

WJEC Chemistry A-level

1.6: The Periodic Table

Detailed Notes

Welsh Specification

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The Periodic Table

The Periodic Table arranges the known elements according to **atomic number**. This increases from left to right along a period and down a group. All the elements along a period have the **same number of electron shells**. All the elements down a group have the same number of outer electrons, this number is indicated by the **group number**.

These electron structures of the elements allows them to be **classified into blocks** within the Periodic Table that indicate which energy level the **outer electron** occupies:

- s-block = groups 1 and 2**
- p-block = groups 3 to 0**
- d-block = transition metals**

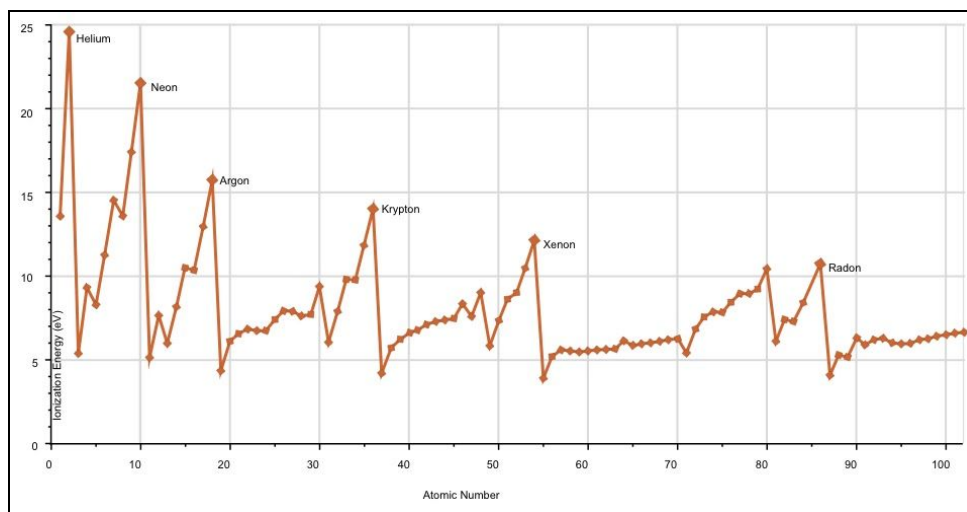
These different electron configurations are often linked to other trends within the Periodic Table. **Periodicity** is the study of these trends.

Trends in the Periodic Table

Ionisation Energy

Along a period, ionisation energy **increases**. The decreasing **atomic radius** and increasing **nuclear charge** means that the outer electrons are **held more strongly** and therefore **more energy is required** to remove the outer electron and ionise the atom.

Down a group, ionisation energy **decreases**. The **nuclear attraction** between the nucleus and outer electrons reduces and increasing amounts of **shielding** means less energy is required to remove the outer electron.



https://commons.wikimedia.org/wiki/File:Ionization_energies.svg

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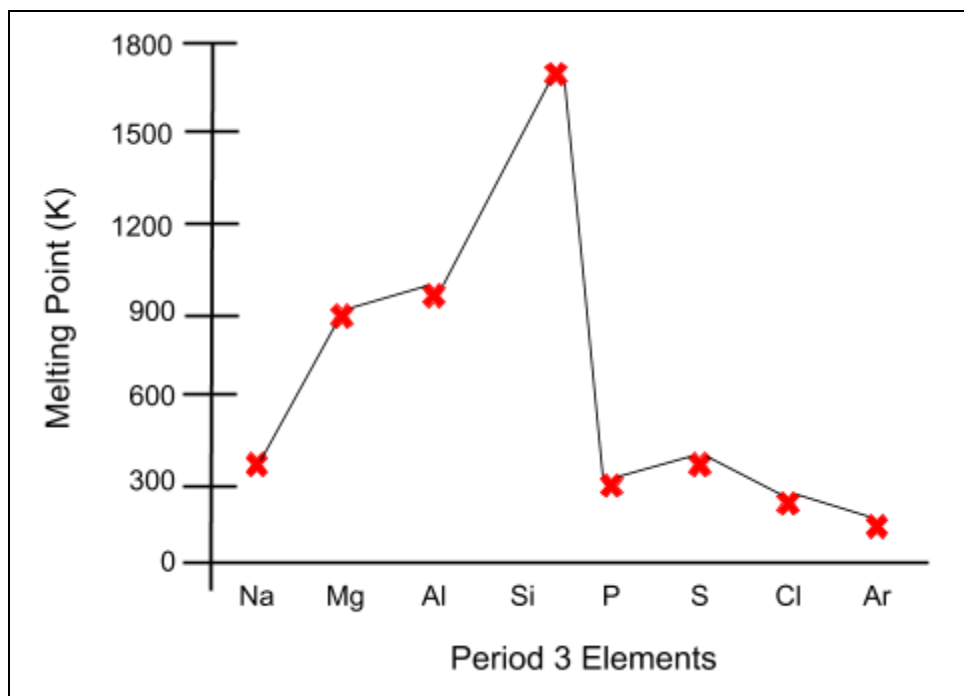


Melting & Boiling Points

Along a Period

The melting/boiling points of elements in the Periodic Table vary depending on the **type of structure** and **bond strength** that the element has. A good example of how this changes along a period can be seen with the period three elements.

Example:



Sodium, magnesium and aluminium are all **metals** with metallic bonding. Their melting points increase due to **increasingly positively charged ions** (Na = +1, Mg = +2, Al = +3). This also means **more electrons** are released as delocalised electrons so the **attractive electrostatic forces** increase from Na to Al.

Silicon is **macromolecular** meaning it has a **very strong covalent structure**. These covalent bonds require a lot of energy to break, giving it a very high melting point.

Phosphorus, sulphur and chlorine are all **simple covalent** molecules, with weak **Van der Waals** forces. These intermolecular forces don't require much energy to overcome so these molecules have relatively low melting points. Their relative melting points decrease from silicon (S_8) to phosphorus (P_4) to chlorine (Cl_2) and finally to argon (Ar). This is due to the **decreasing size** of the molecules.





Argon is a **noble gas** that exists as **individual atoms** with a **full outer shell** of electrons. This makes the atom very **stable** and the van der Waals forces between them very **weak**. As a result, the melting point of Argon is very low and it exists as a gas at room temperature.

Down a Group

The trend in melting/boiling points down a group depends on the changes in **bond strength** or the intermolecular force strength between the atoms. It can be demonstrated with the elements of group 7, the halogens.

The group 7 elements are all **simple covalent molecules** held together with **Van der Waals** forces. The strength of these forces **increases as the Ar** of the atom increases. Therefore, the strength of the Van der Waals forces increases down the group so more energy is required to overcome them. This results in a higher boiling point.

Example:

Fluorine is a gas at room temperature, whereas iodine is a solid.

Electronegativity

Every atom has electronegativity, which is defined as:

The power of an atom to attract negative charge towards itself within a covalent bond.

This 'power' is different for every atom depending on its **size and nuclear charge**.

Electronegativity **increases along a period** as atomic radius decreases and the number of protons increases, which increases the nuclear attraction. Electronegativity **decreases down a group** because shielding increases so the nuclear attraction decreases.

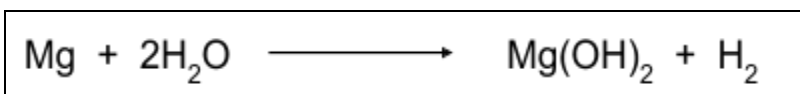
Group 2 Elements

The Group 2 metals lose two electrons to form **2+ ions** when they react, in order to achieve a **full outer shell**. Atomic radius of group 2 metals increases down the group due to **additional electron shells**. Increased electron **shielding** down the group makes the outer electrons easier to lose because there is a weaker nuclear attraction. Therefore, reactivity of the group 2 metals increases down the group.

Reactions with Water

The group 2 metals react with water in a **redox reaction** to produce a **metal hydroxide** and hydrogen.

Example:





The magnesium is oxidised from oxidation state 0 to oxidation state +2.

(Oxidation is loss of e⁻)

The other group 2 elements react in the same way.

Magnesium reacts very slowly in this way with **liquid water**. The reaction can be much faster with **steam**, as it provides the reaction with **extra energy**. When steam is used, the magnesium burns with a **bright white flame** to form hydrogen and magnesium oxide, a white powder.

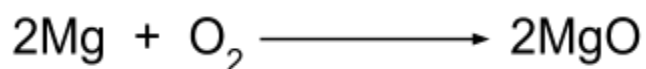
Example:



Reactions with Oxygen

The group 2 elements react with **pure oxygen** to produce simple oxides as they have an oxidation number of **+2** that balances with the -2 oxidation number of oxygen.

Example:



Barium is a slight exception to this as it will form **barium peroxide**, where the oxygen ion has an oxidation number of **-1**.

Example:



Strontium can also react in this way to form **strontium peroxide**, but only when heated under **high pressures**.

Group 2 Hydroxides

Group 2 ions react with OH⁻ ions to form **hydroxides**. When writing the formula of these compounds, the two hydroxide ions required must be bracketed together.

Example:





The **solubility** of group 2 hydroxides varies and means these compounds have different uses. Solubility **increases down the group** meaning magnesium hydroxide, $\text{Mg}(\text{OH})_2$, is the least soluble and barium hydroxide, $\text{Ba}(\text{OH})_2$, the most soluble.

As a result, magnesium hydroxide is used in **medicine as an antacid** as it is alkaline and can neutralise acids. It is used in a similar way in **agriculture, to neutralise acidic soils**.

Group 2 Sulfates

Group 2 element ions react with **SO_4^{2-} ions** to form **hydroxides**. These group 2 sulfates **decrease in solubility** down the group meaning magnesium sulfate, MgSO_4 , is the most soluble and barium sulfate (BaSO_4) is the least soluble.

The insolubility of barium sulfate means it is very useful in medicine as **barium meals**. Barium meals are a form of **medical tracer** that allow internal tissues and organs to be imaged. Barium sulfate is toxic if it enters the bloodstream, however, since it is insoluble, it cannot be absorbed into the blood. This makes it safe to use in the body.

Barium chloride is used as a **test for sulfate ions** as it reacts with sulfate ions to form barium sulfate, which is a white precipitate.

Example:



Group 2 Flame Tests

The group two elements can be identified as they produce different **coloured flames**.

- Lithium (Li^+) – crimson flame
- Sodium (Na^+) – yellow flame
- Potassium (K^+) – lilac flame
- Magnesium shows no colour
- Strontium (Sr^{2+}) – orange flame
- Calcium (Ca^{2+}) – orange-red flame
- Barium (Ba^{2+}) – pale green flame

Thermal stability of Group 2 carbonates and hydroxides

Group 2 carbonates undergo thermal decomposition, and require higher temperatures as you go down the group. This is because **thermal stability** of the group 2 carbonates **increases** down the group. As you go down the group the metal ion increases in size which means the charge has a **less polarising effect** on the carbonate ion, making it harder for the carbonate to be decomposed.





Group 2 hydroxides also form oxides when they undergo thermal decomposition. The trends of the hydroxides are the same as those for the carbonates.



Reactivity trends of Group 1 and Group 2

The reactivity of group 1 and group 2 metals **increases down each group**.

Group 1 metals have one electron in their outer shell so they react in similar ways to form **1+ ions**. They react violently with water, forming alkaline solutions:



Group 2 metals have two electrons in their outer shell, so they react in similar ways to form **2+ ions**. They are **less reactive than group 1** elements as it is harder to lose two electrons.

Reactivity increases down both groups as the **outer electron(s) becomes more easily lost**. This is because shielding and atomic radius increases, leading to a **weaker nuclear attraction**.

Basic character

The **oxides** and hydroxides formed from group 1 and group 2 elements have **basic character**. Group 1 oxides/hydroxides are **stronger bases** than group 2 oxides/hydroxides. This is because the group 2 ions have charges of 2+. This larger charge means that there is a greater force of attraction in the compound and so it is harder for the hydroxide/oxide ions to be released.

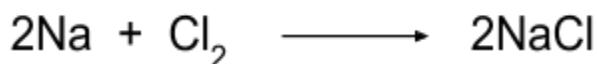
Group 7 Elements

The group 7 elements are **highly reactive non-metals** that need to gain an electron to form a 1- ion and achieve a full outer shell of electrons. Reactivity of the group 7 elements **decreases down the group**. This is because the atomic radius increases due to additional electron shells making the positive attraction of the nucleus is weaker. These electron shells also increase the amount of **shielding**, making it harder to attract electrons in a reaction.

Reactions with Metals

The halogens react with metals to produce **salts** such as sodium chloride.

Example:





The difference in volatility of these reactions can be seen in the reactions with iron wool.

Example:

Halogen	Reaction
Fluorine	Burns with cold iron wool
Chlorine	Burns vigorously with hot iron wool
Bromine	Burns quickly with hot iron wool
Iodine	Reacts slowly with hot iron wool producing iodine vapours

Clearly, from the table, **volatility decreases** down group 7.

Halogen Oxidising Power

The halogens act as good **oxidising agents** as they **accept electrons** from the species being oxidised and are **reduced** as a result. This oxidising power decreases down the group as their ability to attract electrons decreases due to **shielding** and a **greater atomic radius**.

The relative oxidising strengths mean a halogen will **displace any halide beneath it** in the Periodic Table.

Example:

Cl_2 will displace Br^- and I^- ions.
 Br_2 will displace I^- ions
 I_2 won't displace any halide ions.

Reactions with Silver Nitrate

Acidified silver nitrate contains aqueous **Ag^+ ions** and is used to **test for halide ions**. It reacts with these ions to form **different coloured precipitates** depending on the ion present.

The precipitates formed may not be clear to distinguish so they can be tested further using **ammonia**.





Example:

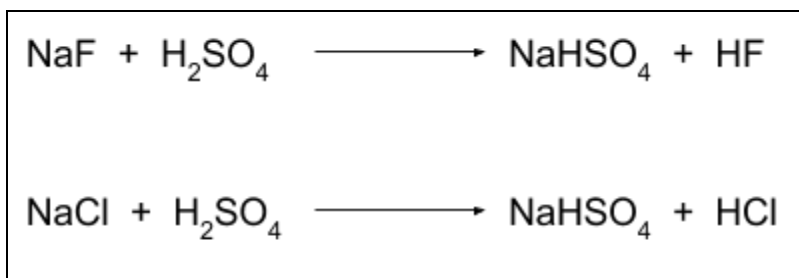
	Cl ⁻	Br ⁻	I ⁻
+ AgNO ₃	White precipitate (AgCl)	Cream precipitate (AgBr)	Yellow Precipitate (AgI)
+ dilute NH ₃	Precipitate dissolves	No Change	No Change
+ conc. NH ₃	Precipitate dissolves	Precipitate dissolves	No Change

Redox Displacement Reactions

Halide ions are good **reducing agents** as they **donate electrons** to the species being reduced and are themselves **oxidised**. This reducing power increases down the group as electrons are easier to lose from larger ions since there is a reduced nuclear attraction due to **shielding** and a **larger atomic radius**.

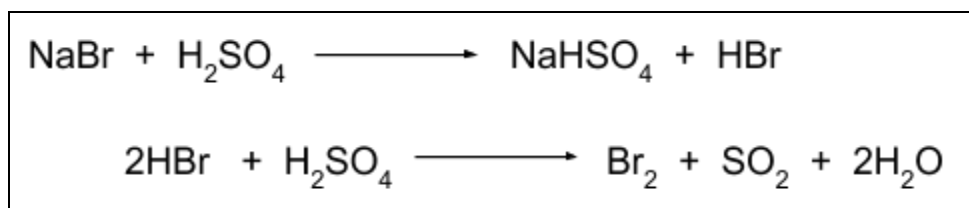
Redox displacement reactions are best demonstrated with sulphuric acid:

1. Fluoride and Chloride ions.



These reactions stop here because HF and HCl are **not strong enough reducing agents** to reduce the sulfuric acid. HF and HCl can both be identified as **misty fumes**.

2. Bromide ions.

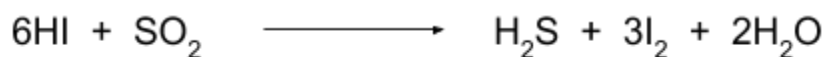
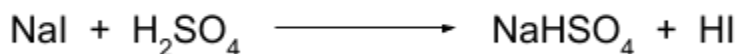


Misty fumes of HBr are produced in the first reaction followed by **choking fumes** of SO₂ and **orange fumes** of Br₂ in the second reaction.





3. Iodide ions.

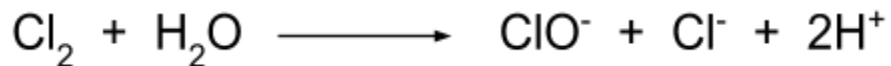


Misty fumes of HI are produced in the first reaction. **Choking fumes** of SO_2 , **toxic gas H_2S** (smells of bad eggs) and **solid iodine** are produced in the further two reactions.

Halogens in Water Treatment

Chlorine is used in small quantities to kill bacteria in **water treatment processes**. When the chlorine reacts with cold water, chlorate(I) ions (ClO^-) are produced. These are very good at **killing bacteria**, helping to treat the water and make it **safe** to drink.

Example:



This is a **disproportionation** reaction as the chlorine is both **oxidised and reduced**. The oxidation state goes from zero to both +1 and -1.

This poses some risks as chlorine can be **toxic** in larger quantities, however the benefits of clean, treated water outweigh the risks.

Fluoride ions are also used in water treatment to prevent teeth decay.

Some people disagree about whether ions should be added to water and believe everyone should have the **right to choose** what goes into the water systems.





Soluble and insoluble salts

Soluble salt

Soluble salts can be formed by reacting **acids with insoluble or soluble bases**. A soluble salt can be separated from a solution by crystallisation or evaporation:

Evaporation:

The solution is slowly heated in an evaporating dish so that the solvent evaporates and the crystals start to form. The solution is continued heating until all that is left are the dry salt crystals.

Crystallisation:

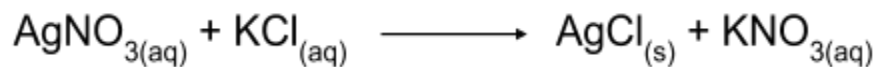
The solution is slowly heated in an evaporating dish so that some of the solvent evaporates. Once the crystals start to form, the heat is removed and the solution is left to cool. Once the crystals have formed, the crystals are filtered out of the solution and left to dry in a warm place.

Insoluble salt

An insoluble salt can be formed by a **precipitation** reaction:

Precipitation:

A precipitation reaction can occur if **two solutions** containing **different salts** are reacted together. A common example is the reaction of aqueous silver nitrate with aqueous potassium chloride to produce a precipitate of silver chloride:



After precipitation, the **insoluble salt** can be **filtered** from the solution and **gravimetric analysis** can be carried out to determine the amount or concentration of ionic compounds in the solutions. For example, the **concentration of an ion in a solution** can be found by precipitating the ion and then weighing the result.

