Structure 3.1 · The periodic table: classification of elements

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

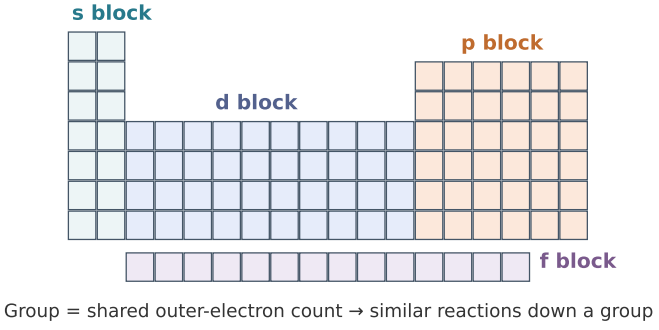
This handout teaches Structure 3.1, The periodic table: classification of elements, for the course first assessed in 2025. Sections marked **AHL** are for Higher Level only.

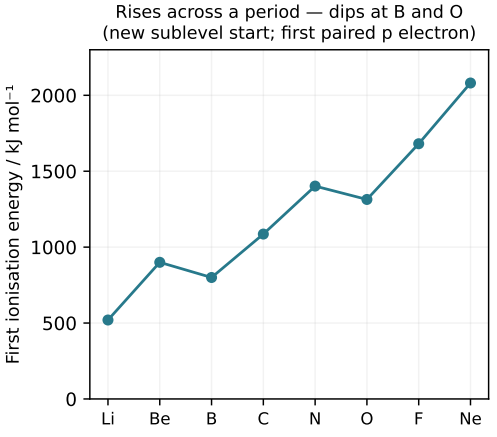
Prerequisites: electron configurations and ionization energy from Structure 1.3, and ionic and covalent bonding from Structures 2.1 and 2.2. You will use the periodic table, ionic radii and electronegativity values from the data booklet.

Organizing the periodic table

One table, many predictions

In 1869 Dmitri Mendeleev arranged the known elements in order of atomic mass and left gaps where the pattern said an element should be. He predicted the properties of "eka-silicon", and germanium, found in 1886, matched them closely. The **periodic table** works because the properties of elements repeat in a regular way: **periodicity** .





The table follows the order in which sublevels fill: s, p, d and f blocks

Periods, groups and blocks

The modern table orders the elements by atomic number Z.

- A **period** is a horizontal row. The period number is the number of the outer energy level: sodium, $[Ne]\,3s^1$, is in period 3.

- A **group** is a vertical column, numbered 1 to 18. Elements in a group have the same number of outer electrons, so they react in similar ways. For groups 1 and 2 the number of valence electrons equals the group number; for groups 13 to 18 it is the group number minus 10.
- A **block** is named after the sublevel being filled: s (groups 1 and 2), p (groups 13 to 18), d (groups 3 to 12) and f (the lanthanoids and actinoids).

Worked example: chlorine is $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$. The outer level is $n = 3$, so period 3; the last electron enters a p sublevel, so p block; seven outer electrons, so group 17.

Metals, metalloids and non-metals

Metals are on the left and in the middle of the table; non-metals are on the right. A diagonal band of **metalloids** (B, Si, Ge, As, Sb, Te) separates them. Metalloids have some properties of both: silicon looks shiny but is brittle, and it is a semiconductor.

Some groups have names: group 1 the **alkali metals**, group 2 the alkaline earth metals, group 17 the **halogens** and group 18 the noble gases.

Periodic trends

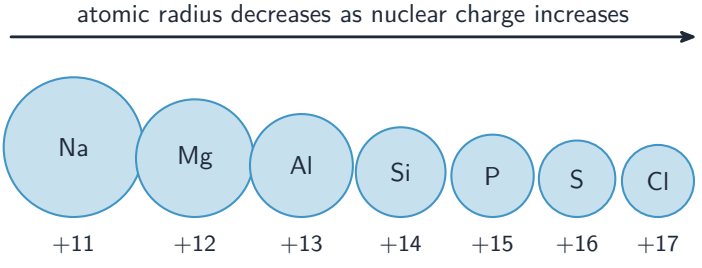
Two ideas explain every trend

Every trend in this section comes from the balance of two effects on the outer electrons:

- **nuclear charge**: more protons pull the electrons more strongly;
- **shielding**: inner electrons repel the outer electrons and reduce the pull they feel.

Across a period, the nuclear charge rises by one each time but the electrons go into the **same** energy level, so shielding hardly changes. The outer electrons feel a stronger pull. Down a group, each element has one more full **inner** level: the outer electrons are further away and more shielded, so they feel a weaker pull, even though the nuclear charge is larger.

Atomic radius



Across period 3 the growing nuclear charge pulls the same outer level closer

- **Across a period** the atomic radius decreases: Na 186 pm to Cl 102 pm.
- **Down a group** the atomic radius increases: one extra occupied energy level each time.

Ionic radius

- A **cation** is smaller than its atom: it has lost its whole outer level, and the remaining electrons are pulled in more strongly. Na 186 pm, Na^+ 102 pm.
- An **anion** is larger than its atom: the extra electrons repel each other, and the same nuclear charge holds more electrons. Cl 102 pm, Cl^- 181 pm.
- For ions with the **same electron configuration** (isoelectronic ions), the radius falls as the nuclear charge rises: $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$, all with the configuration of neon.

Ionization energy, electron affinity and electronegativity

Property	Meaning	Across a period	Down a group
first ionization energy	energy to remove one mole of electrons from one mole of gaseous atoms	generally increases	decreases
electron affinity	energy change when one mole of gaseous atoms each gain one electron	generally more negative (more released)	generally less negative
electronegativity	ability of an atom to attract a shared pair of electrons	increases	decreases

All three follow from the same idea: a stronger pull on the outer electrons makes them harder to remove and makes the atom attract extra electrons more strongly. The first electron affinity is usually negative (exothermic), because the incoming electron is attracted by the nucleus.

Fluorine is the most electronegative element (4.0); caesium and francium are the least.

Metallic and non-metallic character

Metals react by **losing** electrons. Metallic character therefore increases as ionization energy decreases: **down** a group and **to the left** across a period. Non-metallic character increases up and to the right. This is why the most reactive metals (Cs, Fr) are at the bottom left and the most reactive non-metal (F) is at the top right.

Group 1 and group 17

The alkali metals

Group 1 metals are soft, have low melting points and react quickly with water to give a hydroxide and hydrogen:

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

Reactivity **increases** down the group. The outer electron is further from the nucleus and more shielded, so it is lost more easily. Lithium fizzes steadily; potassium melts, catches fire and burns with a lilac flame.

The halogens

Halogens exist as diatomic molecules: F_2 and Cl_2 are gases, Br_2 a liquid and I_2 a solid at room temperature (London forces increase down the group). Halogens react by **gaining** an electron, so their reactivity **decreases** down the group: the incoming electron is further from the nucleus and more shielded.

A more reactive halogen **displaces** a less reactive one from a solution of its halide ions:

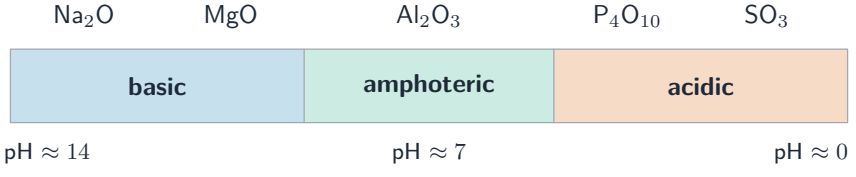
$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$$

The solution turns orange-yellow as bromine forms. Bromine cannot displace chloride ions, because chlorine is more reactive.

Oxides of the elements

From basic to acidic

Metal oxides are mostly ionic and **basic** : they react with acids, and those that dissolve give alkaline solutions. Non-metal oxides are covalent and **acidic** : they react with bases and dissolve to give acidic solutions.



metal oxides (left) are basic; non-metal oxides (right) are acidic

Across period 3 the oxides change from basic through amphoteric to acidic

Oxide	Bonding	Reaction with water	Character
Na_2O	ionic	$Na_2O + H_2O \rightarrow 2NaOH$	basic
MgO	ionic	$MgO + H_2O \rightarrow Mg(OH)_2$ (slightly soluble)	basic
Al_2O_3	ionic with covalent character	insoluble	**amphoteric**
SiO_2	giant covalent	insoluble	acidic (reacts with hot NaOH)
P_4O_{10}	molecular covalent	$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$	acidic
SO_3	molecular covalent	$SO_3 + H_2O \rightarrow H_2SO_4$	acidic

An amphoteric oxide reacts with both acids and bases:

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O \qquad Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$$

■ Acid deposition and ocean acidification

Non-metal oxides in the air make rain more acidic. Burning fuels that contain sulfur forms SO_2, and the high temperature in engines forms nitrogen oxides. In the air they form sulfuric and nitric acid:

$$SO_2 + H_2O \rightarrow H_2SO_3 \qquad 4NO_2 + O_2 + 2H_2O \rightarrow 4HNO_3$$

This **acid deposition** damages limestone buildings, forests and lakes. Carbon dioxide dissolving in the oceans forms carbonic acid, $CO_2 + H_2O \rightleftharpoons H_2CO_3$, which lowers the pH of seawater: **ocean acidification** . It makes it harder for corals and shellfish to build calcium carbonate shells.

Oxidation states

■ Rules for assigning oxidation states

An **oxidation state** is the charge an atom would have if all its bonds were ionic. It is written with the sign first: +2, −1.

1. An uncombined element is 0: Fe, O_2, Cl_2.
2. A simple ion has its charge: Na^+ is +1, O^{2-} is −2.
3. In a compound, oxygen is usually −2 (except in peroxides, −1, and with fluorine), hydrogen is usually +1 (except in metal hydrides, −1), and fluorine is always −1.
4. The oxidation states add up to zero in a compound, and to the charge in an ion.

Worked example: the oxidation state of Mn in MnO_4^-. Four oxygens give $4(-2) = -8$. The total must be -1, so Mn $+ (-8) = -1$, and Mn is $+7$.

Naming with Roman numerals

When an element can have more than one oxidation state, the name shows it in Roman numerals: iron(II) chloride is $FeCl_2$, iron(III) chloride is $FeCl_3$. Oxyanions are named the same way: SO_4^{2-} is sulfate(VI) and SO_3^{2-} is sulfate(IV), commonly called sulfite; NO_3^- is nitrate(V).

AHL: Transition elements

What makes an element a transition element

A **transition element** is an element that has an incomplete d sublevel, or that forms at least one stable ion with an incomplete d sublevel. In the first row, titanium to copper are transition elements.

- Scandium forms only Sc^{3+}, [Ar], with an empty d sublevel: **not** a transition element.
- Zinc forms only Zn^{2+}, $[Ar]\,3d^{10}$, with a full d sublevel: **not** a transition element.

When a transition element forms an ion, it loses its **4s electrons first**: Fe is $[Ar]\,3d^6\,4s^2$ and Fe^{2+} is $[Ar]\,3d^6$.

Characteristic properties

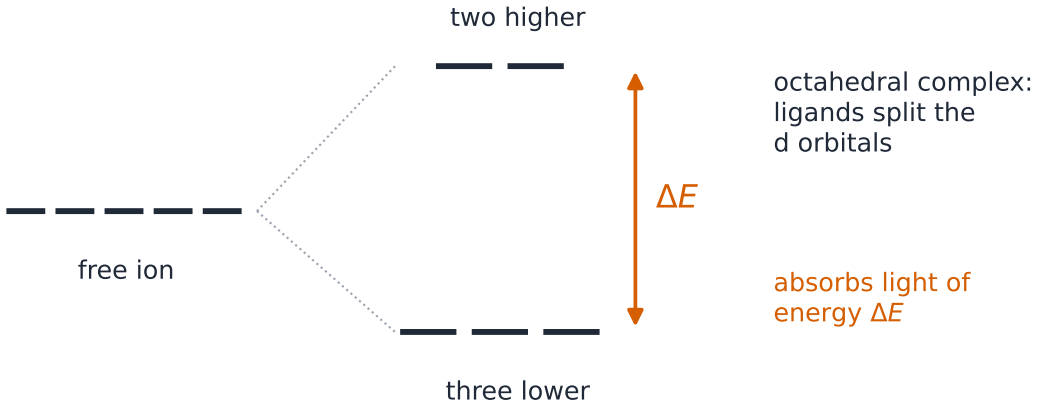
Transition elements share four properties:

1. **Variable oxidation states.** The 4s and 3d sublevels are close in energy, so successive ionization energies rise gradually and different numbers of electrons can be lost. Manganese shows every state from +2 to +7, as in Mn^{2+}, MnO_2 and MnO_4^-.
2. **Complex ions.** A central metal ion bonds to **ligands** : molecules or ions with a lone pair that form coordination bonds to the metal, such as H_2O, NH_3, Cl^- and CN^-. The number of coordination bonds is the **coordination number** . Example: $[Cu(H_2O)_6]^{2+}$ is octahedral, with coordination number 6.
3. **Coloured compounds**, explained below.
4. **Catalytic activity.** Variable oxidation states and the ability to adsorb molecules on the surface make transition elements good catalysts: iron in the Haber process, V_2O_5 in the Contact process, nickel in the hydrogenation of alkenes, and MnO_2 in the decomposition of hydrogen peroxide.

The charge on a complex ion is the oxidation state of the metal plus the charges of the ligands. In $[CuCl_4]^{2-}$, four Cl^- give -4, so copper is +2.

Why transition metal complexes are coloured

In a free ion the five d orbitals have the same energy. In a complex, the lone pairs of the ligands repel the d electrons. In an octahedral complex, two of the d orbitals point straight at the ligands and are raised in energy; the other three are lowered. The d orbitals **split** into two sets separated by an energy gap ΔE.



An electron absorbs light of energy ΔE to move to a higher d orbital

An electron in a lower d orbital absorbs a photon of visible light with energy equal to ΔE and moves to a higher d orbital. The light that is **not** absorbed passes through, so we see the **complementary colour** . $[Cu(H_2O)_6]^{2+}$ absorbs orange-red light and looks blue.

The size of ΔE, and so the colour, depends on:

- the **metal** and its **oxidation state**: Fe^{2+} solutions are pale green, Fe^{3+} yellow-brown;
- the **ligand**: adding ammonia to blue $[Cu(H_2O)_6]^{2+}$ gives deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$, because NH_3 splits the d orbitals more strongly;
- the **coordination number** and shape of the complex.

There is no colour if no d-to-d transition is possible: Sc^{3+} ($3d^0$) and Zn^{2+} ($3d^{10}$) compounds are white or colourless.

Magnetic properties

Electrons spin, and an unpaired electron behaves like a tiny magnet.

- A substance with **unpaired** electrons is **paramagnetic** : it is attracted into a magnetic field. The more unpaired electrons, the stronger the attraction. Fe^{3+}, $3d^5$, has five unpaired electrons and is strongly paramagnetic.
- A substance with **all electrons paired** is **diamagnetic** : it is weakly repelled by a magnetic field. Zn^{2+}, $3d^{10}$, is diamagnetic.

Iron, cobalt and nickel are also **ferromagnetic**: the unpaired electrons of neighbouring atoms line up, so the metals can become permanent magnets.

Practice

1. Without a periodic table, state the period, group and block of the element with configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$.
2. Explain why the atomic radius decreases from sodium to chlorine.
3. Put these in order of increasing radius and explain: Mg^{2+}, Na^+, F^-, O^{2-}.
4. Explain why potassium reacts more vigorously with water than sodium.
5. Write an equation for the reaction of chlorine with aqueous potassium iodide, and state what you would see.
6. Write equations to show the reaction of (a) sodium oxide and (b) sulfur trioxide with water, and state the effect of each on the pH.
7. Deduce the oxidation state of: (a) Cr in $Cr_2O_7^{2-}$, (b) S in $Na_2S_2O_3$, (c) N in NH_4^+.
8. **AHL.** Explain why zinc is a d-block element but not a transition element.
9. **AHL.** State the oxidation state of the metal and the coordination number in $[Fe(CN)_6]^{3-}$.
10. **AHL.** Explain why $[Cu(H_2O)_6]^{2+}$ is blue, and why the colour changes when ammonia is added.
11. **AHL.** Deduce the number of unpaired electrons in Mn^{2+} and state whether it is paramagnetic or diamagnetic.

Answers.

1. Period 3 (outer level $n = 3$), group 16 (six outer electrons), p block. It is sulfur.
2. The nuclear charge increases from +11 to +17, but the outer electrons stay in the same energy level, with about the same shielding. The outer electrons are pulled closer to the nucleus.
3. $Mg^{2+} < Na^+ < F^- < O^{2-}$. All four have the neon configuration (10 electrons). The nuclear charges are 12, 11, 9 and 8; the larger the nuclear charge, the more strongly the same electrons are pulled in.
4. Potassium's outer electron is in the fourth energy level, further from the nucleus and more shielded than sodium's. It is held less strongly, so it is lost more easily.
5. $Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$. The colourless solution turns brown as iodine forms.
6. (a) $Na_2O + H_2O \rightarrow 2NaOH$: the pH rises (alkaline, about 14). (b) $SO_3 + H_2O \rightarrow H_2SO_4$: the pH falls (strongly acidic, about 0 to 1).
7. (a) $2Cr + 7(-2) = -2$, so Cr is +6. (b) $2(+1) + 2S + 3(-2) = 0$, so S is +2. (c) $N + 4(+1) = +1$, so N is −3.

8. **AHL.** Zinc is in the d block because its last electrons fill the 3d sublevel. But its only ion, Zn^{2+}, is $[Ar]\,3d^{10}$, and the atom is $[Ar]\,3d^{10}\,4s^2$: neither has an incomplete d sublevel.
9. **AHL.** Six CN^- ligands give -6; the total is -3, so iron is $+3$. Coordination number 6.
10. **AHL.** The water ligands split the 3d orbitals of Cu^{2+}. An electron absorbs orange-red light of energy ΔE to move to a higher d orbital, and we see the complementary colour, blue. Ammonia ligands split the d orbitals by a different amount, so a different wavelength is absorbed and the colour changes to deep blue.
11. **AHL.** Mn is $[Ar]\,3d^5\,4s^2$; Mn^{2+} loses the 4s electrons, giving $3d^5$. The five electrons occupy the five d orbitals singly: five unpaired electrons, so it is paramagnetic.

Exam tips

- Explain every trend with **nuclear charge**, **shielding** and **distance**; say what happens to the attraction for the outer electrons.
- Across a period, electrons go into the **same** energy level; down a group, a **new** level is added.
- Isoelectronic ions: compare nuclear charge only; the electron number is the same.
- Metal oxides are basic, non-metal oxides acidic, Al_2O_3 amphoteric; write equations for both of its reactions.
- Write oxidation states with the sign **first** ($+3$), and charges with the sign last ($3+$).
- **AHL.** Transition metals lose **4s** electrons before 3d electrons.
- **AHL.** For colour, say that the d orbitals **split**, light of energy ΔE is **absorbed**, and the **complementary** colour is seen; never say the complex "emits" its colour.