Edexcel Chemistry A-Level

Topic 4: Inorganic Chemistry and the Periodic Table

Detailed Notes

Topic 4A: The Elements of Group 1 & 2

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. The various physical properties of these elements follow trends down the group:

Atomic Radius

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

Reactivity

Increased electron **shielding** down the group makes the outer electrons **easier to lose**. Therefore reactivity of the group II metals **increases** down the group.

Ionisation Energy

The first ionisation energy of group II metals **decreases** down the group due to a greater atomic radius and increased amounts of shielding.

Melting Point

The group II elements are **metallic** meaning the larger the ions within the metallic structure, the weaker the attractive forces as the attractive force has to **act over a much greater distance**. Therefore the melting points **decrease** down the group.

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 II metals are known as the alkaline earth metals.

Example:

$$\boxed{Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2}$$

The Magnesium is oxidised from OS 0 to OS +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:

$$Mg + H_2O \longrightarrow MgO + H_2$$

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:

$$X + 2HCl \longrightarrow XCl_2 + H_2$$

Sulfuric Acid:

$$X + H_2SO_4 \longrightarrow XSO_4 + H_2$$

Nitric Acid:

$$X + 2HNO_3 \longrightarrow H_2 + X(NO_3)_2$$

Solubility of Group 2 Hydroxides

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

Solubility of Group 2 Sulfates

Group II 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:

$$Ba^{2+} + SO_4^{2-} \longrightarrow BaSO_{4(s)}$$

Thermal Decomposition

Group 2 carbonates and nitrates undergo **thermal decomposition** to produce solid metal oxides and a variety of gases. This is done by simply **heating** the substance in **aerobic** conditions (plenty of oxygen).

The group 2 **carbonates** produce **carbon dioxide** along with the **metal oxide** which forms as a white powder.
Example:

$$XCO_3 \longrightarrow XO + CO_2$$

More heat is required for this reaction **down group 2** as the ions increase in size and the carbonate increases in thermal stability.

The group 2 **nitrates** produce **nitrogen dioxide and oxygen** along with the white powdered **metal oxide**.
Example:

$$2X(NO_3)_2 \longrightarrow 2XO + 4NO_2 + O_2$$

More heat is required for these reactions **down group 2** as the ions increase in size and the nitrate increases in thermal stability.

Group 1 Elements

The Group 1 metals lose an electron to **form 1+ ions** when they react in order to achieve a full outer shell.

Thermal Decomposition

Group 1 carbonates and nitrates also undergo **thermal decomposition** upon heating in **aerobic** conditions. In these reactions, **lithium** acts in a similar way to the group 2 elements, with lithium nitrate producing lithium oxide, nitrogen oxide and oxygen and lithium carbonate producing lithium oxide and carbon dioxide.

Example:

$$4LiNO_3 \longrightarrow 2LiO_2 + 4NO_2 + O_2$$

The rest of the group 1 metals don't react in this same way as the reactions **rarely go to completion**. Group 1 nitrates produce a **metal nitrite and oxygen** where as the remaining group 1 carbonates simply **won't decompose** upon heating in a lab as they require **extremely high temperatures** to do so.

Example:

$$2XNO_3 \longrightarrow 2XNO_2 + O_2$$

Flame Tests

Group 1 and group 2 elements can be identified, along with group 2 elements using flame tests as each one has a unique colour.

Example:

Group 1	**Colour**
Lithium	Red
Sodium	Orange/Yellow
Potassium	Lilac
Group 2	**Colour**
Magnesium	No colour
Calcium	Brick red
Strontium	Crimson red
Barium	Pale green

Topic 4B: The Elements of Group 7 (halogens)

Trends in Halogen Properties

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.

Atomic Radius

The atomic radius of group 7 elements **increases down the group** due to additional electron shells.

Ionisation Energy

The first ionisation energy of group II metals **decreases down the group** due to a greater atomic radius and increased amounts of shielding.

Boiling Point

The group 7 elements are **simple covalent molecules** held together with **van der waals** forces. The strength of these intermolecular forces increases as the Ar of the molecule increases. Therefore the strength of the van der waals forces **increases down the group** meaning more energy is required to overcome them, resulting in a higher boiling point. Fluorine is a gas at room temperature whereas iodine is a solid.

Reactivity

The group 7 elements need to gain an electron. As atomic radius increases this becomes harder as the positive attraction of the nucleus is weakened by additional **shielding**. Therefore it is harder to attract an electron so **reactivity decreases** down the group.

Oxidising Power of the Halogens

The halogens act a **good oxidising agents** as they accept electrons from the species being oxidised and are 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.
Example:

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.

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

Example:

1. Fluoride and Chloride ions.

$$NaF + H_2SO_4 \longrightarrow NaHSO_4 + HF$$

$$NaCl + H_2SO_4 \longrightarrow NaHSO_4 + HCl$$

2. Bromide ions.

$$NaBr + H_2SO_4 \longrightarrow NaHSO_4 + HBr$$

$$2HBr + H_2SO_4 \longrightarrow Br_2 + SO_2 + 2H_2O$$

3. Iodide ions.

$$NaI + H_2SO_4 \longrightarrow NaHSO_4 + HI$$

$$2HI + H_2SO_4 \longrightarrow I_2 + SO_2 + 2H_2O$$

$$6HI + SO_2 \longrightarrow 2H_2S + 3I_2 + 2H_2O$$

The **greater** the reducing power, the **longer** the reaction as the halide is powerful enough to reduce more species.

Disproportionation Reactions

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:

$$\begin{array}{lcl} Cl_2 + H_2O & \longrightarrow & ClO^- + Cl^- + 2H^+ \\ 0 & & +1 \qquad -1 \end{array}$$

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:

$$2NaOH + Cl_2 \longrightarrow NaClO + NaCl + H_2O$$

Precipitation 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**.

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

Topic 4C: Analysis of Inorganic Compounds

Tests for Anions

Sulfate (SO_4^{2-})

These are tested for using **$BaCl_2$** wich reacts to form a **white precipitate** of barium sulfate.
Example:

$$\boxed{BaCl_2 + XSO_4 \longrightarrow BaSO_4 + XCl_2}$$

Hydroxide (OH^-)

These ions indicate that the substance is alkaline. Therefore they can be identified with **damp red litmus paper, which turns blue** or using universal indicator, which turns blue-purple.

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 it as carbon dioxide.
Example:

$$\boxed{XCO_3 + 2HCl \longrightarrow CO_2 + H_2O + XCl_2}$$

Tests for Cations

Ammonium (NH_4^+)

If ammonium ions are present, ammonia gas is given off, which is a base. 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.

Alternatively, they can be tested for in the same way but by **adding NaOH** to produce the ammonia gas faster.