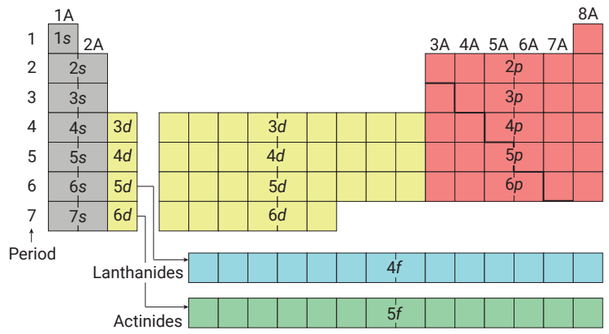
1.6 The Periodic Table

Structure

Elements are arranged according to atomic number. The vertical columns are called **groups**. All the elements in the eight main groups contain the same outer electron configuration.

The horizontal rows are called **periods**. All the elements in a period have the same number of quantum shells containing electrons. The table is also divided into **blocks**. The name of the block shows the orbital in which the elements' outer electrons lie.



Trends

Ionisation energy generally increases across a period because there is an increase in nuclear charge in the same energy level.

There is a decrease between Group 2 and 3 because Group 3's outer electron is partly shielded by the s electrons and between Groups 5 and 6 because of electron-electron repulsion between

the electron pair in one p orbital.

Ionisation energy decreases down a group because the outer electron has increased shielding from inner electrons.

Electronegativity increases across a period because there is an increase in nuclear charge, but the bonding electrons are always shielded by the same inner electrons.

Melting and boiling temperatures generally increase from the first to the fourth element, followed by a large decrease to the fifth element and a small general decrease to the eighth element. This is because the bonding changes from metallic to giant covalent to simple molecular covalent.

Reduction and oxidation (Redox)

Oxidation is loss of electrons and reduction is gain of electrons

An **oxidising agent** is a species that accepts electrons; it becomes reduced itself in the process.

A **reducing agent** is a species that donates electrons; it becomes oxidised itself in the process.

Another way to tell if a reaction is a redox reaction is to work out the **oxidation numbers** of the atoms or ions.

If the oxidation number increases, the species is oxidised; if it

decreases, the species is reduced.

Chemistry of Group 1 and Group 2 metals.

Reaction with water

Group 1 metals react vigorously with cold water to form the hydroxide and hydrogen.

e.g. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

The reaction increases in vigour as you go down the group.

	Float	Fizz	Move	Melt	Burn
Li	✓	✓	✓		
Na	✓	✓	✓	✓	
K	✓	✓	✓	✓	✓

Group 2 metals react less vigorously, in fact magnesium reacts very slowly, while calcium produces a steady stream of bubbles and a white precipitate of calcium hydroxide. Again, the hydroxide and hydrogen are formed, and the reaction increases as you go down the group.

e.g. $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$

Reactivity increases as you go down a group because when the s-block metals react they lose electrons to form positive ions. Since ionisation energies decrease down a group, the energy needed to form positive ions decreases.

Group 1 metals are more reactive than Group 2 metals because they lose only one electron while Group 2 metals lose two.

Reaction with oxygen

All Group 1 and Group 2 metals burn to form solid white oxides.

e.g. $4Li + O_2 \longrightarrow 2Li_2O$

$2Mg + O_2 \longrightarrow 2MgO$

Oxides and hydroxides

All s-block metal oxides are strong bases and neutralise acids to form salt and water.

e.g. $MgO + 2HCl \longrightarrow MgCl_2 + H_2O$

Group 1 oxides and barium oxide react with water to form a soluble hydroxide or alkali.

e.g. $Na_2O + H_2O \longrightarrow 2NaOH$

Since the hydroxides are soluble, they are alkalis.

Test for cations

Most s-block elements may be identified by a flame test (Mg^{2+} ions give no colour). The characteristic colours for the metal ions are:

Ion	Colour	Ion	Colour
Li^+	red	Ca^{2+}	brick red
Na^+	orange yellow	Sr^{2+}	crimson
K^+	lilac	Ba^{2+}	apple green

Solubility in water

All **Group 1** compounds are **soluble**. Many Group 2 compounds are not. Here are some trends for Group 2 compounds.

- All **nitrates** are **soluble**.
- All **carbonates** are **insoluble**.
- **Hydroxides** are **more soluble** as you go down the group.

Magnesium hydroxide is insoluble; barium hydroxide is soluble.

$$Mg^{2+}(aq) + 2OH^-(aq) \longrightarrow Mg(OH)_2(s)$$

Sulfates become **less soluble** as you go down the group.

So, magnesium sulfate is soluble and barium sulfate is insoluble.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Thermal stability of hydroxides and carbonates

Group 2 hydroxides decompose on heating to the oxide and steam

e.g. $Ca(OH)_2(s) \longrightarrow CaO(s) + H_2O(g)$

All Group 2 carbonates decompose on heating to the oxide and

carbon dioxide

e.g. $MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$

Thermal stability increases in both as you go down the group.

The trend for carbonates can be shown in the laboratory

by heating them and seeing how long it takes the CO_2 formed to turn limewater cloudy.

Chemistry of Group 7 halogens

At room temperature, chlorine is a green gas, bromine a red-brown liquid and iodine a grey solid.

As the number of electrons increases with atomic number, there is an increase in the induced dipole – induced dipole intermolecular forces holding the diatomic molecule together. Thus, the melting and boiling temperatures increase as you go down the group.

Trend in reactivity

The halogens react by gaining electrons to form negative halide ions. Therefore, during reactions, halogens are reduced, and they oxidise the other substance. As you go down the group the outer electrons are shielded more and are further from the nucleus. So, it gets harder to attract electrons, and both reactivity and oxidising power decrease down the group.

The halogens react directly with most metals to form the halide.

e.g. $2Na + Cl_2 \longrightarrow 2NaCl$

Displacement reactions

A halogen in a higher position in the group will oxidise a halide ion from lower in the group. Therefore, a more reactive halogen displaces a less reactive halide from its solution.

When displacement reactions happen there are colour changes.

e.g. $Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$

colourless orange

Test for halide ions

Add a few drops of nitric acid to the aqueous halide ion. Then add silver nitrate solution, followed by dilute aqueous ammonia.

Ion	Addition of $AgNO_3(aq)$	Addition of diluted $NH_3(aq)$
Cl^-	White precipitate	Ppt dissolves
Br^-	Cream precipitate	Ppt dissolves slightly
I^-	Pale yellow precipitate	No change

Chlorine and fluoride in water treatment

Chlorine gas is added to drinking water to kill dangerous bacteria and viruses such as cholera and typhoid, thereby preventing the outbreak of serious diseases. It needs to be added in low concentrations (below 1 part per million) in order to be safe.

Fluoride is added to water to reduce tooth decay, by preventing cavities. It is also said to strengthen bones, which helps prevent osteoporosis. Again, it appears to only have beneficial effects when added below 1 ppm.

Salt formation

Insoluble salts are formed by a precipitation reaction. In this reaction two suitable solutions are mixed to form a soluble salt and an insoluble salt. For example, the formation of calcium carbonate.

$$Ca(NO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaNO_3(aq)$$

The precipitate is filtered, washed and dried.

Soluble salts are usually formed by neutralising an acid. For example, the formation of magnesium sulfate.

$$MgO(s) + H_2SO4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

Any excess solid is filtered. The solution is evaporated and left to cool to form crystals.