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

When **Group 2** metals react, they lose two electrons to **form 2+ ions**.This allows them 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 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.

Ionisation Energy

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

Reactions of Group 2 Elements and 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:

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

The magnesium is oxidised from an oxidation state of 0 to +2

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**. Magnesium burns in steam with a **bright white flame** to form hydrogen and magnesium oxide, a **white powder**.

Example:

$$\boxed{Mg + H_2O \longrightarrow MgO + H_2}$$

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:

$$Mg + Cl_2 \rightarrow MgCl_2$$

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:

$$2M + O_2 \rightarrow 2MO$$

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

$$M + O_2 \rightarrow MO_2$$

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:

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

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

$$\boxed{Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O}$$

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 ($Mg(OH)_2$) 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 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:

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

Thermal Decomposition

Thermal Decomposition of Group 2 Metals

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).

Group 2 **carbonates** produce **carbon dioxide** along with the **metal oxide** which forms as a white powder.

Example:

$$XCO_3 \longrightarrow XO + CO_2$$

As you go down Group 2, **more heat** is required for this reaction as the ions increase in size and the carbonates increase in thermal stability.

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

Example:

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

Again, as you go down Group 2, **more heat** is required for this reaction as the ions increase in size and the nitrates increase in thermal stability.

Thermal Decomposition of Group 1 Metals

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:

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$

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

Example:

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

Explanation of the Thermal Stability Trend

The elements of Group 1 and Group 2 form more **stable** carbonates and nitrates as you go down their groups, and so require more heat energy to undergo **thermal decomposition**.

In Groups 1 and 2, as you go down the group, the ionic radius increases for the same overall charge. This means smaller ions have a **higher charge density.**

The smaller ions are able to **polarise** the negative carbonate and nitrate ions more and the more this negative ion is polarised, the less heat is required to separate the two ions. Therefore smaller ions form **less stable** carbonates and nitrates, and larger cations form more stable carbonates and nitrates.

Flame Tests

Group 1 and Group 2 elements can be identified using flame tests as each metal has a **unique flame colour**.

Procedure

1. Take a **nichrome wire** and clean it by placing it into a solution of concentrated hydrochloric acid and then into a blue bunsen burner flame.
2. Repeat this **cleaning** until there is no colour produced in the flame.
3. Dip the wire into the unknown metal compound and place it into the flame. **Observe** the flame colour produced.

The colours observed in the flame tests of Group 1 and Group 2 metal compounds:

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

The formation of these colours in the flame test can be explained by **electron transitions.**

Electrons exist in orbitals and when energy (such as that supplied by a bunsen burner flame) is absorbed by the species, some **electrons** will be **promoted** to **orbitals** which have a **higher energy.** The electrons will then drop back down to the original orbital they were in and during this process they **release energy**, some of which is in the form of **light**.

The colour produced depends on the **wavelength** of the light energy emitted. When some electrons drop back down to their original orbital, the energy emitted may not have a wavelength in the visible light part of the **electromagnetic spectrum**, which explains why not all atoms and ions produce a colour in the flame test.

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

Trends in Halogen Properties

The Group 7 elements are **highly reactive non-metals.** 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,** held together with **van der waals** forces. The strength of these intermolecular forces increases as the relative atomic mass of the molecule increases. Therefore, 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

The Group 7 elements 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**.

Oxidising Power of the Halogens

The halogens act as 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. Thus, **Cl_2** will displace Br^- and I^- ions, **Br_2** will displace I^- ions and **I_2** will not 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.

Reactions of Group 7 Elements and Ions

Redox Reactions

The redox reactions between Group 7 halides and H_2SO_4 vary depending on the **reducing ability** of the halide. The reducing ability of the halide can be observed by the relative reactions of the halides with sulfuric acid:

Example:

1. **Fluoride and Chloride ions**

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

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

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**

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

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

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

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

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

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

HI will be observed as **misty fumes** from the first reaction. Similarly to above, 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, 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

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:

$$\underset{0}{Cl_2} + H_2O \longrightarrow \underset{+1}{ClO^-} + \underset{-1}{Cl^-} + 2H^+$$

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.

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

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.

$$\underset{0}{3Cl_2} + 6NaOH \rightarrow \underset{+5}{NaClO_3} + \underset{-1}{5NaCl} + 3H_2O$$

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.**

Example:

$$\underset{0}{Mg} + \underset{0}{Cl_2} \rightarrow \underset{+2\ -1}{MgCl_2}$$

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

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.

Example:

$$HCl + NH_3 \rightarrow NH_4Cl$$

$$HBr + NH_3 \rightarrow NH_4Br$$

$$HI + NH_3 \rightarrow NH_4I$$

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**.

Example:

$$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

$$HBr + H_2O \rightarrow Br^- + H_3O^+$$

$$HI + H_2O \rightarrow I^- + H_3O^+$$

Predictive Chemistry

Using the **trends** in physical properties, chemical properties and reactions described for chlorine, bromine and iodine, **predictions** can be made for other halogens, including **fluorine** and **astatine**.

Example:

Use melting and boiling point trends in group 7 to predict the states of fluorine and astatine at room temperature.

Melting and boiling points increase as you go down the group since more intermolecular Van der waals forces form with increasing Ar. Therefore, fluorine would be a gas at room temperature and astatine would be a solid at room temperature.

Topic 4C: Analysis of Inorganic Compounds

Tests for Anions

Sulfate (SO_4^{2-})

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

Example:

$$BaCl_2 + XSO_4 \longrightarrow BaSO_4 + XCl_2$$

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.

Example:

$$XCO_3 + 2HCl \longrightarrow CO_2 + H_2O + XCl_2$$

Tests for Cations

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.