

OCR A Chemistry A-level

Module 3.1: The Periodic Table

Detailed Notes

This work by [PMT Education](https://www.pmt.education) is licensed under [CC BY-NC-ND 4.0](https://creativecommons.org/licenses/by-nc-nd/4.0/)





3.1.1 Periodicity

The Structure of the Periodic Table

The Periodic Table arranges the known elements in order of **increasing proton number**. All the elements along a **period** (row) have the same number of **electron shells**. All the elements down a **group** (column) have the same number of **outer electrons** which is indicated by the group number. Elements in the same group have similar chemical properties.

Periodic Trend in Electron Configuration and Ionisation Energy

Elements are classified into **blocks** within the Periodic Table, and elements in the same block have their outer electrons in the same type of **orbital**.

s-block = Groups 1 and 2

p-block = Groups 3 to 0

d-block = Transition metals

f-block = Lanthanides & actinides

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

The **energy** of the orbitals **increases from s to f**, meaning the orbitals are **filled in this order**. Each orbital is filled before the next one is used to hold electrons. For Period 2, first the s orbitals are filled, followed by the p orbitals. The same applies to Period 3.

Example:

Nitrogen has the atomic number 7 so has 7 electrons and has the configuration: $1s^2 2s^2 2p^3$

Atomic Radius

Along a **period**, atomic radius **decreases**. This is due to an **increased nuclear charge** for the same number of electron shells. The outer electrons are pulled in closer to the nucleus as the increased charge produces a **greater attraction**. As a result, the atomic radius for that element is reduced.

Down a **group**, atomic radius **increases**. With each increment down a group, an electron shell is added. This increases the distance between the outer electrons and the nucleus, **reducing the power of attraction**. More shells also increases electron **shielding** where the inner shells create a 'barrier' that blocks the attractive forces. Therefore, the **nuclear attraction is reduced** further and atomic radius increases.





Ionisation Energy

First ionisation energy is defined as:

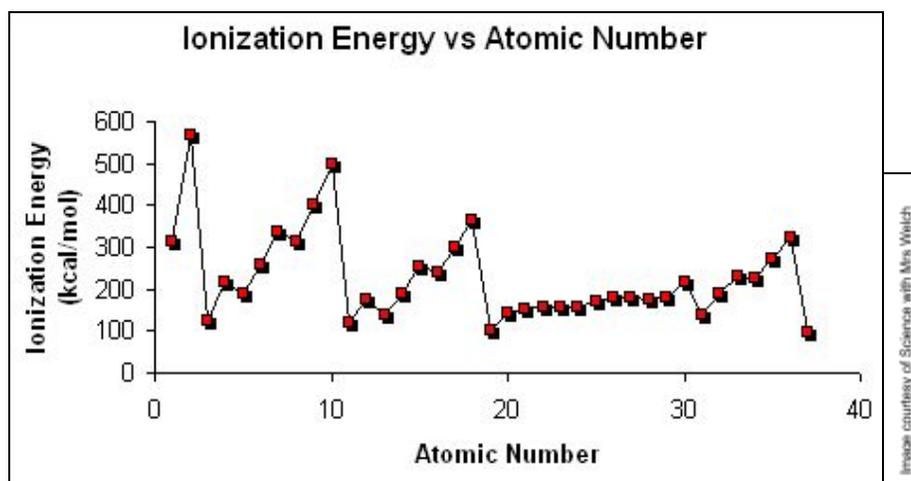
The minimum energy required to remove one mole of electrons from one mole of atoms in a gaseous state. It is measured in kJmol^{-1} .



Successive ionisation energies occur when further electrons are removed. This usually requires **more energy** because as electrons are removed, the **electrostatic force of attraction** between the positive nucleus and the negative outer electron **increases**. More energy is therefore needed to **overcome this attraction**, so ionisation energy increases.

First ionisation energy follows **trends** within the Periodic Table as they are influenced by proton-electron forces of attraction and electron shielding.

- **Along a Period** - first ionisation energy **increases** due to a **decreasing atomic radius** and greater electrostatic **forces of attraction**.
- **Down a Group** - first ionisation energy **decreases** due to an **increasing atomic radius** and **electron shielding** which reduces the effect of the electrostatic forces of attraction.



Successive ionisation energies involve removing one mole of electrons from **one mole of gaseous ions**. For example, the third ionisation energy of potassium would involve removing one electron from $\text{K}^{2+}(\text{g})$ to form $\text{K}^{3+}(\text{g})$.





Successive ionisation energies increase because atomic radius decreases and there is greater attraction between outer shell electrons and the nucleus.

A large **jump** between successive ionisation energies indicates which **group** an element is in. The successive ionisation energies for an element are shown below:

1st	2nd	3rd	4th	5th
801	2427	3660	25026	32827

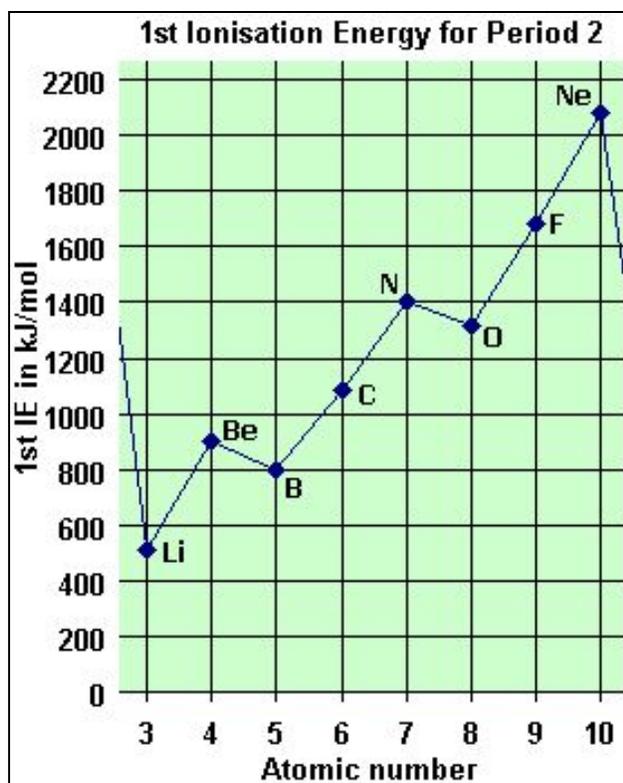
The **large jump between the third and fourth ionisation energy** shows there are 3 electrons that are relatively easy to remove then a fourth one which requires a lot more **energy to remove**. This shows that there are 3 electrons in an outer shell. This means the element is in **Group 3**. There are 5 ionisation energies for this element so it has **5 electrons**. An element with 5 electrons in Group 3 is boron.

Period 2

First ionisation energies follow a general increasing trend along Period 2. This is due to the decreasing atomic radius and increasing nuclear charge meaning outer electrons are **held more strongly**.

Boron and oxygen are **exceptions** to this trend. **Boron** has a **lower first ionisation energy** than expected by the general trend as a result of the energy difference between the 2s and 2p sub-shells. The electron is being removed from a **higher energy level** that is further from the nucleus, so the electron is held less strongly.

Oxygen has a lower first ionisation energy than would be otherwise expected due to **repulsion** within the 2p orbital when two electrons with opposite spins are placed in the same orbital. This repulsion is **destabilising** in comparison to the configuration of nitrogen and allows the electron to be removed from oxygen more easily.





Period 3

First ionisation energies follow a **general increasing trend** along Period 3. This is due to the decreasing atomic radius and increasing nuclear charge meaning outer electrons are **held more strongly**.

Aluminium and sulfur are **exceptions** to this trend. **Aluminium** has a **lower first ionisation energy** than expected by the general trend as a result of the energy difference between the 3s and 3p sub-shells. The electron is being removed from a **higher energy level** that is further from the nucleus, so it is held less strongly.

Sulfur has a lower first ionisation energy than would be otherwise expected due to **repulsion** within the 3p orbital when two electrons with opposite spins are placed in the same orbital. This repulsion is **destabilising** in comparison to the configuration of phosphorus and allows the electron to be removed from sulfur more easily.

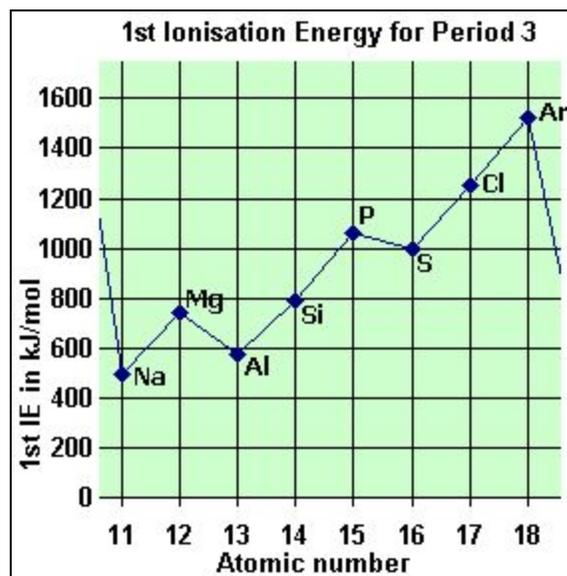


Image courtesy of Doc Brown's Chemistry

Periodic Trend in Structure and Melting Point

Metallic Bonding

Metallic bonding consists of a **lattice of positively charged ions** surrounded by a **'sea' of delocalised electrons**. There are very strong **electrostatic forces of attraction** between the oppositely charged particles.

The **greater the charge** on the positive ion, the **stronger the attractive force** as more electrons are released into the 'sea'. Ions that are **larger in size**, such as barium, produce a **weaker attraction** due to their **greater atomic radius** decreasing the charge density.

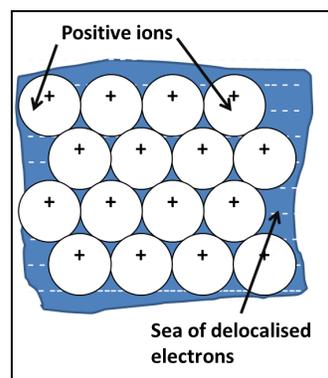


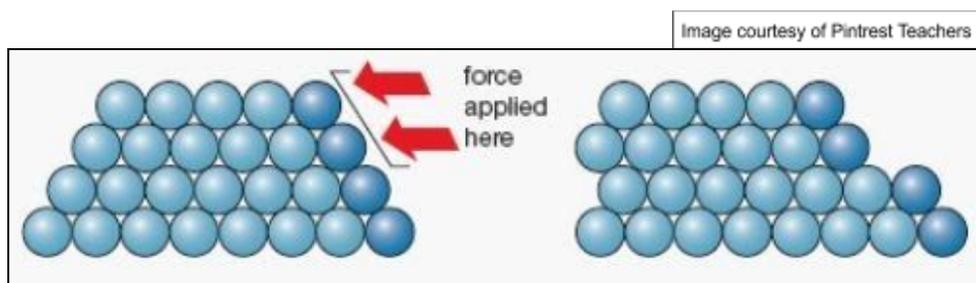
Image courtesy of Science Revision



Metallic Properties

Substances with metallic structure are often **good conductors**. The 'sea' of delocalised electrons is able to move and **carry a flow of charge**.

Metals are also **malleable** as the **uniform layers** of positive ions are able to slide over one another. The delocalised electrons prevent fragmentation as they can move around the lattice.



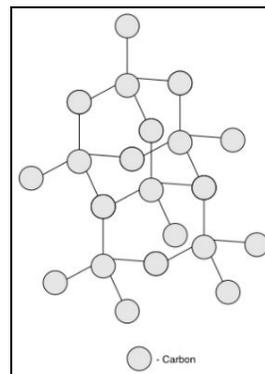
The electrostatic forces of attraction between the positive ions and delocalised electrons are very strong and therefore require a lot of energy to overcome. This means metallic substances have **high melting points** and are nearly always **solid at room temperature**. Mercury is the only metal which is a liquid at room temperature.

Giant Covalent Lattices

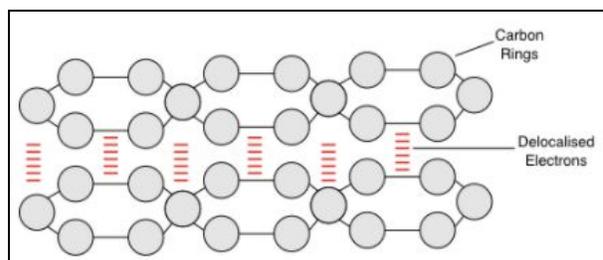
A solid giant covalent lattice is a network of atoms bonded by many **strong covalent bonds**.

Substances that have a macromolecular structure are **covalently bonded** into a **giant lattice** structure. Each atom has **multiple covalent bonds** which are very strong, giving the substance a **very high melting point**.

Diamond is a macromolecular structure made up of carbon atoms, each of which are bonded to four further carbon atoms. This provides a very **rigid** structure, making diamond one of the **hardest materials**.

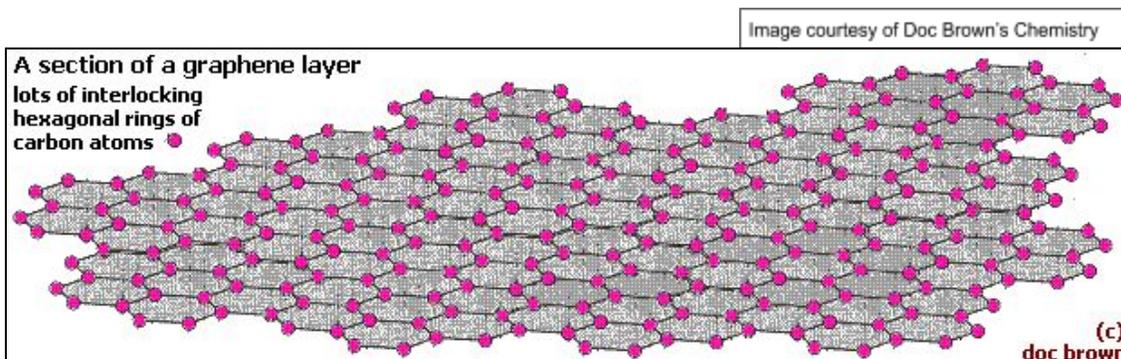


Graphite is another **macromolecular** structure made up of carbon atoms. However, in graphite, each carbon atom is bonded to three other carbon atoms in **flat sheets**. This means there is one **delocalised electron** per carbon atom which can move between layers, meaning graphite can **conduct electricity**. There are strong covalent bonds between carbon atoms but weak van der Waals forces between the layers. These **weak intermolecular forces** allow layers to slide over each other easily, making graphite a good **lubricant**.





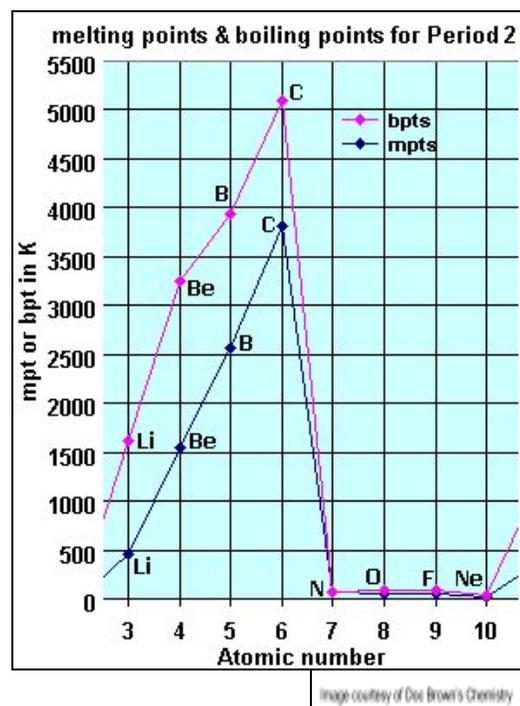
Graphene is a relatively new material consisting of single, **2D sheets of graphite** that are just **one atom thick**. These sheets are formed of **hexagonal carbon rings** that create a very strong, rigid material that is extremely **lightweight**. Delocalised electrons move through each layer allowing it to **conduct** electricity.



Variation in Melting Points across Periods 2 and 3

The melting points of the Period 2 elements peaks towards the middle of the period due to the different **bond strength and structures**:

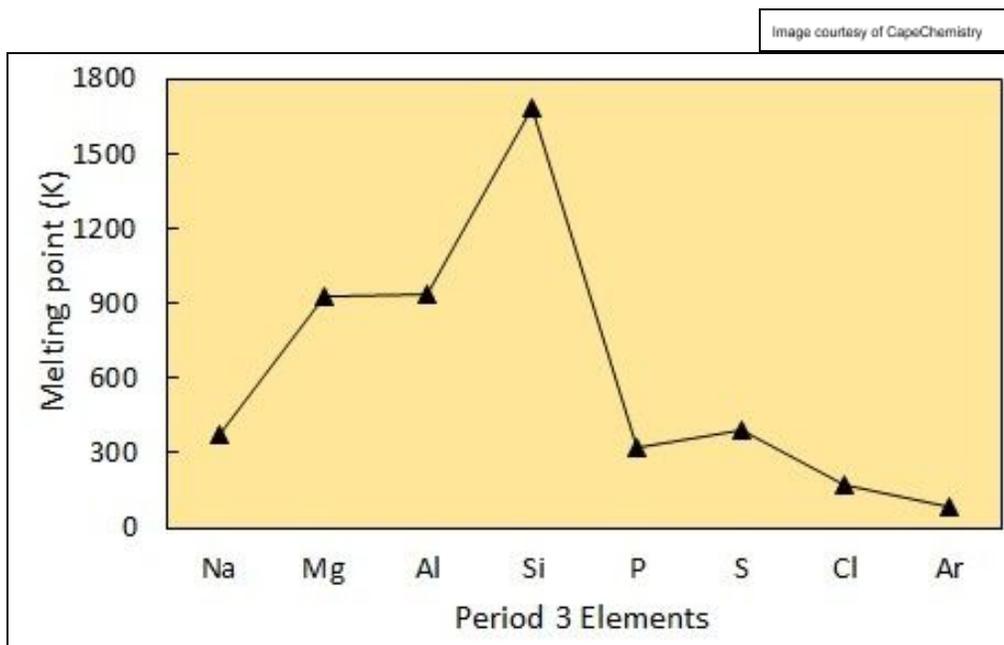
- Lithium and beryllium have **metallic bonding**. Their melting points increase due to a greater **positive charge of the ions** (Li = +1, Be = +2). There are **more electrons released** as free electrons in the beryllium lattice so the attractive electrostatic forces are greater than for lithium.
- Boron** and **carbon** form giant covalent lattices with **very strong covalent bonds**. Covalent bonds require a lot of energy to break, giving them very high melting points.
- Nitrogen, oxygen, fluorine and neon are all **small, simple covalent molecules** with weak **van der Waals** intermolecular forces. These intermolecular forces don't require much energy to overcome so these molecules have relatively low, similar melting points.





The melting points of the Period 3 elements are also linked to the **bond strength and structure**:

- Sodium, magnesium and aluminium are all metals with **metallic bonding**. Their melting points increase due to greater **positive charged ions** (Na = +1, Mg = +2, Al = +3). This also means **more electrons are released** as free 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, similar melting points.
- 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.





3.1.2 Group 2

Redox Reactions and Reactivity of Group 2 Metals

Group 2 metals have an ns^2 valence shell configuration. When Group 2 metals react, they lose two electrons to **form 2+ ions**. This allows them to achieve a full outer shell. All Group 2 elements tarnish in air and form a coating of the metal oxide. The various **physical properties** of these elements follow trends down the group:

Atomic Radius

The atomic radius of Group 2 metals **increases** down the group due to additional electron shells.

Reactivity

Increased electron **shielding** and increased atomic radius down the group makes the outer electrons **easier to lose**. Therefore, reactivity of the Group 2 metals **increases** down the group. The increase in reactivity down the group is demonstrated by the reactions below.

Ionisation Energy

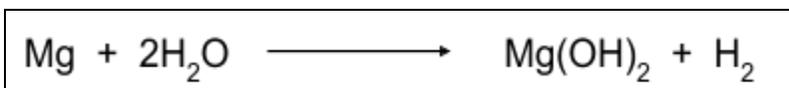
The first and second ionisation energies of Group 2 metals **decrease** down the group due to a greater atomic radius and increased amounts of shielding. This makes it easier for an electron to be removed.

Reactions of Group 2 Compounds

Reactions with Water

The Group 2 metals react with water in a redox reaction to produce a **metal hydroxide and hydrogen**. The metal hydroxide forms as an **alkaline solution**, hence why the Group 2 metals are known as the alkaline earth metals.

Example:



*The Magnesium is oxidised from an oxidation state of 0 to +2.
(Oxidation is loss of e^-)*





Magnesium reacts very slowly in this way with liquid water, however, 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:



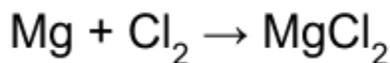
Down the group, the **reactivity** of the Group 2 metals with water **increases**. Calcium, strontium, and barium all react with cold water with **increasingly vigorous** reactions. These reactions all form the metal hydroxide and hydrogen gas.

The **metal oxides** of the Group 2 elements react with water to form basic metal hydroxides. The alkalinity of these increases down the group as the solubility of the metal hydroxides increases.

Reactions with Chlorine

Group 2 metals all react with chlorine gas to form **metal chlorides**, which are all **white precipitates**. As you move down the group the reactions become **more vigorous** because the elements are **more reactive**.

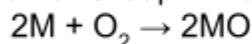
Example:



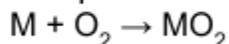
Reactions with Oxygen

Group 2 metals react with oxygen to form **oxides**. Once the reaction has been initiated it is **vigorous**. **Strontium** and **barium** can react with excess oxygen and heat energy to form **metal peroxides**.

The general equation of Group 2 metals with oxygen is:



Strontium or barium may react with oxygen and form the peroxide:





Reactions with Dilute Acids

The Group 2 metals react with dilute acids to produce **bubbles of hydrogen gas** and solutions of **metal compounds**.

Example:

Hydrochloric Acid:



Sulfuric Acid:



Nitric Acid:

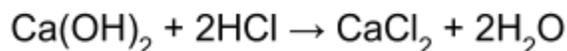


Reactions of Group 2 Hydroxides with Dilute Acid

Group 2 hydroxides react with dilute acid to form **a salt** and **water**. This is a type of **neutralisation** reaction. The salt formed depends on the acid used:

- **Hydrochloric** acid forms **chloride** salts.
- **Sulfuric** acid forms **sulfate** salts.
- **Nitric** acid forms **nitrate** salts.

Example: Calcium hydroxide with hydrochloric acid



Solubility of Group 2 Compounds

Solubility of Group 2 Hydroxides

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

As a result, **magnesium hydroxide** is used in medicine as an **antacid** to treat indigestion as it is alkaline and can neutralise acids. **Calcium carbonate** can also be used as an **antacid**. **Calcium hydroxide** is also used in **agriculture** to neutralise acidic soils.





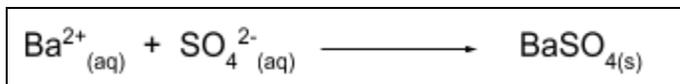
Solubility of Group 2 Sulfates

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**. These 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 because it is insoluble, it **cannot be absorbed** into the blood. Therefore, it is **safe** to use in this way.

Barium chloride is used as a **test for sulfate ions** as it reacts to form barium sulfate which forms as a **white precipitate** when sulfate ions are present.

Example:



3.1.3 The Halogens

Characteristic Physical Properties

The Group 7 elements are **highly reactive non-metals**. The halogens exist as **diatomic** molecules with single covalent bonds. In reactions, in order to achieve a full outer shell, the halogens gain an electron and form a **1- ion**.

Atomic Radius

Atomic radius of Group 7 elements **increases down the group** due to additional electron shells.

Electronegativity

As you go down Group 7, **atomic radius and electron shielding increases**. This means electrons in the outer shells are less strongly attracted to the nucleus, and so are more easily removed. Therefore, electronegativity decreases down Group 7.

Melting and Boiling Point

The Group 7 elements are **simple covalent molecules**, with weak **van der waals** forces. The strength of these intermolecular forces increases as the relative atomic mass of the molecule increases, so the strength of the van der waals forces **increases down the group**. This means **more energy** is required to overcome them, resulting in higher melting and boiling points. This is why **fluorine** is a **gas** at room temperature, whereas **iodine** is a **solid**.





Reactivity

Halogens have s^2p^5 outer shell **electron configuration**. They need to **gain an electron** in order to react. As atomic radius increases, this becomes harder as the positive attraction of the nucleus is weakened by additional **shielding**. Therefore, down Group 7 it is harder to attract an electron so **reactivity decreases**.

Appearance

The colour and state of the halogens at room temperature and pressure are listed below:

Fluorine - pale yellow gas

Chlorine - pale green gas

Bromine - red/brown liquid

Iodine - dark grey solid which sublimes to a purple vapour

Redox Reactions and Reactivity of Halogens and their Compounds

Oxidising Power of the Halogens

The halogens act as good **oxidising agents** as they accept electrons from the species being oxidised and are themselves reduced. 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.

Cl_2 will displace Br^- and I^- ions.

Br_2 will displace I^- ions

I_2 won't displace any halide ions.

Halide Ions

The negative ions of halogens are known as **halide ions**. These 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 due to shielding and a larger atomic radius meaning there is a **weaker attraction** between the nucleus and the outer electrons.





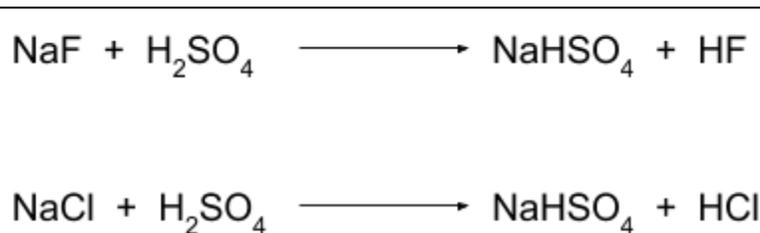
Reactions of Group 7 Elements and Ions

Redox Reactions

The redox reactions between Group 1 and Group 2 halides and H_2SO_4 vary depending on the **reducing ability** of the halide:

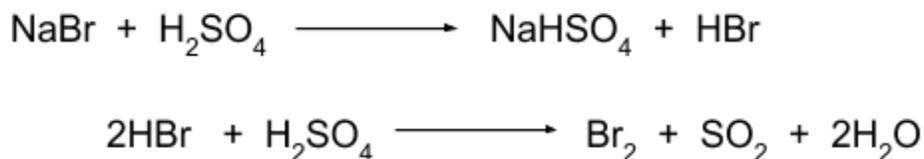
Example:

1. Fluoride and Chloride ions



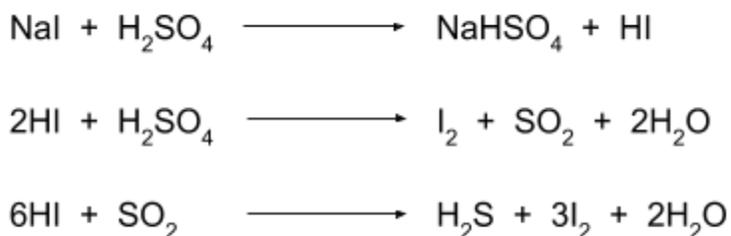
No further reactions take place since HF and HCl are not strong enough reducing agents to reduce H_2SO_4 . HF and HCl will both be observed as **misty fumes**.

2. Bromide ions



HBr will be observed as **misty fumes** from the first reaction. **Orange fumes** of Br_2 and **choking fumes** of SO_2 will be observed in the second reaction, where HBr reduces H_2SO_4 .

3. Iodide ions



HI will be observed as **misty fumes** from the first reaction. HI then reduces H_2SO_4 to **solid iodine** and **choking fumes** of SO_2 . The HI will then further reduce the SO_2 to **toxic gas** H_2S (which smells of bad eggs).



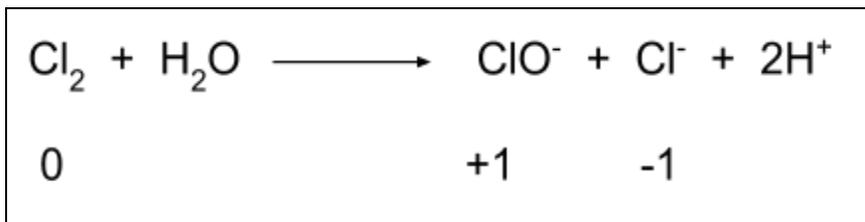


The **greater** the reducing power of the halide, the **further** the reaction will proceed as the halide is powerful enough to reduce more species. These reactions therefore represent the trend in reducing power in halide ions. Clearly, **reducing power increases** down Group 7.

Disproportionation Reactions

A **disproportionation** reaction is a reaction in which a species is both **oxidised** and **reduced**. Chlorine reacts with cold water to produce **chlorate(I) ions (ClO⁻)** and **chloride ions** via a **disproportionation reaction** as the chlorine is both oxidised and reduced. The oxidation state goes from zero to both **+1 and -1**.

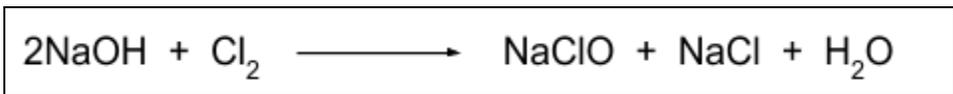
Example:



This reaction is used in **water treatment systems** where chlorine is used in small quantities to kill bacteria. This poses some risks as chlorine can be **toxic** - however the benefits of clean, treated water outweigh the risks.

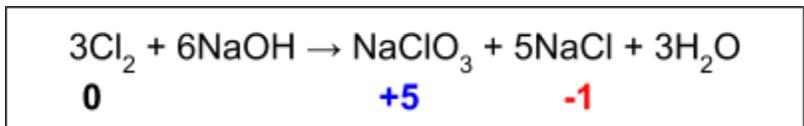
A similar reaction of chlorine with cold dilute sodium hydroxide is used in bleach production. **Sodium chlorate(I)**, a key ingredient in the production of bleach, is produced via this **disproportionation** reaction.

Example:



If chlorine is reacted with **hot, concentrated alkali**, chlorine is disproportionated even further to form one species with an oxidation number of **-1** and another with an oxidation number of **+5**.

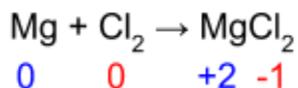
Example:





Oxidation Reactions with Group 1 and Group 2 Metals

Group 1 and Group 2 metals react with chlorine gas to form **metal chlorides**, which are all **white precipitates**. During these reactions **chlorine is reduced** and the **metal is oxidised**.



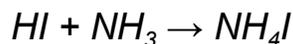
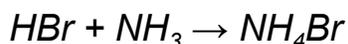
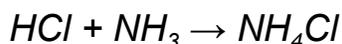
Characteristic Reactions of Halides

When combined with acidified **silver nitrate**, halide ions react to form different **coloured precipitates**, depending on the ion present. The precipitates formed can be used to identify which halide is present in a solution. It may not always be clear to distinguish the colour of the precipitate so they can be tested further using **ammonia**.

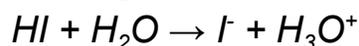
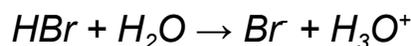
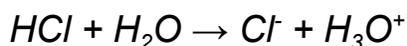
Halide ion	Cl ⁻	Br ⁻	I ⁻
+ AgNO ₃	White precipitate (AgCl)	Cream precipitate (AgBr)	Yellow Precipitate (AgI)
+ dilute NH ₃	Precipitate dissolves	No Change	No Change
+ conc. NH ₃	Precipitate dissolves	Precipitate dissolves	No Change

Reactions of Hydrogen Halides

Hydrogen halides react with ammonia gas to form **ammonium salts**. The hydrogen halides (hydrogen chloride, hydrogen bromide and hydrogen iodide) are strong acids in solution and react with ammonia in an **acid-base reaction** to form a salt.



Hydrogen halides react with water to form **dilute acids**. In solution, these strong acids **dissociate** to release their halide ions and hydrogen ions. The hydrogen ions form a **hydroxonium ion** with water molecules in solution. The resulting solution is **acidic**.





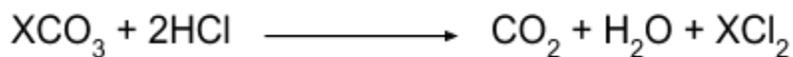
3.1.4 Qualitative Analysis

Tests for Ions

The following tests can be used to **determine** if any of these **ions are present**. They should be performed in the order carbonate, sulfate, then halide if a mixture of ions is present.

Carbonate (CO_3^{2-}) and Hydrogencarbonate (HCO_3^-)

When an acid such as HCl is added, the substance containing the carbonate ions will **fizz and CO_2 gas is given off**. This gas can be collected and bubbled through **limewater** which will turn **cloudy**, confirming its identity as carbon dioxide.



Sulfate (SO_4^{2-})

These are tested for using **acidified BaCl_2** which reacts to form a **white precipitate** of barium sulfate.



Halide ions (Cl^- , Br^- , I^-)

When combined with **acidified silver nitrate**, halide ions react to form different **coloured precipitates** depending on the ion present. The halide can be tested further with **ammonia**.

Halide Ion	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

Ammonium (NH_4^+)

If ammonium ions are present, adding NaOH and gently warming results in the formation of **ammonia gas**, which is **basic**. Therefore, the presence of ammonium ions can be tested by holding **damp red litmus paper** over a petri dish of the substance being tested. It will **turn blue** if ammonium ions are present.

