_8 > P_4 > Cl_2 by molecule size
Ar	Monatomic	Lowest: very weak forces between single atoms

> **Watch the shape of the graph:** Melting point climbs Na–Mg–Al, peaks sharply at Si, then collapses to the molecular non-metals. Always justify your answer by naming the structure, not just the position.

4. Group 1 and Group 17 reactivity

Group trends in reactivity are examined every year. The key point that trips students is that **reactivity increases down Group 1 but decreases down Group 17** — and the explanation is the same idea applied to opposite processes.

Group 1: the alkali metals

Alkali metals react by **losing** their single valence electron to form a 1+ ion. Going down the group the outer electron is further from the nucleus and better shielded, so ionisation energy falls and the electron is lost more readily: **reactivity increases down the group** (Li < Na < K < Rb < Cs).

Reaction with water gives a metal hydroxide (an alkali) and hydrogen gas:

$$\mathbf{2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)}$$

Lithium fizzes gently; sodium melts into a ball that darts across the surface; potassium ignites the hydrogen with a lilac flame; rubidium and caesium react explosively. The vigour of these reactions is the visible evidence of the trend.

Group 17: the halogens

Halogens react by **gaining** one electron to form a 1– ion (or by sharing one in covalent bonds). Going down the group the incoming electron is added to a shell further from the nucleus and more shielded, so the attraction is weaker: **reactivity decreases down the group** (F_2 > Cl_2 > Br_2 > I_2).

Physical state also changes smoothly down the group as London forces grow with molecular size:

Halogen	State at room temperature	Colour
F_2	Gas	Pale yellow
Cl_2	Gas	Pale green
Br_2	Liquid	Red-brown
I_2	Solid	Grey-black (violet vapour)

Displacement reactions

A more reactive halogen displaces a less reactive one from a solution of its salt, because it is the stronger oxidising agent. For example chlorine displaces bromine:

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

The solution turns orange as bromine forms. In ionic form, $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$. Iodine cannot displace bromine or chlorine, confirming the reactivity order.

> **The word 'reactivity' means different things here:** For a metal it means how easily it **loses** electrons, so reactivity rises down Group 1. For a non-metal it means how easily it **gains** electrons, so reactivity falls down Group 17. Never quote one trend as if it were universal.

5. Periodic trends in oxides

The acid–base character of an oxide follows the metal / non-metal divide, so it changes systematically across a period. This is the standard link between position and chemical behaviour.

- **Metal oxides are basic.** They react with acids to form a salt and water, and ionic ones dissolve to give alkaline solutions.
- **Non-metal oxides are acidic.** They react with bases, and many dissolve in water to form acids.
- **Amphoteric oxides** (e.g. Al_2O_3) react with *both* acids and bases, marking the changeover along the period.

Across period 3 the oxides move from basic through amphoteric to acidic:

Oxide	Na_2O, MgO	Al_2O_3	SiO_2	P_4O_{10}, SO_3, Cl_2O_7
Character	Basic	Amphoteric	Weakly acidic	Acidic
Element type	Metal	Metal (borderline)	Metalloid	Non-metal

Representative reactions (qualitative)

- Basic oxide + acid: $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$.
- Acidic oxide + base: $SO_3(g) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$.
- Acidic oxide + water: $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$.
- Amphoteric oxide, both ways: $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$, and $Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$.

> **Environmental link:** Acidic non-metal oxides such as SO_2, SO_3 and the nitrogen oxides dissolve in atmospheric water to form the acids responsible for acid deposition – a favourite application question.

6. (HL) Transition elements

The d-block spans Groups 3–12. A **transition element** is defined as an element that forms at least one stable ion with a **partially filled d subshell**. This precise definition matters in the exam.

> **(HL) Two textbook exceptions:** Scandium and zinc are d-block metals but are **not** classed as transition elements. Sc forms only Sc^{3+} ([Ar], empty d), and Zn forms only Zn^{2+} ([Ar]$3d^{10}$, full d) – neither ion has a partially filled d subshell.

Remember also the filling anomaly: the 4s electrons are removed **before** the 3d electrons when a transition-metal ion forms. So Fe is [Ar]$3d^6 4s^2$ but Fe^{2+} is [Ar]$3d^6$ and Fe^{3+} is [Ar]$3d^5$.

Characteristic properties

Property	Explanation
Variable oxidation states	The 3d and 4s subshells are close in energy, so a variable number of electrons can be lost. Fe (+2, +3), Cu (+1, +2), Mn (+2 to +7) are common examples.
Coloured ions and complexes	Partially filled d orbitals split in a ligand field; ions absorb visible light on d−d transitions (see section 7).
Catalytic activity	Variable oxidation states and the ability to bind reactant molecules let them provide low-energy pathways. Fe in the Haber process, V_2O_5 in the Contact process, Ni in hydrogenation.
Magnetic behaviour	Unpaired d electrons make many complexes paramagnetic (attracted into a magnetic field); species with only paired electrons are diamagnetic.
Complex-ion formation	The small, highly charged metal ions act as Lewis acids, accepting lone pairs from ligands (Lewis bases) to form dative covalent bonds.

Complexes and ligands

A **complex ion** is a central metal ion bonded to a number of **ligands** – molecules or ions with a lone pair (e.g. H_2O, NH_3, Cl^-, CN^-, OH^-). The number of dative bonds to the central ion is the **coordination number**, commonly 6 (octahedral) or 4 (tetrahedral or square planar). Example: $[Cu(H_2O)_6]^{2+}$ is octahedral with coordination number 6.

7. (HL) The origin of colour

In a free gaseous ion the five d orbitals have equal energy (degenerate). When ligands approach, their lone pairs repel the d electrons unequally and the d orbitals **split** into two energy levels separated by a gap ΔE.

A d electron can absorb a photon of visible light whose energy matches ΔE and jump to the higher level (a **d−d transition**). The wavelengths absorbed are removed from white light; the colour we see is the **complementary colour** of the light absorbed. For example a complex that absorbs red light appears green.

Because Sc^{3+} (empty d) and Zn^{2+} (full d) have no possible d−d transition, their compounds are colourless – a neat confirmation of the definition in section 6.

Factors that change the colour

Anything that changes the size of the split ΔE changes the energy, and hence the wavelength, absorbed – so the observed colour changes:

Factor	Effect on colour
Identity of the ligand	Different ligands split the d orbitals by different amounts (the spectrochemical series). Swapping ligands shifts the colour, e.g. pale blue $[Cu(H_2O)_6]^{2+}$ turns deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$.
Oxidation state of the metal	A higher charge pulls ligands closer and increases ΔE. Fe^{2+}(aq) is pale green while Fe^{3+}(aq) is yellow-brown.

Factor	Effect on colour
Geometry / coordination number	Octahedral and tetrahedral fields split the orbitals differently, giving different ΔE and different colours.
Nature of the metal	The number of d electrons and nuclear charge differ from metal to metal, so each ion has its own characteristic colours.

Naming simple complexes (qualitative)

You are expected to interpret, not derive, complex formulae. The overall charge is the sum of the metal-ion charge and the ligand charges. Worked logic:

- $[Cu(H_2O)_6]^{2+}$: six neutral water ligands, so the copper is +2. Coordination number 6, octahedral.
- $[Fe(CN)_6]^{3-}$: six CN^- ligands total 6−; overall 3− means iron is +3. Coordination number 6.
- $[CuCl_4]^{2-}$: four Cl^- total 4−; overall 2− means copper is +2. Coordination number 4, tetrahedral.

8. Worked examples

Worked example 1 — predict and explain a trend

Explain why the atomic radius of sulfur is smaller than that of both sodium (same period) and oxygen (same group).

Sodium → sulfur: across period 3 the nuclear charge rises from 11 to 16 while electrons enter the same third shell, so shielding is nearly constant. The greater effective nuclear charge pulls the shell in, so S is smaller than Na.

Oxygen → sulfur: down Group 16 sulfur has an extra occupied shell (period 3 vs 2) and more shielding, so its valence electrons are further out. Hence S is larger than O.

Answer: the two comparisons use different factors – nuclear charge across, shell number down – and must be argued separately.

Worked example 2 — identify an element from properties

An element X is a soft, silvery solid that reacts vigorously with cold water to give an alkaline solution and a flammable gas, and burns with a lilac flame. Identify the group and a likely identity.

A soft metal that reacts vigorously with water to form an alkali plus hydrogen is a Group 1 alkali metal. The lilac flame colour is characteristic of potassium.

X is potassium (K): $2K + 2H_2O \rightarrow 2KOH + H_2$. The vigour (more than Na, less than Rb) fits its position in the group.

Worked example 3 — write a halogen reaction

Predict whether a reaction occurs when chlorine water is added to potassium iodide solution, and write the equation.

Chlorine is above iodine in Group 17, so it is the more reactive (stronger oxidising) halogen and will displace iodine.

$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$; ionically $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$. The solution turns brown as iodine forms.

Worked example 4 — deduce oxide acidity

Element Q is in period 3 and forms an oxide that reacts with both dilute hydrochloric acid and sodium hydroxide solution. Identify Q and classify the oxide.

An oxide that reacts with both an acid and a base is **amphoteric**. In period 3 the amphoteric oxide is aluminium oxide, Al_2O_3.

Q is aluminium: $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ (acting as a base) and $Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$ (acting as an acid).

Worked example 5 — explain an ionisation-energy dip (HL-style)

The first ionisation energies across period 3 rise overall, yet IE_1(S) is lower than IE_1(P). Explain.

Phosphorus has the configuration $[Ne]3s^23p^3$ – a stable half-filled p subshell with one electron in each p orbital.

Sulfur, $[Ne]3s^23p^4$, must place a second electron into one 3p orbital. The extra electron–electron repulsion in that paired orbital makes it easier to remove, so $IE_1(S) < IE_1(P)$ despite the higher nuclear charge.

Worked example 6 — order an isoelectronic series

Place the ions O^{2-}, F^-, Na^+ and Mg^{2+} in order of decreasing radius.

All four have 10 electrons (isoelectronic). Radius is set by nuclear charge: O has 8 protons, F 9, Na 11, Mg 12. More protons pull the same 10 electrons in more tightly.

Order (largest → smallest): $O^{2-} > F^- > Na^+ > Mg^{2+}$.

9. Common pitfalls

- Forgetting the two IE dips (Groups 13 and 16) and drawing a smooth rise across a period.
- Comparing ionic radii without checking whether the ions are isoelectronic – compare like with like.
- Saying 'reactivity increases down every group'. It increases down Group 1 but decreases down Group 17.

- Confusing electron affinity (adding an electron) with ionisation energy (removing one).
- Removing 3d electrons before 4s when writing transition-metal ions (HL) – the 4s electrons leave first.
- Calling Sc and Zn transition elements (HL). Their common ions have empty or full d subshells.
- Quoting a melting-point 'trend' across period 3 without naming the structure (metallic, giant covalent, molecular, monatomic).

10. Quick-reference trends

Quantity	Across a period (→)	Down a group (→)
Atomic radius	Decreases	Increases
First ionisation energy	Increases (dips at Gp 13, 16)	Decreases
Electron affinity	More exothermic	Less exothermic
Electronegativity	Increases	Decreases
Metallic character	Decreases	Increases
Non-metallic character	Increases	Decreases
Oxide character	Basic → amphoteric → acidic	More basic further down
Group 1 reactivity	—	Increases (electron lost more easily)
Group 17 reactivity	—	Decreases (electron gained less easily)

11. Test yourself

Attempt all eleven questions without notes; full answers follow.

1. State the block, period and group of the element with configuration $[Ar]3d^{10}4s^{2}4p^{4}$.

2. Explain, in terms of structure, why silicon has the highest melting point in period 3.

3. Arrange P^{3-}, S^{2-} and Cl^{-} in order of increasing ionic radius and justify your order.

4. Explain why the first ionisation energy of aluminium is lower than that of magnesium.

5. Write a balanced equation for the reaction of potassium with water and state one observation.

6. Predict whether bromine will displace chlorine from sodium chloride solution, and explain.

7. Classify the oxides Na_2O, Al_2O_3 and SO_3 as basic, amphoteric or acidic, and give one equation for each with a suitable reagent.

8. (HL) Define a transition element and explain why zinc does not qualify.

9. (HL) Explain why aqueous copper(II) ions are blue whereas aqueous zinc ions are colourless.

10. (HL) Deduce the oxidation state of the metal in $[Fe(CN)_6]^{4-}$ and state its coordination number.

11. Explain why electronegativity increases from left to right across a period.

Answers

1. p-block (last electron in 4p); period 4 (highest n = 4); Group 16 (this is selenium, Se). Ignore the filled 3d in counting the block — it is the last subshell filled that matters.

2. Silicon is a giant covalent (macromolecular) solid. Melting requires breaking many strong SiSi covalent bonds throughout the lattice, which needs far more energy than melting the metallic or simple-molecular elements of period 3.

3. Increasing radius: $Cl^- < S^{2-} < P^{3-}$. All three are isoelectronic (18 electrons). Cl has the most protons (17) and P the fewest (15), so Cl^- is pulled in most tightly and P^{3-} least.

4. Mg is [Ne]$3s^2$; Al is [Ne]$3s^2 3p^1$. Aluminium's outer electron is in a 3p orbital, higher in energy and slightly shielded by the 3s pair, so it is removed more easily than an electron from magnesium's full 3s subshell — giving the Group 13 dip.

5. $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$. Observation: the metal fizzes and melts, moving rapidly on the surface, and the hydrogen ignites with a lilac flame.

6. No reaction. Bromine is below chlorine in Group 17, so it is the *less* reactive (weaker oxidising) halogen and cannot displace the more reactive chlorine from its salt.

7. Na_2O basic: $Na_2O + 2HCl \rightarrow 2NaCl + H_2O$. Al_2O_3 amphoteric: $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ (with acid) or $+ 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$ (with base). SO_3 acidic: $SO_3 + 2NaOH \rightarrow Na_2SO_4 + H_2O$.

8. (HL) A transition element forms at least one stable ion with a partially filled d subshell. Zinc forms only Zn^{2+}, which is [Ar]$3d^{10}$ — a completely full d subshell — so it is excluded.

9. (HL) Cu^{2+} has a partially filled 3d subshell ($3d^9$); its d orbitals split in the aqua-ligand field and it absorbs part of the visible spectrum by a d–d transition, appearing blue (the complementary colour). Zn^{2+} is $3d^{10}$ (full), so no d–d transition is possible and it is colourless.

10. (HL) Six CN^- ligands contribute 6–; the overall charge is 4–, so the iron must be +2. Coordination number 6 (octahedral).

11. Across a period the nuclear charge increases while electrons enter the same shell and shielding is almost unchanged. The atom gets smaller and the effective nuclear charge felt by a bonding pair increases, so the atom attracts shared electrons more strongly — higher electronegativity.