Group 7 part 1

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It is important to note that this topic can be split into **halogens** and **halide ions**.

✓ students often wrongly interchange halogens and halides. Please remember that halogens and halides are not the same thing and that they react very differently.

halogens are diatomic elements e.g. F_2, Cl_2

halide ions e.g., F^-, Cl^-

Halogens

States

Halogens exist in different physical states:

F_2 is a yellow gas

Cl_2 is a green gas

Br_2 is a dark orange liquid

I_2 is a grey solid

✓ At_2 is at the bottom of group 7. It is a solid but hardly ever gets mentioned.

Boiling Points

You can also see how the boiling point varies down the group from their states. The gases are at the top (lowest boiling point) versus solids at the bottom (highest boiling points).

This is simply to do with **size** (again)….bigger molecules at the bottom of the group having more **van der Waals** that need to be broken when boiling.

Colours in solution

Halogens are not very soluble in water. They are **non-polar** molecules and water is a very polar solvent. They don't like each other.

But halogens will dissolve happily in **non-polar organic covalent solvents** like hexane as they can form London forces with hexane.

	In Hexane	In Water
Chlorine	colourless	colourless
Bromine	red	orange
Iodine	violet	brown

- ✓ **Iodine** is a favourite here. Chlorine less so as the colours are non-existent and everyone is happy with bromine from the alkene test.

Reactivity

Halogens become less reactive as you go down a group: $\mathbf{F_2 > Cl_2 > Br_2 > I_2 > At_2}$

This is simply to do with **atomic radius** and **electronic configuration**.

For example, fluorine has electronic configuration $\mathbf{1s^2\ 2s^2\ 2p^5}$**.** When fluorine reacts it will become F^-. This means it needs to **gain one electron** to fill its outer shell to form the F^- ion and become $2p^6$.

> The reactivity of the halogens is determined by how easily they can **attract** an electron.

The smaller the atomic radius, then the more this "extra" electron will feel the attractive force of the nucleus and the easier it will be to form the halide ion.

As fluorine is the halogen with the smallest radius, it is therefore the most reactive.

- ✓ slightly off topic but you can reverse this theory for Group 1 or 2. Metal ions are trying to lose electrons and therefore the larger they are, the easier it is to do it. Hence why reactivity increases down Group 1 or 2.

Halogen Reactions

Redox/Displacement

> Halogens are **oxidising agents**

Fluorine is the strongest/most reactive of those in Group 7.

A common reaction to demonstrate the differences in reactivity is to react a halogen with a metal halide (see examples below).

This reaction can be viewed as **displacement** or **redox**.

A halogen will oxidise/displace a halide ion in the salt if the **halogen formed is below** it in the group.

For example:

$$\mathbf{F_2 + 2KCl \rightarrow Cl_2 + 2KF}$$

As **Cl_2 is below F_2** in the group then the Cl^- ion will be oxidised/displaced. F_2 is oxidising/displacing Cl^- to form Cl_2 and F^-.

This is a redox reaction as **Cl^- is oxidised to Cl_2** and **F_2 is reduced to F^-.** They often ask for the ionic equation. Just remove the "K" (spectator ion):

$$\mathbf{F_2 + 2Cl^- \rightarrow Cl_2 + 2F^-}$$

If you did it the other way round:

$$\mathbf{Cl_2 + KF \rightarrow \text{no reaction}}$$

As **F_2 is above Cl_2** in Group 7, then Cl_2 is not a powerful enough oxidising agent to oxidise the F^- ion.

These reactions are of course accompanied by colour changes, so the table we looked at earlier will be useful.

Disproportionation

A disproportionation reaction is where a species is **both** oxidised and reduced at the same time. Yes this is possible! You won't see this reaction very often other than in Group 7 questions.

For oxidation and reduction to happen at the same time, we have to do something like this:

$$\mathbf{Cl_2 \rightarrow Cl^+ + Cl^-}$$

Oxidation state of **0 in Cl_2** to **1+ and 1-**

You probably haven't come across the Cl^+ ion but it can be formed! The halogens can change their oxidation states.

Disproportionation (reaction with cold NaOH)

A favourite question is to ask about the reaction of a halogen with NaOH to make bleach (NaClO), which just happens to be a disproportionation reaction.

They could ask this for any halogen but it's usually Cl_2, as they need the Cl part to make the bleach:

$$\mathbf{Cl_2 + 2NaOH \rightarrow NaClO + NaCl + H_2O}$$

It is worth remembering this equation, it comes up a lot in exams.

NaClO contains the Cl^+ ion and NaCl, the Cl^- ion. The **ClO^-** ion is called the hypochlorate ion and is responsible for killing bacteria.

✓ the solution with the ClO^- ion is too dilute to do any harm to humans.

Disproportionation (reaction with water)

Another disproportionation example can be seen in the reaction of Cl_2 and H_2O.

We know that, for example, they add chlorine to swimming pools to kill bacteria. As above, this is due to the formation of the ClO^- ion:

$$\mathbf{Cl_2 + 2H_2O \rightarrow HClO + HCl}$$

✓ the Cl_2 has gone to Cl^+ and Cl^- again.

In sunlight

$$\mathbf{2Cl_2 + 2H_2O \rightarrow 4HCl + O_2}$$

When the same reaction is carried out on sunlight, **no ClO^-** is formed and **no disproportionation** occurs. It is just a redox reaction. The Cl is reduced and the O is oxidised.

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