

3.1 The periodic table

3.1.1 Periodicity

- Definitions

Term	Definition
Periodicity	A repeating trend in physical and chemical properties of the elements across the periodic table.
Groups	A vertical column in the periodic table. Elements in a group have similar chemical properties and their atoms have the same number of outer shell electrons.
Periods	A horizontal row in the periodic table. Elements show trends in properties across a period.
Shielding effect	The repulsion between electrons in different inner shells. Shielding reduces the net attractive force between the positive nucleus on the outer shell electrons.
Metallic bonding	The strong electrostatic attraction between the regularly arranged metal cations and the delocalised valence electrons between them.
Delocalised electrons	Electrons shared between more than two electrons.
Giant metallic lattice	A three dimensional structure of positive ions and delocalised electrons, bonded together by strong metallic bonds.
Giant covalent lattice	A three dimensional structure of atoms, bonded together by strong covalent bonds.

- History

- Then

- Mendeleev arranged the elements in order of atomic mass
 - Swapped elements to line up to arrange them into groups of similar properties
 - Gaps left where he thought elements would be found
 - Predicted properties for missing elements
 - Newly discovered elements filled in the gap and matched the predicted properties

- Now

- Arranged in increasing atomic number
 - In vertical columns (groups) with same number of outer electrons + similar properties and horizontal rows (periods) giving number of highest energy electron shell

- Arrangement

- In the order of increasing atomic number
 - Periodicity: in periods showing repeating trends in physical and chemical properties e.g. metals → non-metals
 - In groups with similar properties

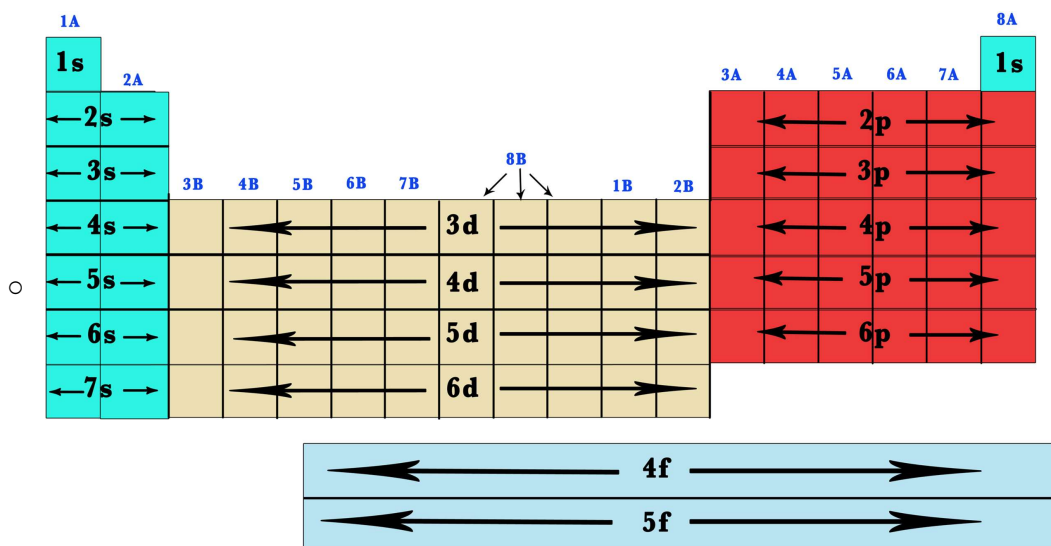
- Electron configuration pattern

- Across period

- Each period starts with an electron in a new highest energy shell
 - Period 2: 2s fills → 2p fills
 - Period 3: 3s fills → 3p fills
 - Period 4: only 4s and 4d occupied in $n = 4$ shell

- Blocks

- * s/p/d/f-block meaning: the **highest energy / outer electron** is in a s/p/d/f-orbital
 - S, p, d and f block



- Name of groups

Group number	Name
1	Alkali metal
2	Alkaline earth metals
3-12	Transition elements
15	Pnictogens
16	Chalcogens
17	Halogens
18	Noble gases

- First ionisation energy

- Energy required to remove one electron from each **atom** in one mole of **gaseous** atoms of an element, forming one mole of **gaseous** 1+ ions
- Unit = kJ mol^{-1}
- Equation: e.g. $\text{Na(g)} \rightarrow \text{Na}^+(\text{g}) + \text{e}^-$ (first ionisation energy = $+496 \text{ kJ mol}^{-1}$)

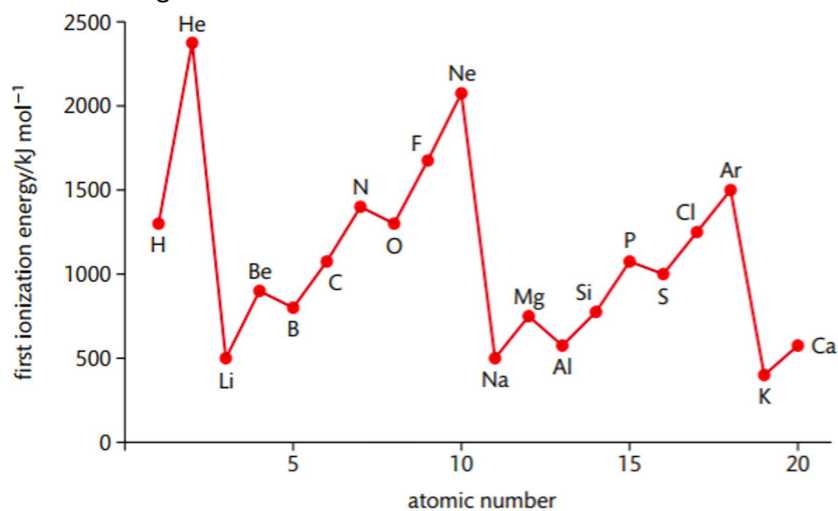
- Factors affecting ionisation energy

- Atomic radius
 - Greater distance between nucleus and outer electrons = less nuclear attraction
 - Large effect on ionisation energy as force of attraction falls sharply with increasing distance
 - * Effect of increase atomic radius outweighs the effect of increasing nuclear charge
- Nuclear charge
 - More protons in nucleus (greater nuclear charge) = greater attraction between the nucleus and the outer electrons = increase in ionisation energy
- Electron shielding
 - Shielding effect: electrons are negatively charged so inner shell electrons repel outer-shell electrons
 - Reduces the attraction between nucleus and outer electrons \rightarrow reduce ionisation energy
 - * Effect of shielding outweighs the effect of increasing nuclear charge

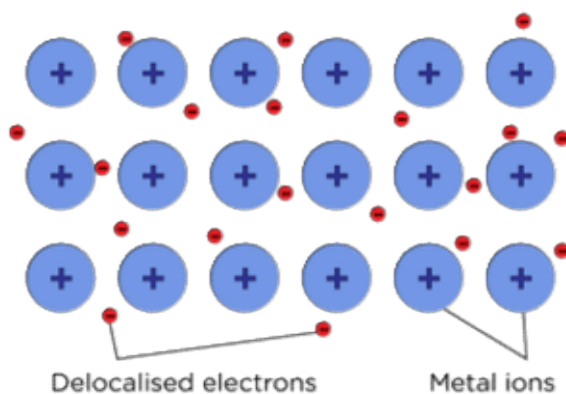
- First ionisation energy trends across a period

- Increases across a period
 - Nuclear charge increases
 - Same number of shells so similar shielding
 - Atomic radius decreases
 - Nuclear attraction increases \rightarrow first ionisation energy increases
- Falls when the p sub-shell is starting to be filled (e.g. $\text{Li} \rightarrow \text{Be}$)
 - 2p / 3p sub-shell has a higher energy than 2s / 3s sub-shell so the electron is easier to remove
 - * Still larger than IE before the decrease

- Falls when pairing of electrons in p sub-shell starts (e.g. N → O)
 - Paired electrons in one of the p orbitals repel one another so it is easier to remove an electron from the atom
- * Still larger than IE before the decrease



- First ionisation energy trend down a group
 - Decrease down a group
 - Atomic radius increases
 - More inner shells so shielding increases
 - * Increase in atomic radius and shielding outweighs the increasing nuclear charge
 - Nuclear attraction on outer electrons decreases → first ionisation energy decreases
- Successive ionisation energy pattern
 - Equation: e.g. $\text{Mg}^+(g) \rightarrow \text{Mg}^{2+}(g) + e^-$
 - Larger than the previous one
 - After the first electron is lost the remaining electrons are pulled closer to the nucleus
 - Nuclear attraction on the remaining electrons increases so more energy needed
 - Large increase when shell change
 - Shell closer to the nucleus → stronger nuclear attraction and less shielding so more energy needed
 - Going to a more inner shell = extremely large increase
 - Smaller but still large increase when going to a new sub-shell / sub-shell become half filled
 - Can be used to work out the number of electrons in each shell + group number of the element
- Metallic bonding structure
 - Regularly arranged metal cations sitting in a sea of delocalised electrons
 - Each atom donate its negative outer shell electrons to a shared pool of electrons which are delocalised throughout the whole structure
 - Cations left behind = nucleus + inner shell electrons
 - Cations are fixed in position
 - Delocalised electrons are mobile and free to move throughout the structure
 - Forms a giant metallic lattice



- Properties of metals

- All conduct electricity
 - Delocalised electrons can move through the structure and carry charge through the structure when a voltage is applied across a metal
 - More delocalised electrons → more electrons can move → better conductivity
 - Conducts electricity both in solid state and when molten
- Most have high melting and boiling points
 - Depends on the strength of metallic bonds
 - Greater cation charge = stronger attractive forces as more electrons are delocalised and forces between electrons + cations are stronger
 - Larger ions = weaker attractive forces due to larger atomic radius decreasing the charge density
 - High temperature needed to provide the large amount of energy to overcome strong electrostatic attraction between the cations and the electrons
 - Melting and boiling points decrease down the group
- Dissolve in liquid metals only
 - Similar force between particles
 - Any interaction between polar / non-polar solvent + solute lead to a reaction rather than dissolving
 - Forces between particles are too large so it is not energetically favourable for them to mix
- Giant covalent structures
 - **Boron, carbon allotropes, and silicon (Si, SiO₂, SiC)**
 - A network of atoms bonded by strong covalent bonds to form a giant covalent lattice
- Diamond / silicon
 - 4 outer shell electrons of each atom form 4 covalent bonds with other carbon / silicon atoms
 - Tetrahedral structure
 - 109.5° bond angle due to electron-pair repulsion
 - High melting and boiling points
 - Atoms held together by strong covalent bonds
 - High temperatures are needed to provide the large quantity of energy needed to break the strong covalent bonds
 - Non-conductors of electricity
 - All 4 outer-shell electrons involved in covalent bond so no charged particles or mobile ions are available for conducting electricity
- Graphite
 - Flat 2D sheets of hexagonally arranged carbon atoms (trigonal planar 120°)
 - Layers bonded by weak London forces
 - Very hard as there are no points of weakness in the structure
 - High melting and boiling points
 - Atoms held together by strong covalent bonds
 - High temperatures are needed to provide the large quantity of energy needed to break the strong covalent bonds
 - Can conduct electricity
 - One electron from each carbon atom is delocalised and is available for conductivity
- Graphene
 - Single layer of graphite
 - Hexagonally arranged (trigonal planar 120°) carbons
 - One of the thinnest + strongest material in existence (atoms held together by strong covalent bond)
 - High melting and boiling points
 - Atoms held together by strong covalent bonds
 - High temperatures are needed to provide the large quantity of energy needed to break the strong covalent bonds
 - Can conduct electricity
 - One electron from each carbon atom is delocalised
 - They can move and conduct electricity
- Applications of graphene

- Electronics
 - Flexible displays
 - Wearables
 - Other next-generation electronic devices
- Atomic radii trend across a period
 - Atomic radii decreases across the period
 - Positive charge in nucleus and negative charge in the valence shell both increases
 - Shielding remains similar as the number of shells doesn't change
 - The attraction between the nucleus and the valence electrons increases
- Melting / boiling point trend across a period
 - Increases from Group 1 to 14
 - Sharp decrease between Group 14 to 15 - change from giant to simple molecular structures
 - Comparatively low from Group 15 to 18
 - The exact boiling points depend on the type of covalent bonding
 - Giant covalent bonding = very high melting and boiling points
 - Simple covalent bonding = depends on strength of intermolecular forces (London forces) which depends on the mass of the nucleus
 - Smaller molecular radius = lower boiling points (hence Group 18 has the lowest boiling points)

3.1.2 Group 2

- Definitions

Term	Definition
• Oxidising agent	A reagent that oxidises another species
Reducing agent	A reagent that reduces another species

- Comparison to Group 1 elements
 - Better electrical conductivity
 - More outer shell electrons → 2 electrons delocalised from each metal atom instead of 1
 - More delocalised electrons to move and carry electric current
 - Higher melting and boiling points
 - **More outer shell electrons** → 2 electrons delocalised from each metal atom instead of 1
 - Higher charge on metal cation
 - Stronger attraction between the metal ions and the delocalised electrons
- Redox reaction of Group 2 elements
 - Two outer shell in the outer s sub-shell
 - Lose two electrons to form 2+ ion with an electron configuration of a noble gas
 - Another species gain these two electrons and is reduced
 - Reducing agent as it reduce another species
- Redox with oxygen
 - $M(s) + O_2(g) \rightarrow 2MO(s)$
 - Each metal atom's oxidation number increases from 0 on LHS to +2 in MO on RHS
 - Each O decreases from 0 in $O_2(g)$ on LHS to -2 in MO(s) on RHS
- Redox with water
 - $M(s) + 2H_2O(l) \rightarrow M(OH)_2(aq) + H_2(g)$
 - Each metal atom increases from 0 in M(s) on LHS to +2 in $M(OH)_2(aq)$ on RHS
 - 2 hydrogen atoms in H_2O decreases from +1 in $H_2O(l)$ on LHS to 0 in $H_2(g)$ on RHS
 - The other 2 H atoms do not change their oxidation state and forms $M(OH)_2(aq)$
 - Observations
 - More vigorous fizzing / bubbling down the group
 - Metal being more soluble down the group / dissolve faster down the group
 - Solution has a higher pH / more alkaline down the group
- Redox with dilute acids
 - Metal + acid → salt + hydrogen
 - Oxidation number change
 - Each metal atom increases from 0 in metal on LHS to +2 in salt on RHS

- Each H decreases from +1 in acid on LHS to 0 in H₂ on RHS
- * **Barium / calcium / strontium sulfate** are **insoluble in water** so reaction stops quickly after the metal solid is coated with insoluble sulfate
- Reactivity trend
 - Reactivity increases down the group
 - First and second ionisation energy decreases down the group / easier to remove outer shell electrons
 - Increased atomic radius
 - Increased shielding
 - The effect of increasing nuclear charge is outweighed by the effects of increasing atomic radius and shielding
 - Attraction between the nucleus and outer electrons decreases
- Reaction of Group 2 oxides with water
 - $\text{MO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{M}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$
 - OH⁻ ions are released and form alkaline solutions of the metal hydroxide
 - Hydroxides only slightly soluble in water
 - When the solution become saturated, addition of further metal oxide causes some ions to come out of the solution and form a solid precipitate: $\text{M}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{M(OH)}_{2(\text{s})}$
 - Solubility of hydroxides increases down the group so the solution can contain more OH⁻ ions and become more alkaline
 - $\text{Mg(OH)}_{2(\text{s})}$ slightly soluble → low OH⁻ concentration, pH ≈ 10
 - $\text{Ba(OH)}_{2(\text{s})}$ more soluble → higher OH⁻ concentration, pH ≈ 13
- Uses of Group 2 compounds as bases
 - Ca(OH)_2 in agriculture to neutralise acidic soil
 - Added to fields as lime
 - $\text{Ca(OH)}_{2(\text{s})} + 2\text{H}^{+}(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{H}_2\text{O(l)}$
 - Sodium or potassium hydroxides are not suitable as they would kill soil organisms and plants (too alkaline)
 - Mg(OH)_2 , MgCO_3 and CaCO_3 are used as antacids in treating indigestion
 - They are weak bases that neutralise excess stomach acids (HCl) that causes the indigestion
 - e.g. milk of magnesia = suspension of Mg(OH)_2 in water
 - $\text{Mg(OH)}_{2(\text{s})} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_{2(\text{aq})} + 2\text{H}_2\text{O(l)}$
 - $\text{CaCO}_{3(\text{s})} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_{2(\text{aq})} + \text{H}_2\text{O(l)} + \text{CO}_{2(\text{g})}$
 - KOH and NaOH are not suitable as ingesting them would cause poisoning and potentially death

3.1.3 The halogens

- Trend in boiling points
 - Boiling point increases down the group
 - Halogens exist as diatomic molecules at RTP
 - Number of electron shells in the atom increases going down the group
 - The atom gets bigger + heavier
 - Number of electrons in the diatomic molecules increases
 - Stronger induced dipole-dipole interactions
 - More energy required to break the intermolecular forces

F₂	Gas	Boiling point increases down the group
Cl₂	Gas	
Br₂	Liquid	
I₂	Solid	
At₂	Solid	

- Appearance under different states

Halogen	Colour in natural state	Colour in aqueous solutions	Colour in organic solvents
F ₂	Pale yellow gas	/	/

• Cl ₂	Pale green gas	Pale green	Pale green
Br ₂	Red-brown liquid	Orange	Orange
I ₂	Shiny grey-black solid	Brown	Violet

- Redox reaction of halogens
 - Halogens all have s²p⁵ electron configuration in their outer shell
 - They gain 1 electron to form 1- ion during redox reactions and gain the electron configuration of the nearest noble gas (reduced)
 - Another species loses electrons to halogen atoms so it is oxidised
 - Halogens are oxidising agents as they oxidise other species
 - Halide ions combine with metal ions to form white (ionic) solids which are **mostly soluble**
- Halogen-halide displacement reactions
 - Solution of hydrogen added to other halide solutions
 - If the halogen added is more reactive than the halide in the solution
 - It will displace the halide in the solution
 - The solution changes colour (from ... to ...)
 - Results
 - Cl₂ reacts with Br⁻ (Cl₂(aq) + 2Br⁻(aq) → 2Cl⁻(aq) + Br₂(aq) orange)
 - Cl₂ reacts with I⁻ (Cl₂(aq) + 2I⁻(aq) → 2Cl⁻(aq) + I₂(aq) violet)
 - Br₂ reacts with I⁻ only (Br₂(aq) + 2I⁻(aq) → 2Br⁻(aq) + I₂(aq) violet)
 - I₂ doesn't react at all
 - Element displaced can be identified by adding cyclohexane (non-polar solvent, dissolve more readily) + mix
 - Use the colour of the top layer to identify the element being displaced
 - Show reactivity: Cl₂ > Br₂ > I₂
- Trend in reactivity
 - Reactivity decreases down the group
 - Atomic radius increases
 - More inner shells so shielding increases
 - Effect of increasing nuclear charge is outweighed by increasing atomic radius and shielding
 - Less nuclear attraction between the nucleus and the outer shell
 - Harder for elements to capture an electron from another species and form 1- ions
- Disproportionation
 - A redox reaction in which the same element is both oxidised and reduced
- Chlorinating water
 - Chlorine reacted with water

$$\text{Cl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HClO}(\text{aq}) + \text{HCl}(\text{aq})$$
 - | | | | |
|---|---|----|-----------|
| 0 | → | -1 | reduction |
| 0 | → | +1 | oxidation |
 - Used in water treatment systems to kill harmful bacteria
 - Bacteria killed by chloric(I) acid (HClO) and chlorate(I) ions (ClO⁻)
 - Chloric(I) acid also acts as a weak bleach e.g. indicator paper will turn red then white
- Benefits and risk of chlorinating water
 - Benefits
 - Kill bacteria in water treatment
 - Reduces the risk of waterborne diseases
 - Risks
 - Hazards of toxic chlorine gas → respiratory irritant in small concentrations, can be fatal if in large concentrations
 - Formation of chlorinated hydrocarbons may cause cancer
- Manufacturing bleach
 - React chlorine with **cold and dilute aqueous NaOH solution**

$$\text{Cl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{NaClO}(\text{aq}) + \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$$
 - | | | | |
|---|---|----|-----------|
| 0 | → | -1 | reduction |
| 0 | → | +1 | oxidation |

- NaClO solution = bleach
- Halide test
 - See 3.1.4

3.1.4 Qualitative analysis

- Test for cations
 - Test for ammonium ion (NH_4^+)
 - $\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - **Aqueous sodium hydroxide** added to a **solution** of ammonium ion
 - Ammonia gas is produced (unlikely to see bubbles as it is very soluble in water)
 - Ammonia can be smelled / turns damp red litmus paper blue
- Test for anions (must be in this order)
 1. Carbonate test
 - Add dilute **nitric acid (not HCl or H_2SO_4 if doing sulfate or halide test later)** to the **solid or solution** being tested
 - If there are bubbles then the compound could be a carbonate
 - $2\text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - Check the gas produced by bubbling it through lime water (calcium hydroxide)
 - Lime water will turn cloudy if the gas is carbon dioxide as a white precipitate of calcium carbonate forms
 - $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
 2. Sulfate test
 - Add barium chloride (barium nitrate if halide test needed later) to the **solution** of the compound
 - If white precipitate forms then SO_4^{2-} is present
 - $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
 3. Halide test
 - Add aqueous silver nitrate (AgNO_3) to an **aqueous solution** of a halide
 - Silver halide precipitates as different colour
 - $\text{Ag}^+(\text{aq}) + \text{X}^-(\text{aq}) \rightarrow \text{AgX}(\text{s})$
 - AgCl = white
 - AgBr = cream
 - AgI = yellow
 - Add aqueous ammonia to test the solubility of the precipitate (colours can be difficult to tell apart)
 - AgCl = soluble in dilute and concentrated $\text{NH}_3(\text{aq})$ only
 - AgBr = soluble in concentrated $\text{NH}_3(\text{aq})$ only, insoluble in dilute $\text{NH}_3(\text{aq})$
 - AgI = insoluble in dilute and concentrated $\text{NH}_3(\text{aq})$
- Correct sequence needed
 - Neither sulfate / halide produce bubbles with dilute acid → cannot have incorrect conclusion
 - BaCO_3 is white and insoluble in water so it gives a false positive result on carbonates as well in sulfate tests
 - Ag_2CO_3 and Ag_2SO_4 both insoluble in water and form precipitates in the test so it gives a false positive result for carbonates and sulfates in halide tests