AQA Chemistry A-level

3.2.3: Group 7 - The Halogens

Detailed Notes

3.2.3.1 - Trends in 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.

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.

Ionisation Energy

The first ionisation energy of group 7 elements **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.

Oxidising Power

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.

These redox reactions with H_2SO_4 have to be known:

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.

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

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

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

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

Silver Nitrate

Acidified silver nitrate is used to **test for halide ions** as it reacts 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**.

	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

3.2.3.2 - Chlorine and Chlorate(I) ions

Chlorine reacts with cold water to produce **Chlorate(I) ions (ClO^-)** and **chloride ions**.

Example:

$$Cl_2 + H_2O \longrightarrow ClO^- + Cl^- + 2H^+$$

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

In the presence of **UV light**, chlorine decomposes water to produce **oxygen and hydrochloric acid**. The chlorine is reduced in this reaction.
Example:

$$2Cl_2 + 2H_2O \xrightarrow{UV} 4HCl + O_2$$

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

Bleach Production

Chlorine can be mixed with **cold, aqueous sodium hydroxide** to produce **sodium hypochlorite**. This is a key ingredient in the production of bleach.

Example:

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

Tests for ions

You need to know how to test for certain anions and cations for your required practical.

Anions - Halides

These are tested for using acidified **silver nitrate and ammonia**. The silver nitrate is acidified so that any other impurities that could form a precipitate are removed.

	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

Anions - Sulfate (SO_4^{2-})

These are tested for using **$BaCl_2$** which reacts to form a **white precipitate**.

Anions - Hydroxide (OH^-)

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

Anions - Carbonate (CO_3^{2-})

When an acid such as HCl is added, the substance containing the carbonate ions will **fizz (effervescence) 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.

Cations - Group 2

The group 2 ions can be identified with a series of **flame tests**.

Calcium (Ca^{2+})	Brick red
Strontium (Sr^{2+})	Red
Barium (Ba^{2+})	Pale green

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 **red litmus** 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.