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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Structure 3.1
The Periodic Table

Booklet 1 · Understandings 3.1.1 – 3.1.5

Periods, groups and blocks, electron configuration and position, periodic trends, group 1 and group 17 chemistry, and the oxides from basic to acidic

Complete revision notes & memory points · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Core material. Understandings 3.1.1 to 3.1.5 are common to both levels. Every student must know all of it.

HL Short preview panels point ahead to the HL understandings 3.1.7 to 3.1.10 (covered in Booklets 2 and 3). SL students can skip these panels.

Inside this booklet

1. What the syllabus actually asks of you
2. 3.1.1 — The structure of the periodic table
3. 3.1.2 — Electron configuration and position
4. 3.1.3 — Periodicity: the five trends
5. 3.1.4 — Trends down a group: groups 1, 17 and 18
6. 3.1.5 — From metallic to non-metallic: the oxides
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1 What the syllabus actually asks of you

The guiding question for Structure 3.1 is: *how does the periodic table help us to predict patterns and trends in the properties of elements?* Instead of learning the chemistry of more than 100 elements one by one, you learn the chemistry of a few groups and the trends that link them: the behaviour of potassium can be predicted from that of lithium and sodium. The periodic table is a map: an element's position tells you its electron configuration, and the configuration tells you how the element will behave. If you understand **why** the trends happen, you can predict the properties of an element you have never studied.

Understanding	What it says	What you must be able to do	Level
3.1.1 Periods, groups and blocks	The periodic table consists of periods, groups and blocks.	Identify the positions of metals, metalloids and non-metals. Recognize the s, p, d and f blocks. (A periodic table is in the data booklet.)	SL+HL
3.1.2 Configuration and position	The period number shows the outer energy level occupied by electrons. Elements in a group have a common number of valence electrons.	Deduce the electron configuration of an atom up to Z = 36 from its position in the table, and the position from the configuration. Groups are numbered 1 to 18. Know the names alkali metals, halogens, transition elements and noble gases.	SL+HL
3.1.3 Periodicity	Periodicity refers to trends in the properties of elements across a period and down a group.	Explain the periodicity of atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.	SL+HL
3.1.4 Group trends	Metallic character increases down group 1; non-metallic character decreases down group 17.	Describe and explain the reactions of group 1 metals with water, and of group 17 elements with halide ions.	SL+HL
3.1.5 Oxides	Metallic and non-metallic properties show a continuum, from basic metal oxides through amphoteric oxides to acidic non-metal oxides.	Deduce equations for the reactions with water of the oxides of group 1 and group 2 metals, carbon and sulfur. Include acid rain caused by gaseous non-metal oxides, and ocean acidification caused by rising CO_2 levels.	SL+HL

Linking questions and connections for 3.1.1 – 3.1.5

- **Nature of Science, Structure 1.2:** how has the organization of elements in the periodic table helped the discovery of new elements? → section 2.4
- **Structure 1.3:** electron configurations, sublevels and the order of filling are the foundation of this whole topic. → section 3
- **Inquiry 2, Tool 2:** why are simulations often used to explore the reactivity trends of groups 1 and 17? → section 5.5
- **Structure 2.1, 2.2:** how do differences in bonding explain the different properties of metal and non-metal oxides? → section 6.5
- **Structure 2.4:** the period 3 oxides on the bonding triangle. **Reactivity 3.1 and 3.2:** acids, bases and redox (the halogen displacement reactions are redox reactions).

What the data booklet gives you — and what it does not

The data booklet contains a full **periodic table** (atomic numbers and relative atomic masses) and tables of **electronegativity, first ionization energy and electron affinity** (Section 9) and of **atomic and ionic radii** (Section 10). You can read values from it, but you must be able to **explain** each trend yourself. The data booklet does **not** give group names, electron configurations, observations for reactions, colours of halogens, or the equations for oxides reacting with water: you must learn these.

2 Structure 3.1.1 — The structure of the periodic table SL + HL

2.1 Periods, groups and blocks

The modern periodic table lists the 118 known elements in order of increasing **atomic number**, Z (the number of protons). The table is arranged so that elements with similar properties fall into the same column.

Definition

Periodicity is the regular repetition of the properties of the elements, arising from the repeating pattern in their electron configurations.

Because there are no gaps in the sequence of atomic numbers from 1 to 118, chemists can be confident that no more elements are waiting to be found in nature. The only way to extend the table is to **make new elements artificially** in nuclear reactions.

Definitions to learn

- A **period** is a horizontal row of the periodic table. There are 7 periods.
- A **group** is a vertical column. There are 18 groups, numbered 1 to 18. Elements in the same group have similar chemical properties.
- A **block** is a region of the table named after the sublevel (s, p, d or f) that is being filled by the highest-energy electron of the atoms in that region.

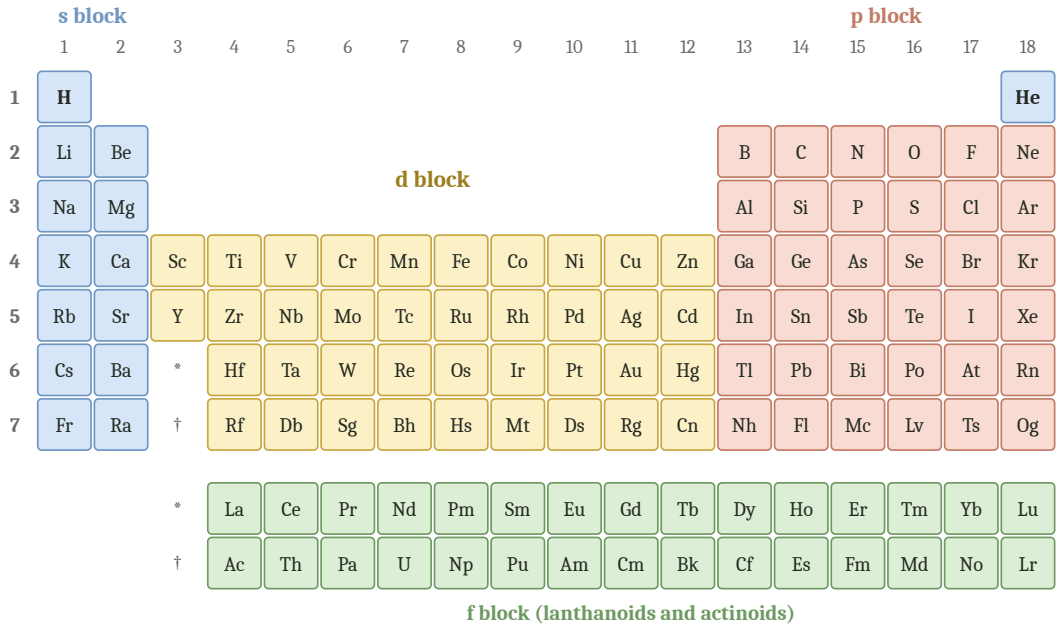


Figure 1 — The four blocks. Groups 1 and 2 (plus helium) form the s block, groups 13 to 18 the p block, groups 3 to 12 the d block, and the lanthanoids and actinoids the f block. The f block is normally printed below the main table to keep the table a manageable width.

Block	Groups	Sublevel being filled	Elements per period
s	1 and 2 (and He)	s: holds 2 electrons	2
p	13 to 18	p: holds 6 electrons	6
d	3 to 12	d: holds 10 electrons	10
f	lanthanoids, actinoids	f: holds 14 electrons	14

The width of each block is therefore not an accident: it is the maximum number of electrons in that kind of sublevel. The shape of the periodic table is a direct consequence of electron configurations (Structure 1.3).

2.2 Metals, metalloids and non-metals

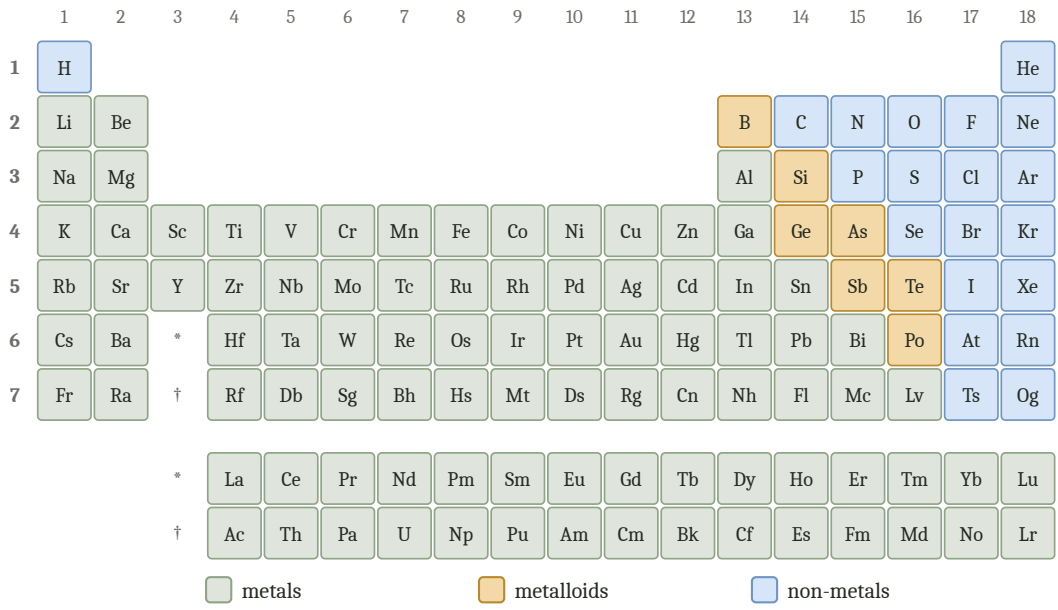


Figure 2 — About three quarters of the elements are metals. A diagonal "staircase" of metalloids separates the metals (left and bottom) from the non-metals (top right).

Metals are found in the s block (except hydrogen and helium), the whole of the d block and the f block, and on the left of the p block (for example aluminium, tin and lead). The non-metals are in the top right of the p block, plus hydrogen. Between them runs a diagonal band of **metalloids**: boron, silicon, germanium, arsenic, antimony and tellurium (polonium is sometimes included). Metalloids have properties that are intermediate between those of metals and non-metals: physically they tend to resemble metals, while chemically they behave more like non-metals. (Textbooks differ slightly over the list: some leave out boron, some include polonium.) Silicon, for example, looks shiny like a metal, but it is brittle and is a **semiconductor**: it conducts electricity much less well than a metal, and its conductivity increases as temperature rises.

Property	Metals	Non-metals
Appearance	shiny (lustrous)	usually dull; many are gases
Physical state	solids at room temperature (except mercury)	gases, molecular solids, or bromine (a liquid)
Conductivity	good conductors of heat and electricity	poor conductors (graphite is an exception)
Mechanical	malleable and ductile	brittle when solid
Ions formed	lose electrons to form **cations**	gain electrons to form **anions**
Oxides	**basic** (section 6)	**acidic** (section 6)
Electronegativity	low	high

2.3 The awkward cases: hydrogen and helium

Hydrogen ($1s^1$) has one valence electron, so it is often placed above group 1. But it is a non-metal that exists as H_2 molecules, and it can either lose an electron (H^+) or gain one (H^-, in metal hydrides such as NaH). It has no true group.

Helium ($1s^2$) is an s-block element with two electrons, which might suggest group 2. It is placed in **group 18** because its outer level (n = 1) is **full**, which makes it an unreactive noble gas like neon and argon. Its **properties** decide its position.

2.4 Nature of Science — a table that predicted the future

Several chemists in the 19th century arranged elements by relative atomic mass and noticed that properties repeated at regular intervals. In 1869 **Dmitri Mendeleev** produced a table that stood out for two bold decisions:

- he **left gaps** where no known element fitted, and **predicted the properties** of the missing elements from those of their neighbours. When gallium (1875), scandium (1879) and germanium (1886) were discovered, their properties closely matched his predictions, which persuaded other chemists that the table reflected something real. A fourth gap was not filled until technetium was made artificially in 1937. When gallium was first isolated, its density was measured as about 4.7 g cm^{-3}; Mendeleev had predicted about 5.9 g cm^{-3} and suggested the sample was impure. A purer sample gave 5.9 g cm^{-3}: the model corrected the measurement;
- he **swapped the order** of some elements, such as tellurium and iodine, so that they sat in groups with similar properties, even though this went against the order of atomic mass.

The reason for the swaps became clear in 1913, when Henry Moseley showed that each element has a characteristic nuclear charge. Ordering by **atomic number** instead of atomic mass removed the anomalies. This is a good example of a scientific model that was **predictive** before the underlying explanation (protons and electron configurations) was known. The table continues to grow. Early elements were often isolated by one chemist; the newest are made, an atom or two at a time, by large international teams, so they are as much invented as discovered. Each claim must be verified by IUPAC, a form of **peer review**, before the element is accepted and named. Period 7 was completed in 2016 with the naming of nihonium, moscovium, tennessine and oganesson.

Nature of Science and TOK: a law is not an explanation

Mendeleev's **periodic law** states that properties repeat at intervals when the elements are arranged in order. Like other scientific laws, it **describes** a pattern but does not **explain** it. The explanation had to wait for the discovery of the electron (Thomson, 1897) and the atomic models of Rutherford and Bohr; today we trace the pattern to electron configurations. Unlike a law such as the conservation of mass, the periodic law is not exact: the intervals vary (2, 8, 8, 18, 18, 32). TOK: the table was found by **induction** from observed properties and later explained by **deduction** from atomic structure. How far do the classification systems we choose shape the conclusions we reach?

Examiner traps in 3.1.1

- The periodic table is ordered by **atomic number**, not by relative atomic mass. The Ar/K, Co/Ni and Te/I pairs are "out of mass order".
- Rows are **periods** and columns are **groups**. Do not mix them up in a written answer.
- Groups are numbered **1 to 18**. Old numbering (group VII, group 0) will not be credited.
- Helium is in the **s block** but in **group 18**. Both statements are true.
- Metalloids are not "half metals": state their **intermediate properties**, for example semiconductivity.

3 Structure 3.1.2 — Electron configuration and position SL + HL

3.1 What the period and group numbers tell you

The two rules to learn

- The **period number** equals the number of the **outer (highest) occupied principal energy level**, n. Sodium is in period 3 because its outer electron is in n = 3.
- Elements in the same **group** have the same number of **valence electrons** (outer electrons) and the same outer configuration. For groups 1 and 2, the number of valence electrons equals the group number. For groups 13 to 18, it equals the group number minus 10 (helium is the exception, with 2).

Because chemical reactions involve the valence electrons, elements in the same group react in similar ways. That is why the table "works".

Group	Name	Outer configuration	Typical behaviour
1	**alkali metals**	ns^1	reactive metals; form 1+ ions
2	alkaline earth metals	ns^2	reactive metals; form 2+ ions
3–12	**transition elements**	$(n–1)d^{1–10}\ ns^{1–2}$	metals; variable oxidation states, coloured compounds (HL detail in Booklet 3)
17	**halogens**	$ns^2\ np^5$	reactive non-metals; form 1– ions
18	**noble gases**	$ns^2\ np^6$ (He: $1s^2$)	very unreactive; monatomic gases

The four names in bold are required by the guide. "Alkaline earth metals" is not required, but you will often meet it. Strictly, zinc (group 12) is not a transition element, because neither Zn nor Zn^{2+} has an incomplete d sublevel (Structure 2.3, HL).

3.2 Reading a configuration from the table

Read across the table, period by period, adding electrons until you reach the element. Each block tells you which sublevel the electrons are entering. Remember that 4s fills before 3d, so in period 4 the d block corresponds to the **3d** sublevel.

1. Find the element. Its **period** gives the outer level, n. Its **block** gives the sublevel the last electron enters.
2. Write the configuration of the previous noble gas in square brackets as a **core**: [He], [Ne] or [Ar].
3. Count across the current period: s block = *n*s, d block = (*n*–1)d, p block = *n*p. Add electrons until you reach the element.
4. Check: total electrons = Z. Write the full form if asked ("full electron configuration" means no noble gas core).

Worked example — selenium, Z = 34

Selenium is in period 4, group 16, p block. The core is [Ar] (18 electrons). Crossing period 4: $4s^2$ (K, Ca), then $3d^{10}$ (Sc to Zn), then 4p: Ga, Ge, As, Se is the fourth p-block element, so $4p^4$.

Condensed: [Ar] $3d^{10}\ 4s^2\ 4p^4$ Full: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$

Check: 18 + 10 + 2 + 4 = 34 ✓. Six valence electrons ($4s^2\ 4p^4$), as expected for group 16.

Worked example — the reverse: which element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$?

Highest n = 4, so **period 4**. The last sublevel filling is 3d, so it is in the **d block**. Six d electrons means the sixth element of the d block: Sc, Ti, V, Cr, Mn, **Fe**. Total electrons = 26 = Z for iron. It is in **group 8** (2 + 6 = 8 electrons beyond the [Ar] core).

Two exceptions you must know (from Structure 1.3)

Chromium is [Ar] $3d^5\ 4s^1$, not $3d^4\ 4s^2$; copper is [Ar] $3d^{10}\ 4s^1$, not $3d^9\ 4s^2$. A half-full or full 3d sublevel gives a lower overall energy. Reading straight off the table gives the "expected" configuration, so these two need to be remembered separately.

Model answer: "Explain why chlorine and bromine have similar chemical properties." (2 marks)

Both are in group 17 and have the same outer electron configuration, $ns^2\ np^5$ (seven valence electrons) **[1]**. Chemical reactions involve the valence electrons, so both react in the same way, for example by gaining one electron to form a 1– ion **[1]**.

Examiner traps in 3.1.2

- The period number is the number of the **outer occupied energy level**, which is also the number of occupied levels. It is not the number of valence electrons.
- In period 4, the d block is **3d**, not 4d. "4d" for iron is a common slip.
- When asked for the **full** configuration, do not use [Ar]. When the configuration is written in order of energy level, 3d is often written before 4s; both orders are accepted.
- For groups 13 to 18, subtract 10 from the group number to get valence electrons: aluminium (group 13) has 3, not 13.
- A **full** inner d sublevel does not count as valence electrons in a p-block element. Bromine, [Ar] $3d^{10}\ 4s^2\ 4p^5$, has **7** valence electrons ($4s^2\ 4p^5$), not 17.

4 Structure 3.1.3 — Periodicity: the five trends SL + HL

Periodicity is the repeating pattern of properties across periods and down groups. Every trend in this section is explained by the same three factors. Learn them first, and the five trends follow.

4.1 Nuclear charge, shielding and effective nuclear charge

The three factors

- **Nuclear charge** — the number of protons. More protons means a stronger attraction for every electron.
- **Shielding** — electrons in inner (lower) energy levels repel the outer electrons and partly cancel the nuclear attraction they feel. Electrons in the **same** level shield each other only weakly.
- **Distance** — the further an electron is from the nucleus (the higher its energy level), the weaker the attraction.

Combining the first two gives the **effective nuclear charge**, Z_{eff}: the net positive charge actually felt by an outer electron.

A simple estimate is **Z_{eff} ≈ number of protons – number of inner electrons**.

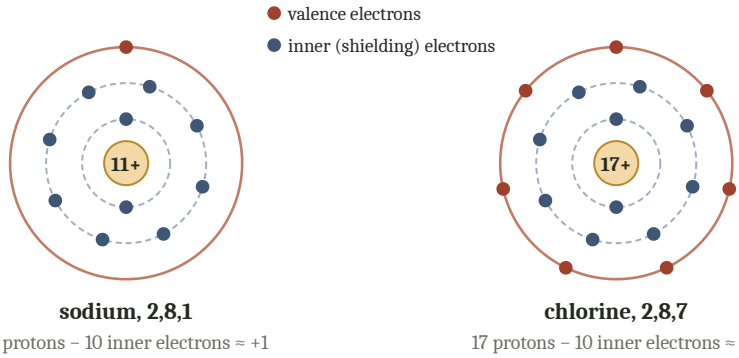


Figure 3 — Sodium and chlorine both have 10 inner electrons (2,8) shielding the outer level. The outer electrons of chlorine feel a much larger effective nuclear charge (about +7) than the outer electron of sodium (about +1).

The two patterns that explain everything

- **Across a period** (left to right): the nuclear charge increases by one each step, but the new electrons go into the **same** energy level, so shielding stays about the same. Z_{eff} **increases** and the outer electrons are pulled **closer** and held **more strongly**.
- **Down a group**: each step adds a whole new energy level. The outer electrons are **further** from the nucleus and more shielded. The extra protons are cancelled by the extra inner electrons, so Z_{eff} stays **about the same** (about +1 for every group 1 atom, about +7 for every halogen). The increase in distance wins, so the outer electrons are held **less strongly**.

4.2 Atomic radius

Definition

The **atomic radius** is half the distance between the nuclei of two bonded atoms of the same element. (An atom has no sharp edge, so its radius has to be measured this way.)

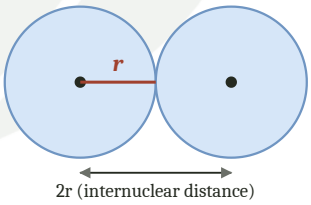


Figure 4 — Atomic radius.

- **Decreases across a period.** Z_{eff} increases while the outer electrons stay in the same energy level, so they are pulled closer to the nucleus. Na 160 pm, Cl 100 pm: a chlorine atom is only about 60% of the radius of a sodium atom.
- **Increases down a group.** Each element has one more occupied energy level, so the outer electrons are further from the nucleus; Z_{eff} is about constant. Li 130 pm, Cs 238 pm.

The data booklet (Section 10) gives no atomic radii for the noble gases. Their atoms do not form covalent bonds with each other, so there is no bond length to halve.

4.3 Ionic radius

Atoms change size when they become ions. Two rules cover all cases:

- **Cations are smaller than their parent atoms.** A period 3 metal atom loses its whole outer level (Na: 2,8,1 → Na^+: 2,8). The remaining electrons are in a lower level and are attracted by the same number of protons with less electron–electron repulsion.

- **Anions are larger than their parent atoms.** The added electrons go into the existing outer level. The nuclear charge is unchanged, but there is more repulsion between the electrons, so the electron cloud spreads out.

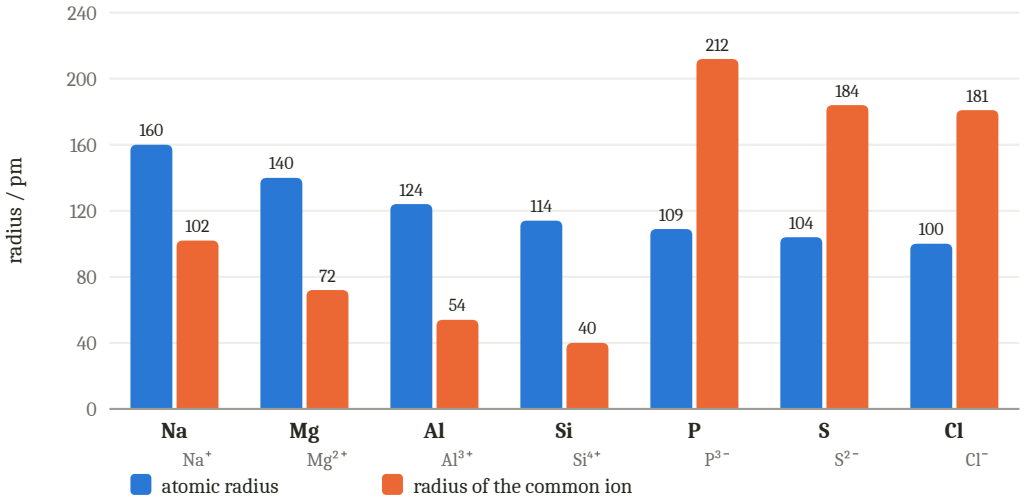


Figure 5 — Atomic radius (blue) and the radius of the common ion (orange) across period 3, in pm. The cations Na^+ to Si^{4+} are much smaller than their atoms; the anions P^{3-} to Cl^- are much larger.

Look at the orange bars in Figure 5. There are three distinct trends:

1. **Na^+ to Si^{4+}: radius decreases.** All four ions have the same configuration (2,8), but the number of protons rises from 11 to 14, so the same electrons are pulled in more strongly.
2. **Si^{4+} to P^{3-}: a large jump up.** The anions have one more occupied energy level (2,8,8) than the cations (2,8). Silicon shows the effect most dramatically: the hypothetical Si^{4+} ion has a radius of about 40 pm, while Si^{4-} would be about 271 pm, nearly seven times larger, with the same nucleus.
3. **P^{3-} to Cl^-: radius decreases.** Again all three ions have the same configuration (2,8,8), and the nuclear charge increases from 15 to 17.

Down any group, ionic radius **increases**, for the same reason as atomic radius: one more energy level each step ($Li^+ < Na^+ < K^+$; $F^- < Cl^- < Br^-$).

Isoelectronic species

Species with the **same number of electrons** are **isoelectronic**. In an isoelectronic series, the radius decreases as the number of protons increases, because the same electrons are attracted by a greater nuclear charge.

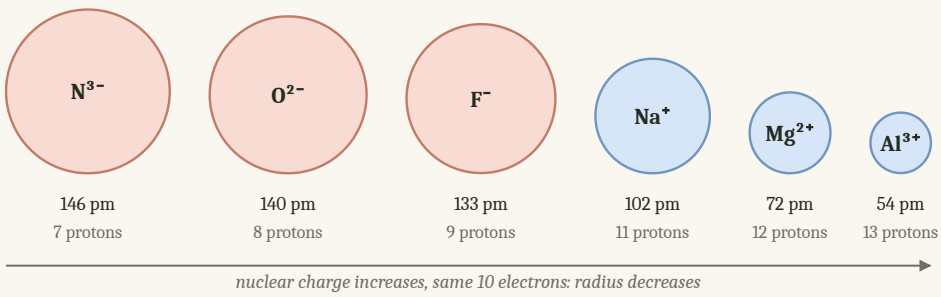


Figure 6 — Six ions, each with 10 electrons (configuration 2,8).

4.4 First ionization energy

Definition to learn

The **first ionization energy** is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms in their ground state, forming one mole of gaseous 1+ ions.

$$X(g) \rightarrow X^{+}(g) + e^{-} \quad \text{for example, } Na(g) \rightarrow Na^{+}(g) + e^{-} \quad \Delta H = +496 \text{ kJ mol}^{-1}$$

Ionization energy is always **positive** (endothermic): energy must be supplied to pull an electron away from the attraction of the nucleus.

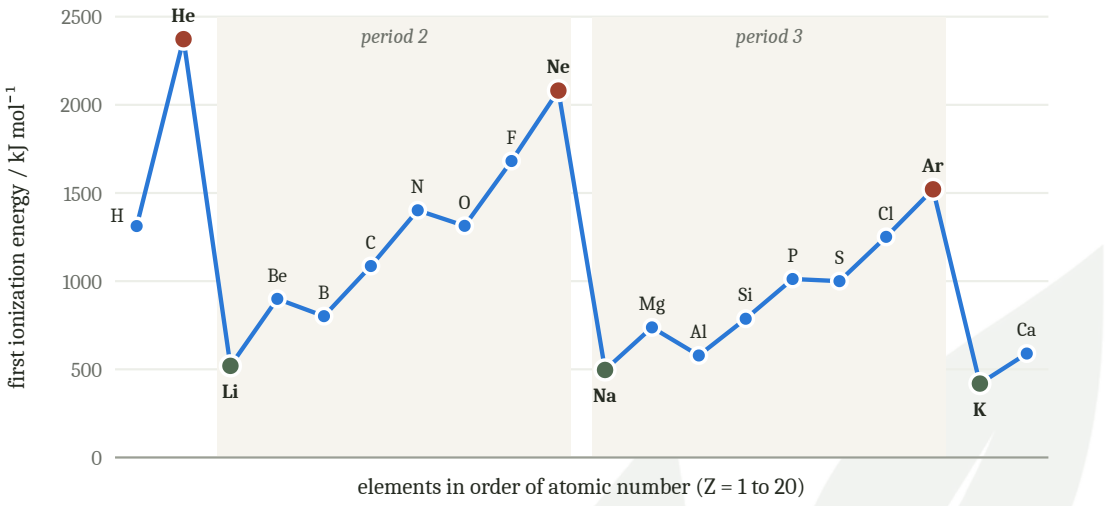


Figure 7 — First ionization energies of hydrogen to calcium. Noble gases (red) are at the peaks; alkali metals (green) are at the troughs.

- **General increase across a period.** Z_{eff} increases and the atomic radius decreases, so the outer electron is held more strongly and more energy is needed to remove it. The noble gas at the end of each period has the highest value.
- **Decrease down a group.** The outer electron is in a higher energy level, further from the nucleus and more shielded, so it is held less strongly (He 2372 → Ne 2081 → Ar 1521 kJ mol^{-1}).
- The ionization energy trend is the **mirror image** of the atomic radius trend: both measure how strongly the nucleus holds the outer electrons.
- **The sharp drop** from a noble gas to the next alkali metal (Ne → Na) happens because the outer electron of the alkali metal is the first electron in a **new** energy level.

HL **Preview of 3.1.7: the two small dips**

The increase across a period is not smooth. There are small dips from group 2 to group 13 (Be → B, Mg → Al) and from group 15 to group 16 (N → O, P → S). These are evidence for sublevels, and they are explained in HL 3.1.7 in Booklet 2. At SL you only need the general trend.

4.5 Electron affinity

Definition to learn

The **first electron affinity** is the energy change when one mole of electrons is added to one mole of gaseous atoms, forming one mole of gaseous 1− ions.

$$X(g) + e^- \rightarrow X^-(g) \quad \text{for example, } Cl(g) + e^- \rightarrow Cl^-(g) \quad \Delta H = -349 \text{ kJ mol}^{-1}$$

For most atoms the first electron affinity is **negative** (exothermic): the incoming electron is attracted by the effective nuclear charge, so energy is released.

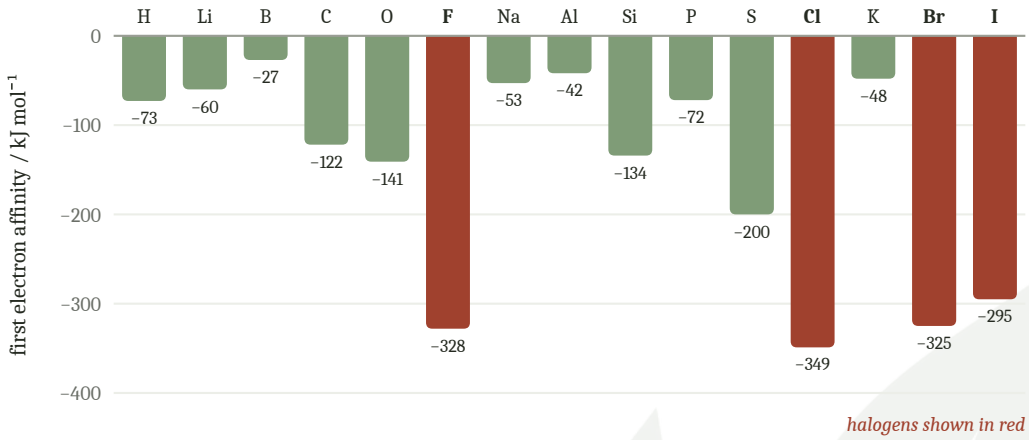


Figure 8 — First electron affinities. The more negative the value, the more energy is released when the atom gains an electron. Noble gases, beryllium, nitrogen and magnesium are not shown (see the box below).

- **Generally more negative across a period.** Z_{eff} increases and the atom is smaller, so the added electron is attracted more strongly. The halogens release the most energy, because one extra electron completes their outer p sublevel.
- **Generally less negative down a group.** The added electron enters a level further from the nucleus and is attracted less strongly (Cl −349, Br −325, I −295 kJ mol^{-1}).
- **Fluorine is an exception:** its value (−328) is less negative than that of chlorine. Fluorine is so small that the added electron is crowded into a small 2p sublevel and repelled strongly by the electrons already there.

Why some elements have no electron affinity value

- **Noble gases:** the outer level is full, so an added electron would have to enter a new, further, well-shielded level. They do not form stable negative ions.
- **Group 2 (Be, Mg), ns^2:** the s sublevel is full, so the added electron must go into a higher-energy p orbital, which is further out and shielded by the ns electrons. The value is close to zero or positive.
- **Nitrogen, 2p^3:** each 2p orbital already holds one electron, so the added electron must pair up and is repelled. The value is close to zero and uncertain.

Second electron affinities are positive

Adding a second electron means forcing an electron onto an ion that is **already negative**. The repulsion must be overcome, so energy has to be supplied: $O^-(g) + e^- \rightarrow O^{2-}(g)$ is endothermic (about +753 kJ mol^{-1}). You will use this in Born–Haber cycles (Reactivity 1.2, HL).

4.6 Electronegativity

Definition to learn

Electronegativity is the ability of an atom to attract a shared pair of electrons in a covalent bond. It has no units; IB uses the scale devised by Linus Pauling in 1932 from bond energy data, on which fluorine, the most electronegative element, has a value of 4.0.

Group	1	2	13	14	15	16	17
Period 1	**H** 2.2						
Period 2	**Li** 1.0	**Be** 1.6	**B** 2.0	**C** 2.6	**N** 3.0	**O** 3.4	**F** 4.0
Period 3	**Na** 0.9	**Mg** 1.3	**Al** 1.6	**Si** 1.9	**P** 2.2	**S** 2.6	**Cl** 3.2
Period 4	**K** 0.8	**Ca** 1.0					**Br** 3.0
Period 5							**I** 2.7
Period 6	**Cs** 0.8						

Figure 9 — Electronegativity values (data booklet, Section 9). Darker cells are more electronegative.

- **Increases across a period** (excluding the noble gases). Z_{eff} increases and the radius decreases, so the nucleus attracts the bonding pair more strongly.
- **Decreases down a group.** The bonding pair is further from the nucleus and more shielded from it.
- The most electronegative elements are in the **top right** (F, O, N, Cl); the least are in the **bottom left** (Cs, Fr). Noble gases are usually not given values, because most of them form no compounds.

Electron affinity or electronegativity?

Students often confuse the two. **Electron affinity** is a measured **energy change** (kJ mol^{-1}) for an **isolated gaseous atom** gaining a whole electron. **Electronegativity** is a **relative number with no units** describing how strongly an atom **in a covalent bond** attracts a **shared** pair of electrons. Ionization energy and electron affinity are properties of **gaseous atoms** and can be measured directly; electronegativity is not.

4.7 All five trends on one page

Property	Across a period	Down a group	Main reason
Atomic radius	↓	↑	across: Z_{eff} increases, same level; down: extra energy level
Ionic radius	↓ *	↑	across: more protons, same electrons (isoelectronic); down: extra energy level. *Within the cations and within the anions; anions are much larger than cations.
First ionization energy	↑ (general)	↓	outer electron held more / less strongly
First electron affinity	more negative (general)	less negative	added electron attracted more / less strongly
Electronegativity	↑	↓	bonding pair attracted more / less strongly
Metallic character	↓	↑	outer electrons lost less / more easily (section 5.1)

Model answer: "Explain why the first ionization energy of magnesium is lower than that of chlorine." (2 marks)

Chlorine has a greater nuclear charge (17 protons compared with 12) but its outer electrons are in the same energy level (n = 3) as those of magnesium, with similar shielding **[1]**. So the effective nuclear charge is greater, the atom is smaller, and the outer electron is held more strongly in chlorine, so more energy is needed to remove it **[1]**.

Model answer: "Explain why a potassium ion is larger than a calcium ion." (2 marks)

K^+ and Ca^{2+} are isoelectronic; both have the configuration 2,8,8 (18 electrons) **[1]**. Calcium has one more proton (20 compared with 19), so its nucleus attracts the same number of electrons more strongly, pulling them closer **[1]**.

Examiner traps in 3.1.3

- "The atoms get bigger across a period because they have more electrons" is **wrong**. They get **smaller**: the extra electrons enter the same level and the nuclear charge increases.
- Always state the reason in terms of **attraction**: name the nucleus and the electron, and say whether the attraction is stronger or weaker. "More shells" alone is not enough; add "further from the nucleus" or "more shielding".
- Gaseous state symbols are required in ionization energy and electron affinity equations. Leaving out (g) loses the mark.
- Ionization energy is per **mole**, not per atom, and it is the **minimum** energy.
- Do not use "electronegativity" when the question is about isolated atoms gaining electrons, and do not give it units.
- In ionic radius questions, compare the **number of protons** for the same number of electrons, or the **number of occupied energy levels**. Do not just say "it has a negative charge".

5 Structure 3.1.4 — Trends down a group: groups 1, 17 and 18 SL + HL

5.1 Metallic and non-metallic character

Metallic character is the tendency of an element to **lose** electrons and form cations. **Non-metallic character** is the tendency to **gain** electrons and form anions (or to attract shared electrons strongly). Both follow from the trends in section 4:

- Down a group, ionization energy and electronegativity **decrease**, so electrons are lost more easily: **metallic character increases**. Group 1 metals become more reactive down the group.
- Down a group, the attraction for an extra electron **decreases**, so **non-metallic character decreases**. Group 17 elements become less reactive down the group. Iodine has a slight metallic lustre, and astatine is thought to be partly metallic.
- Across a period, metallic character decreases and non-metallic character increases: Na, Mg and Al are metals, Si is a metalloid, P, S, Cl and Ar are non-metals.

5.2 Group 18 — the noble gases

The noble gases have a **full outer energy level** ($ns^2\ np^6$, or $1s^2$ for helium). This gives them the highest first ionization energy in each period, and they have no tendency to gain electrons. They exist as **monatomic** gases, held together in the liquid and solid only by weak London dispersion forces, so their boiling points are very low and increase down the group as the atoms get larger. For a long time they were thought to form no compounds at all, but since 1962 compounds of xenon with fluorine and oxygen (such as XeF_4) have been made. Xenon is the most likely to react because its outer electrons are furthest from the nucleus and have the lowest ionization energy in the group.

5.3 Group 1 — the alkali metals

The alkali metals have one valence electron (ns^1), which they lose easily to form **1+ ions** with a noble gas configuration. They are so reactive that they are never found uncombined in nature and are stored under oil to keep out air and water. They form **ionic** compounds with non-metals, and their single mobile outer electron also explains their good electrical and thermal conductivity.

Metal	Atomic radius / pm	First IE / kJ mol^{-1}	Melting point / °C	Density / g cm^{-3}
Li	130	520	181	0.53
Na	160	496	98	0.97
K	200	419	63	0.86
Rb	215	403	39	1.53
Cs	238	376	29	1.93

Physical properties. The alkali metals are soft (they can be cut with a knife), shiny when freshly cut but tarnish quickly in air, and have low densities: Li, Na and K float on water. Their melting points are low and **decrease** down the group, because the cations become larger while the charge stays at 1+, so the metallic bonding gets weaker (Structure 2.3).

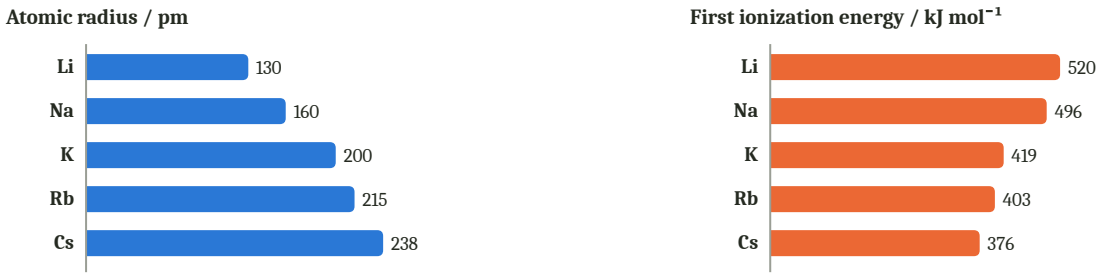


Figure 10 — Down group 1 the atoms get larger and the first ionization energy falls: the outer electron is lost more easily.

Reaction with water

Every alkali metal reacts with water to form a metal hydroxide (an alkali) and hydrogen gas:

$$\mathbf{2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)} \quad \text{e.g. } 2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

$$\text{Ionic equation: } 2M(s) + 2H_2O(l) \rightarrow 2M^+(aq) + 2OH^-(aq) + H_2(g)$$

The OH^- ions make the solution alkaline: universal indicator turns purple (pH about 13–14) and phenolphthalein turns pink. The metal is **oxidized** (it loses an electron) and hydrogen in water is **reduced**.

Metal	Observations when a small piece is added to water
Li	floats; fizzes steadily; slowly gets smaller and disappears; does not melt
Na	floats; melts into a silvery ball (the reaction is exothermic and Na has a low melting point); moves rapidly over the surface; fizzes vigorously; may burn with a yellow-orange flame
K	floats; melts; the hydrogen ignites at once with a **lilac** flame; moves rapidly and may spit or crackle
Rb, Cs	sink (denser than water); react explosively. Too dangerous for a school laboratory

Model answer: "Explain why potassium reacts more vigorously with water than sodium does." (3 marks)

In each reaction the metal atom loses its single outer electron to form a 1+ ion **[1]**. The outer electron of potassium is in a higher energy level (n = 4 compared with n = 3), further from the nucleus and more shielded by inner electrons; the effective nuclear charge is about the same **[1]**. So the outer electron of potassium is held less strongly and is lost more easily: potassium has a lower first ionization energy **[1]**.

5.4 Group 17 — the halogens

The halogens have seven valence electrons ($ns^2\ np^5$). They exist as **diatomic molecules** (F_2, Cl_2, Br_2, I_2) and react by **gaining one electron** to form halide ions, X^-. They form **ionic** compounds with metals and **covalent** compounds with other non-metals. All are toxic. The name *halogen* means "salt maker"; chlorine and iodine are named from the Greek for yellow-green and violet, and bromine from the Greek for "stench".

Halogen	Appearance at room temperature	Electronegativity	Boiling point / °C
F_2	pale yellow gas	4.0	−188
Cl_2	yellow-green gas	3.2	−34
Br_2	dark red-brown liquid, gives off orange vapour	3.0	59
I_2	shiny grey-black solid, sublimes to a violet vapour	2.7	184



Figure 11 — Down group 17 electronegativity falls while boiling point rises (boiling points shown in kelvin so that every bar starts at absolute zero).

Physical trend. The colours get darker and the boiling points rise down the group. The molecules have more electrons, so the **London dispersion forces** between them become stronger (Structure 2.2).

Chemical trend. Reactivity **decreases** down the group. A halogen reacts by attracting an extra electron into its outer level. Down the group the outer level is further from the nucleus and more shielded, while Z_{eff} stays at about +7, so the incoming electron is attracted less strongly. Fluorine is the most reactive non-metal of all; iodine is the least reactive common halogen.

Putting the two groups together: the most vigorous reactions are between a metal from the **bottom** of group 1 and a halogen from the **top** of group 17. Potassium reacts with chlorine far more readily than lithium does with bromine. The most reactive metal (francium) and the most reactive non-metal (fluorine) sit at opposite corners of the table.

Displacement reactions with halide ions

A more reactive halogen **displaces** a less reactive halogen from a solution of its halide ions. The more reactive halogen is **reduced** (it gains electrons) and the halide ion is **oxidized** (it loses electrons). These are redox reactions (Reactivity 3.2).

$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq) \quad (e.g.\ Cl_2 + 2KBr \rightarrow 2KCl + Br_2)$$

$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$$

$$Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$$

Halogen added (rows) to halide solution (columns)	KCl(aq), Cl^-	KBr(aq), Br^-	KI(aq), I^-
Cl_2(aq) (very pale green)	—	**reaction:** solution turns yellow-orange (Br_2)	**reaction:** solution turns brown (I_2)
Br_2(aq) (yellow-orange)	no reaction	—	**reaction:** solution turns brown (I_2)
I_2(aq) (brown)	no reaction	no reaction	—

Because bromine water and dilute iodine solution can look similar, a few drops of a non-polar solvent such as cyclohexane are often added and the mixture shaken. The halogen moves into the upper organic layer, where the colours are distinct: chlorine is very pale green, **bromine is orange** and **iodine is violet (purple)**. The halide salt solutions themselves are colourless.

Model answer: "Chlorine water is added to aqueous potassium iodide. Describe what you would observe, write an ionic equation and explain the reaction." (4 marks)

The colourless solution turns brown (violet in a cyclohexane layer) **[1]**. $Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$ **[1]**. Chlorine is more reactive than iodine: its atoms are smaller, so an incoming electron is closer to the nucleus and is attracted more strongly **[1]**. So chlorine is reduced to chloride ions and the iodide ions are oxidized to iodine **[1]**.

Extension: testing for halide ions with silver nitrate

Adding aqueous silver nitrate to a halide solution acidified with dilute nitric acid gives a precipitate: $Ag^+(aq) + X^-(aq) \rightarrow AgX(s)$. Silver chloride is **white**, silver bromide is **cream** and silver iodide is **yellow**. Silver fluoride is soluble, so fluoride gives no precipitate. This is a useful way to identify the halide ion formed in a displacement reaction.

TOK

Chlorine was used as a chemical weapon in the First World War. Should scientists be held morally responsible for the applications of their discoveries, and should research be subject to ethical limits?

5.5 Why simulations? (Inquiry 2, Tool 2)

The trends in groups 1 and 17 are clearest for the heaviest elements, but those are the ones that cannot be studied safely in a school. Rubidium and caesium explode on contact with water; fluorine attacks almost everything, including glass; astatine and francium are intensely radioactive and only ever made in tiny amounts. Videos and computer simulations let you observe and compare reactions that would be too dangerous, too expensive or impossible to carry out, and to **extrapolate** trends from the elements you can test. Their limitation is that a simulation can only show what its model contains: it is not an experimental result.

Examiner traps in 3.1.4

- The two groups have **opposite** reactivity trends: group 1 reactivity **increases** down the group (electrons lost more easily); group 17 reactivity **decreases** (electrons gained less easily). Explain each one in terms of the relevant process.
- Products of a group 1 metal with water are the **hydroxide** and **hydrogen**, not the oxide.
- Observations must be things you can **see or hear**: "fizzing", "floats", "melts into a ball", "lilac flame". "Hydrogen is produced" is not an observation; "a gas is produced" or "bubbles" is.
- Halogens are diatomic: write Cl_2, not Cl, and balance the equation with $2Br^-$.
- In a displacement reaction the **halogen** is reduced and the **halide ion** is oxidized. Spectator ions (K^+) are left out of ionic equations.
- Fluorine's electron affinity is less exothermic than chlorine's, yet fluorine is the more reactive. Do not use electron affinity values alone to argue about reactivity.

6 Structure 3.1.5 — From metallic to non-metallic: the oxides SL + HL

6.1 A continuum of properties

Metallic and non-metallic properties do not switch over suddenly. Across period 3 they change gradually, and one of the clearest places to see this is in the **oxides**. Most elements form oxides, and whether an oxide behaves as a base or as an acid depends on the bonding in it, and so on the position of the element in the table.

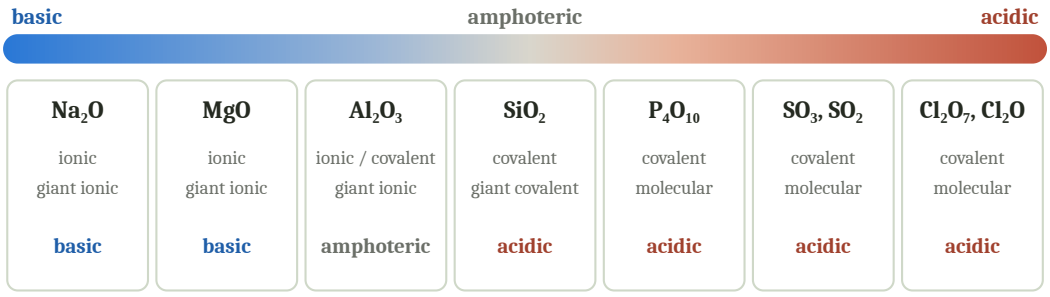


Figure 12 — The highest oxides of the period 3 elements. Ionic metal oxides on the left are basic; covalent non-metal oxides on the right are acidic; aluminium oxide in the middle is amphoteric.

Definitions to learn

- A **basic oxide** reacts with acids to form a salt and water. Soluble basic oxides react with water to form an **alkaline** solution containing OH^- ions. Metal oxides are basic.
- An **acidic oxide** reacts with bases. Soluble acidic oxides react with water to form an **acidic** solution containing H^+ ions. Most non-metal oxides are acidic.
- An **amphoteric oxide** can react both as a base (with acids) and as an acid (with bases). Aluminium oxide is the example to know.

The bonding shows up in a property you can measure: only the ionic oxides conduct electricity when **molten**, because only they contain ions that become free to move.

Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10} / P_4O_6	SO_3 / SO_2	Cl_2O_7 / Cl_2O
Oxidation state of the element	+1	+2	+3	+4	+5 / +3	+6 / +4	+7 / +1
Conductivity when molten	high	high	high	very low	none	none	none
Structure	giant ionic (high melting points)			giant covalent (very high m.p.)	molecular covalent (low melting points)		
Acid–base character	basic		amphoteric	acidic			

Notice that the **highest** oxidation state of each period 3 element in its oxide matches the number of valence electrons: +1 for group 1, +2 for group 2, +3 for group 13, +4 for group 14, and so on up to +7 for chlorine. Oxidation states are the subject of Booklet 2. Down a group the electronegativity of the element falls, so its oxides become **more ionic** and more basic.

6.2 Basic oxides: groups 1 and 2

Group 1 and group 2 oxides react with water to form hydroxides. Learn the general patterns, then apply them to any metal in the group. A base that dissolves in water is called an **alkali**: it forms hydroxide ions in solution.

Group	General equation	Examples
Group 1 (M_2O)	$M_2O(s) + H_2O(l) \rightarrow 2MOH(aq)$	$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$ $Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$ $K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$
Group 2 (MO)	$MO(s) + H_2O(l) \rightarrow M(OH)_2(aq)$	$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$ $MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(s/aq)$ $BaO(s) + H_2O(l) \rightarrow Ba(OH)_2(aq)$

Sodium oxide dissolves to give a strongly alkaline solution (pH about 13–14). Magnesium oxide is only slightly soluble, so the solution is only weakly alkaline (pH about 9–10). Calcium oxide (quicklime) reacts vigorously and exothermically with water to give calcium hydroxide; its solution is limewater. All basic oxides react with acids to form a salt and water, whether or not they dissolve in water. In every case the oxide ion combines with two hydrogen ions:

$$O^{2-}(s) + 2H^+(aq) \rightarrow H_2O(l) \quad \text{e.g. } Li_2O(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2O(l);\ MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$

6.3 Aluminium oxide: amphoteric

Aluminium oxide is insoluble in water, so it has no effect on pH. However, it reacts with acids **and** with strong bases:

$$\text{as a base (with acid): } Al_2O_3(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2O(l)$$

$$\text{e.g. } Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l), \text{ or } Al_2O_3(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$$

$$\text{as an acid (with alkali): } Al_2O_3(s) + 3H_2O(l) + 2OH^-(aq) \rightarrow 2Al(OH)_4^-(aq)$$

$$\text{e.g. } Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq) \text{ (sodium aluminate)}$$

Aluminium oxide behaves like this because its bonding is ionic with significant covalent character (section 6.5).

6.4 Acidic oxides: carbon, sulfur and others

Non-metal oxides that dissolve in water react with it to form **oxoacids**. The equations for carbon and sulfur are required; the others are useful for understanding the trend.

Oxide	Reaction with water	Acid formed
CO_2	$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$	carbonic acid (weak)
SO_2	$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$	sulfurous acid (sulfuric(IV) acid)
SO_3	$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$	sulfuric acid (sulfuric(VI) acid)
P_4O_{10}	$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$	phosphoric acid
N_2O_5	$N_2O_5(s) + H_2O(l) \rightarrow 2HNO_3(aq)$	nitric acid
NO_2	$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$	nitrous and nitric acids
SiO_2	insoluble in water; reacts with hot concentrated NaOH: $SiO_2 + 2NaOH \rightarrow Na_2SiO_3 + H_2O$	none in water, but acidic towards bases

The acids then ionize and release H^+ ions, for example $H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$. Not every non-metal oxide is acidic: carbon monoxide (CO), nitrogen monoxide (NO) and dinitrogen monoxide (N_2O) are **neutral** oxides that do not react with water.

6.5 Why the bonding decides the acid–base character

This answers the linking question with Structure 2.1 and 2.2. Across period 3, the electronegativity difference between the element and oxygen falls, so the bonding in the oxides changes from ionic to covalent (Structure 2.4, bonding triangle).

Ionic metal oxides (Na_2O, MgO, CaO) contain the oxide ion, O^{2-}. This ion is a strong base: it takes a proton from a water molecule, producing hydroxide ions:

$$O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$$

Covalent non-metal oxides (CO_2, SO_2, SO_3) contain no oxide ions. The central atom carries a partial positive charge, and water molecules react with it to form an oxoacid, which then releases H^+ ions into solution.

Aluminium oxide sits between the two: its bonding is ionic with significant covalent character, so it shows both behaviours. Silicon dioxide is a **giant covalent** structure, so although it is acidic, the strong Si–O bonds throughout the lattice prevent it from dissolving in water.

Model answer: "Explain, with reference to bonding, why sodium oxide forms an alkaline solution in water but sulfur trioxide forms an acidic solution." (4 marks)

Sodium oxide is ionic (large electronegativity difference) and contains O^{2-} ions **[1]**. The oxide ions react with water to form hydroxide ions: $Na_2O + H_2O \rightarrow 2NaOH$, or $O^{2-} + H_2O \rightarrow 2OH^-$ **[1]**. Sulfur trioxide is a covalent molecular oxide (small electronegativity difference) with no oxide ions **[1]**. It reacts with water to form sulfuric acid, which releases H^+ ions: $SO_3 + H_2O \rightarrow H_2SO_4$ **[1]**.

6.6 Acid rain

All rain is slightly acidic, with a pH of about **5.6**, because it dissolves carbon dioxide from the air to form carbonic acid. **Acid rain** has a pH below 5.6. It forms when gaseous non-metal oxides, mainly **sulfur dioxide** and **nitrogen oxides** (NO_x), dissolve in water in the atmosphere.

Acid rain and acid deposition

Acid deposition is the wider term for every way acidic substances leave the atmosphere. **Wet deposition**: acidic rain, snow, sleet, hail, fog, mist and dew. **Dry deposition**: acidic gases and particles that settle as dust and smoke and later dissolve in water to form acids. **NO_x** in atmospheric chemistry means the total of NO and NO_2.

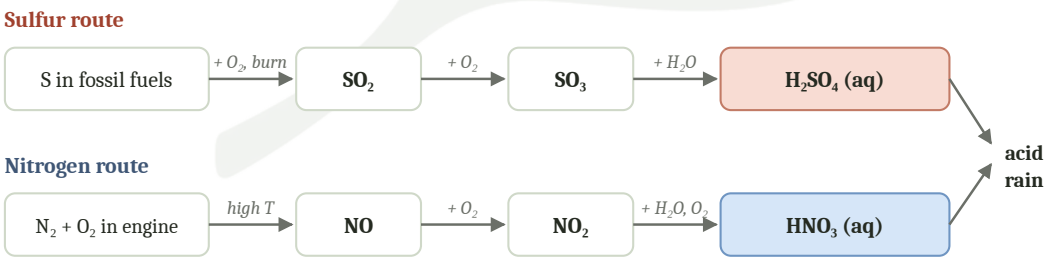


Figure 13 — The two main routes to acid rain.

Sulfur oxides come mainly from burning fossil fuels that contain sulfur, especially coal and heavy fuel oil in power stations (coal is thought to account for about half of global SO_2 emissions), and from smelting sulfide ores; volcanoes are a natural source.

$$S(s) + O_2(g) \rightarrow SO_2(g)$$
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

SO_2 can also dissolve directly to give sulfurous acid, H_2SO_3, which is then oxidized to sulfuric acid in the air.

Nitrogen oxides form when nitrogen and oxygen from the air combine at the very high temperatures inside car engines and power stations; lightning is a natural source.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$
$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$
(or directly: $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$)
$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$
$$4NO_2(g) + O_2(g) + 2H_2O(l) \rightarrow 4HNO_3(aq)$$

Effects. Acid rain reacts with limestone and marble buildings and statues, which are made of calcium carbonate: $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$. It corrodes metals such as iron, lowers the pH of lakes and rivers (harming fish and other aquatic life), and removes nutrients such as Mg^{2+} and Ca^{2+} from soils while releasing toxic Al^{3+} ions, which damages trees.

Reducing it. Remove sulfur from fuels before burning; remove SO_2 from power station waste gases by reacting it with a basic oxide or carbonate, such as $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$; fit catalytic converters, which convert NO into N_2; and switch to renewable energy sources. Acidified lakes can be treated by adding lime (a basic oxide neutralizing an acid).

Nature of Science and TOK

Pollutants can travel hundreds of kilometres before they fall, so one country's emissions can damage another country's lakes and forests. Research into acid deposition therefore led to national laws and international agreements, a clear example of science shaping political and economic decisions. TOK: all rain is acidic, yet "acid rain" refers only to some of it. How does the language scientists choose influence how science is communicated to the public?

6.7 Ocean acidification

The concentration of carbon dioxide in the atmosphere has risen from about 280 ppm before the industrial revolution to more than 420 ppm today, mainly because of fossil fuel combustion. The oceans absorb a large share of the CO_2 released by human activity (estimates are typically a quarter to a third; some textbooks quote higher figures). As it dissolves, it lowers the pH of seawater:

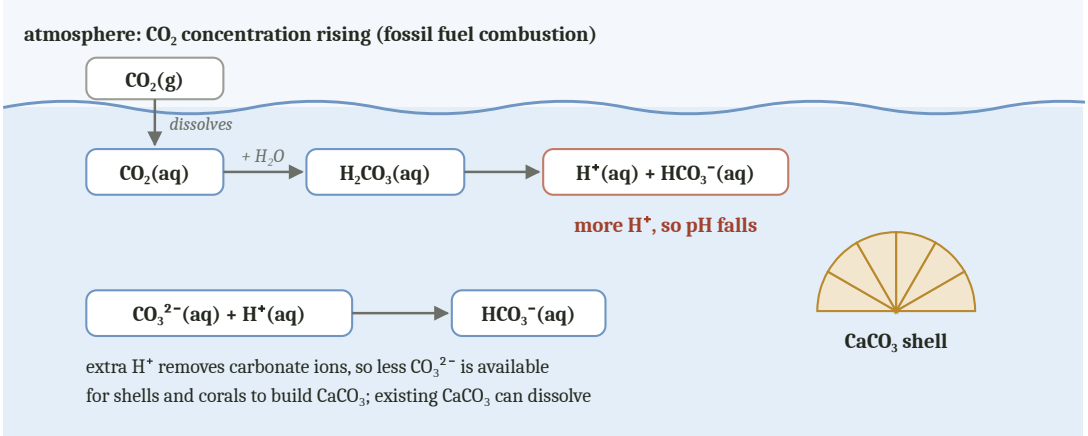


Figure 14 — Dissolved CO_2 forms carbonic acid, which releases H^+ ions. The extra H^+ ions convert carbonate ions into hydrogencarbonate ions.

$$CO_2(g) \rightleftharpoons CO_2(aq)$$
$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$
$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$
$$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$$

All of these are equilibria. Adding more CO_2 pushes the first three to the right and increases $[H^+]$. The extra H^+ ions then push the last equilibrium to the **left**: they combine with carbonate ions to form hydrogencarbonate, $H^+(aq) + CO_3^{2-}(aq) \rightarrow HCO_3^-(aq)$, so the carbonate concentration falls.

The average pH of surface seawater has fallen from about 8.2 to about 8.1. That sounds small, but pH is a logarithmic scale, so it means the concentration of H^+ ions has increased by roughly 25–30%. Seawater is still alkaline: "acidification" describes the **direction** of the change, not a pH below 7.

Why it matters. Corals, molluscs and some plankton build their shells and skeletons from **calcium carbonate**, using carbonate ions from seawater. The extra H^+ ions remove carbonate ions, so less carbonate is available to build shells. In more acidic water, existing calcium carbonate can also dissolve: $CaCO_3(s) + CO_2(aq) + H_2O(l) \rightarrow Ca^{2+}(aq) + 2HCO_3^-(aq)$. Damage to coral reefs and plankton affects whole marine food chains, and higher acidity is also suspected of causing reproductive problems in some fish. The effect is greatest near the surface, where the CO_2 is absorbed.

Examiner traps in 3.1.5

- Do not write "$Na_2O + H_2O \rightarrow Na_2OH$". Each Na^+ pairs with one OH^-, so the product is **2NaOH**. Group 2 gives $M(OH)_2$.
- The reaction of CO_2 with water is **reversible**: use $\rightleftharpoons$. Carbonic acid is a weak acid.
- SO_2 gives **sulfurous** acid, H_2SO_3; SO_3 gives **sulfuric** acid, H_2SO_4. Check that the number of oxygen atoms balances.
- Rain with pH 5.6 is **not** acid rain; that is normal, unpolluted rain. Carbon dioxide is **not** a significant cause of acid rain.
- Nitrogen oxides come from N_2 in the **air** reacting at high temperature, not from nitrogen in the fuel.
- Explaining the acid–base trend requires **bonding**: ionic oxide → O^{2-} ions → OH^-; covalent oxide → oxoacid → H^+. "Because it is a metal" is not an explanation.

7 Key points to memorise

Everything below is either wording that examiners look for, or knowledge you must carry into the exam yourself.

A Definitions — near-verbatim

- **Period:** a horizontal row; the period number is the outer occupied energy level. **Group:** a vertical column; same number of valence electrons.
- **Block:** region of the table named after the sublevel (s, p, d, f) being filled.
- **Effective nuclear charge:** the net nuclear charge felt by an outer electron, after shielding by inner electrons.
- **Atomic radius:** half the distance between the nuclei of two bonded atoms of the same element.
- **First ionization energy:** minimum energy to remove one mole of electrons from one mole of gaseous atoms in the ground state. $X(g) \rightarrow X^{+}(g) + e^{-}$
- **First electron affinity:** energy change when one mole of electrons is added to one mole of gaseous atoms. $X(g) + e^{-} \rightarrow X^{-}(g)$
- **Electronegativity:** the ability of an atom to attract a shared pair of electrons in a covalent bond.
- **Amphoteric:** able to react as both an acid and a base.

B The trends and their single explanation

- Across a period: nuclear charge up, same level, similar shielding → Z_{eff} up → radius down; IE, EN up; EA more negative; metallic character down.
- Down a group: extra level, more shielding, Z_{eff} about constant → radius up; IE, EN down; EA less negative; metallic character up.
- Cations smaller than atoms (outer level lost); anions larger than atoms (more repulsion). Isoelectronic: more protons, smaller.
- Exceptions to remember: F has a less exothermic EA than Cl; IE dips at groups 2→13 and 15→16 (HL 3.1.7).

C Groups 1, 17 and 18

- Group 1 + water: $2M + 2H_2O \rightarrow 2MOH + H_2$. Li fizzes; Na melts into a ball and darts; K ignites with a lilac flame. Reactivity **increases** down the group (IE decreases).
- Halogens: F_2 pale yellow gas, Cl_2 yellow-green gas, Br_2 red-brown liquid, I_2 grey-black solid. Reactivity **decreases** down the group (weaker attraction for an added electron).
- Displacement: $Cl_2 + 2Br^{-} \rightarrow 2Cl^{-} + Br_2$; $Cl_2 + 2I^{-} \rightarrow 2Cl^{-} + I_2$; $Br_2 + 2I^{-} \rightarrow 2Br^{-} + I_2$. In cyclohexane: Br_2 orange, I_2 violet.
- Noble gases: full outer level, monatomic, very unreactive; xenon forms some compounds.

D Oxide equations

- $Na_2O + H_2O \rightarrow 2NaOH$ | $MgO + H_2O \rightarrow Mg(OH)_2$ | $CaO + H_2O \rightarrow Ca(OH)_2$
- $CO_2 + H_2O \rightleftharpoons H_2CO_3$ | $SO_2 + H_2O \rightarrow H_2SO_3$ | $SO_3 + H_2O \rightarrow H_2SO_4$
- Period 3: Na_2O, MgO basic; Al_2O_3 amphoteric; SiO_2, P_4O_{10}, SO_3, Cl_2O_7 acidic.
- Acid rain: pH < 5.6, from SO_2 (sulfur in fuels) and NO_x ($N_2 + O_2$ in engines). Ocean acidification: more CO_2 → more H^+ → less CO_3^{2-} for $CaCO_3$ shells.

E Command terms — what each one is actually asking for

- **Identify** (the position of a metalloid) — pick out the answer; no explanation needed.
- **Deduce** (an electron configuration or an equation) — reach the answer by reasoning from the information given; show the working, for example the noble gas core and the sublevels.
- **Describe** (a reaction) — give the observations: what you see, hear or feel, with colours and states.
- **Explain** (a trend) — give the cause: nuclear charge, shielding, distance and the strength of attraction. Compare both species.
- **Outline** (the causes of acid rain) — a brief account: name the gases, their sources and the acids formed.

8 Check yourself

Cover the answers on the next pages. Marks are shown in brackets where a question is written in exam style. Everything in this booklet is for both SL and HL students.

1. Which element is a metalloid?

   A aluminium B germanium C selenium D tin

2. An element is in period 4 and group 15. Identify the element, state its block, and deduce its full electron configuration. **[3]**

3. An atom has the configuration [Ne] $3s^2\ 3p^4$. State its period, group and block, and identify the element. **[2]**

4. Explain why helium is placed in group 18 rather than group 2. **[2]**

5. Explain how Mendeleev's periodic table allowed new elements to be discovered. **[2]**

6. Which atom has the largest atomic radius?

   A Na B Mg C K D Ca

7. Explain why the atomic radius decreases from sodium to chlorine. **[2]**

8. Explain why a sodium ion is smaller than a sodium atom, but a chloride ion is larger than a chlorine atom. **[3]**

9. Which shows the ions in order of increasing radius?

   A $O^{2-} < F^- < Na^+ < Mg^{2+}$ B $Mg^{2+} < Na^+ < F^- < O^{2-}$ C $Na^+ < Mg^{2+} < O^{2-} < F^-$ D $F^- < O^{2-} < Mg^{2+} < Na^+$

10. Define the first ionization energy of an element, and write an equation, with state symbols, for the first ionization of magnesium. **[3]**

11. Explain why the first ionization energy decreases down group 1. **[2]**

12. Which equation represents the first electron affinity of chlorine?

A $Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$ B $Cl(g) \rightarrow Cl^+(g) + e^-$ C $Cl(g) + e^- \rightarrow Cl^-(g)$ D $Cl(aq) + e^- \rightarrow Cl^-(aq)$

13. Explain why the second electron affinity of oxygen is endothermic. **[2]**

14. Which properties increase down group 17? I. atomic radius II. electronegativity III. boiling point

A I and II only B I and III only C II and III only D I, II and III

15. A small piece of potassium is added to water containing universal indicator. Describe **three** observations and write an equation, with state symbols. **[4]**

16. Explain why rubidium is more reactive than sodium. **[3]**

17. Which mixture results in a reaction?

A $Br_2(aq) + KCl(aq)$ B $I_2(aq) + KBr(aq)$ C $Cl_2(aq) + KBr(aq)$ D $I_2(aq) + KCl(aq)$

18. Bromine water is added to aqueous sodium iodide. Write an ionic equation, state the colour change, and identify the species oxidized and the species reduced. **[4]**

19. Explain why the reactivity of the halogens decreases down group 17. **[2]**

20. Write balanced equations for the reactions with water of: (a) lithium oxide (b) calcium oxide (c) sulfur trioxide (d) carbon dioxide. **[4]**

21. Which oxide is amphoteric?

A MgO B Al_2O_3 C SiO_2 D P_4O_{10}

22. Explain, with reference to bonding, why magnesium oxide is basic but carbon dioxide is acidic. **[3]**

23. (a) State why unpolluted rain has a pH of about 5.6. (b) Identify **one** sulfur-containing and **one** nitrogen-containing pollutant that cause acid rain, and state the source of each. (c) Write an equation for the reaction of acid rain with marble. **[5]**

24. Explain how an increasing concentration of atmospheric carbon dioxide affects marine organisms that build shells of calcium carbonate. Include **two** equations. **[4]**

25. How many valence electrons are present in an atom of the element with atomic number 51? **[1]**

26. Which are properties of gaseous atoms? I. ionization energy II. electron affinity III. electronegativity

A I and II only B I and III only C II and III only D I, II and III

27. Which changes are endothermic? I. $Ca(g) \rightarrow Ca^+(g) + e^-$ II. $I(g) + e^- \rightarrow I^-(g)$ III. $O^-(g) + e^- \rightarrow O^{2-}(g)$

A I and II only B I and III only C II and III only D I, II and III

28. List Cl, Cl^+ and Cl^- in order of decreasing radius, and explain the order. **[2]**

29. The radius of the Si^{4+} ion is about 40 pm, but that of Si^{4-} is about 271 pm. Explain the large difference. **[2]**

30. Which pair of elements reacts most readily?

A Li + Br_2 B Li + Cl_2 C K + Br_2 D K + Cl_2

31. Which oxides produce an acidic solution when added to water? I. SiO_2 II. N_2O_5 III. SO_2

A I and II only B I and III only C II and III only D I, II and III

32. Write ionic equations to show the amphoteric behaviour of aluminium oxide with (a) $H^+(aq)$ and (b) $OH^-(aq)$. **[2]**

9 Answers

1. **B.** Al and Sn are metals; Se is a non-metal.
2. Arsenic **[1]**; p block **[1]**; $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$ **[1]**.
3. Period 3, group 16, p block **[1]**; sulfur **[1]**.
4. Helium has a full outer energy level ($1s^2$) **[1]**, so it is unreactive and has the properties of the other noble gases, not of the reactive group 2 metals **[1]**.
5. Mendeleev left gaps for undiscovered elements **[1]** and predicted their properties from those of neighbouring elements, which helped chemists to search for and recognize them, for example gallium and germanium **[1]**.
6. **C.** K is in period 4 and group 1 (one more level than Na; lower Z_{eff} than Ca).
7. The nuclear charge increases while the outer electrons are added to the same energy level, with similar shielding **[1]**, so the effective nuclear charge increases and the outer electrons are pulled closer to the nucleus **[1]**.
8. $Na \rightarrow Na^+$: the outer (n = 3) energy level is lost, so the remaining electrons are closer to the nucleus **[1]**; the same nuclear charge now attracts fewer electrons with less repulsion **[1]**. $Cl \rightarrow Cl^-$: an extra electron is added to the outer level with the same nuclear charge, so the repulsion between electrons increases and the electron cloud expands **[1]**.
9. **B.** All four have 10 electrons; more protons means a smaller ion.
10. The minimum energy needed to remove one mole of electrons **[1]** from one mole of gaseous atoms in the ground state **[1]**. $Mg(g) \rightarrow Mg^+(g) + e^-$ **[1]**.
11. Down the group the outer electron is in a higher energy level, further from the nucleus and more shielded by inner electrons **[1]**; the effective nuclear charge stays about the same, so the outer electron is attracted less strongly and less energy is needed to remove it **[1]**.
12. **C.** A involves Cl_2 molecules and two electrons; B is ionization; D is not in the gas phase.
13. The O^- ion is already negatively charged **[1]**, so it repels the incoming electron and energy must be supplied to overcome this repulsion **[1]**.
14. **B.** Electronegativity decreases down the group.
15. Any three of: floats; melts into a ball; moves around on the surface; fizzes; burns with a lilac flame; indicator turns purple **[3 max]**. $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ **[1]**.
16. Both react by losing their single outer electron **[1]**. The outer electron of rubidium is in a higher energy level (n = 5), further from the nucleus and more shielded; the effective nuclear charge is about the same **[1]**. So it is held less strongly and lost more easily: Rb has a lower first ionization energy **[1]**.
17. **C.** Chlorine is more reactive than bromine.
18. $Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$ **[1]**. Yellow-orange to brown **[1]**. Iodide ions are oxidized **[1]**; bromine is reduced **[1]**.
19. Down the group the outer energy level is further from the nucleus and more shielded (Z_{eff} about +7 throughout) **[1]**, so an incoming electron is attracted less strongly and the atom gains an electron less readily **[1]**.
20. (a) $Li_2O + H_2O \rightarrow 2LiOH$ (b) $CaO + H_2O \rightarrow Ca(OH)_2$ (c) $SO_3 + H_2O \rightarrow H_2SO_4$ (d) $CO_2 + H_2O \rightleftharpoons H_2CO_3$ **[1]** each.
21. **B.**
22. MgO is ionic and contains O^{2-} ions, which react with water or acids as a base: $O^{2-} + H_2O \rightarrow 2OH^-$ **[1]**. CO_2 is a covalent molecular oxide with no oxide ions **[1]**; it reacts with water to form carbonic acid, which releases H^+ ions **[1]**.

23. (a) Carbon dioxide from the air dissolves to form (weak) carbonic acid **[1]**. (b) Sulfur dioxide, from burning sulfur-containing fossil fuels such as coal **[1]**; nitrogen monoxide or nitrogen dioxide, from N_2 and O_2 reacting at high temperature in engines **[1]**. (c) $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$, or the equivalent with $2HNO_3$ **[2]**, with 1 mark for correct formulas and 1 for balancing.

24. More CO_2 dissolves in seawater and forms carbonic acid: $CO_2 + H_2O \rightleftharpoons H_2CO_3$ **[1]**, which releases H^+ ions: $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$, so the pH falls **[1]**. The H^+ ions react with carbonate ions, $H^+ + CO_3^{2-} \rightarrow HCO_3^-$, so less carbonate is available to build $CaCO_3$ **[1]**, and existing shells and coral skeletons can dissolve **[1]**.

25. Five. Z = 51 is antimony, [Kr] $4d^{10}\ 5s^2\ 5p^3$, in group 15; the full 4d sublevel does not count.

26. **A.** Electronegativity describes an atom in a covalent bond, not a gaseous atom.

27. **B.** Ionization always needs energy; adding an electron to a negative ion needs energy to overcome repulsion. II is exothermic.

28. Cl^- > Cl > Cl^+ **[1]**. All have 17 protons; the more electrons there are, the greater the electron–electron repulsion and the larger the electron cloud **[1]**.

29. Si^{4+} has lost its third energy level (2,8), while Si^{4-} has a full third level (2,8,8) **[1]**. Si^{4-} has 18 electrons held by only 14 protons, so there is much more repulsion, while Si^{4+} has 10 electrons strongly attracted by 14 protons **[1]**.

30. **D.** Potassium is the more reactive metal and chlorine the more reactive halogen.

31. **C.** SiO_2 is acidic but insoluble in water, so it does not change the pH.

32. (a) $Al_2O_3 + 6H^+ \rightarrow 2Al^{3+} + 3H_2O$ **[1]** (b) $Al_2O_3 + 3H_2O + 2OH^- \rightarrow 2Al(OH)_4^-$ **[1]**.

Where this fits in Structure 3.1

- **Booklet 1** (this booklet) — 3.1.1 to 3.1.5: the table, configuration and position, periodicity, group trends and the oxides.
- **Booklet 2** — 3.1.6 oxidation states (SL + HL) and HL 3.1.7: discontinuities in first ionization energy.
- **Booklet 3** — HL 3.1.8 to 3.1.10: the transition elements, their variable oxidation states and coloured complexes.
- Links back: **Structure 1.3** (electron configurations), **Structure 2.4** (period 3 oxides on the bonding triangle). Links forward: **Reactivity 3.1** (acids and bases), **Reactivity 3.2** (redox).

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