AQA GCSE Chemistry: Combined Science

Your notes

1.2 The Periodic Table

Contents

Your notes

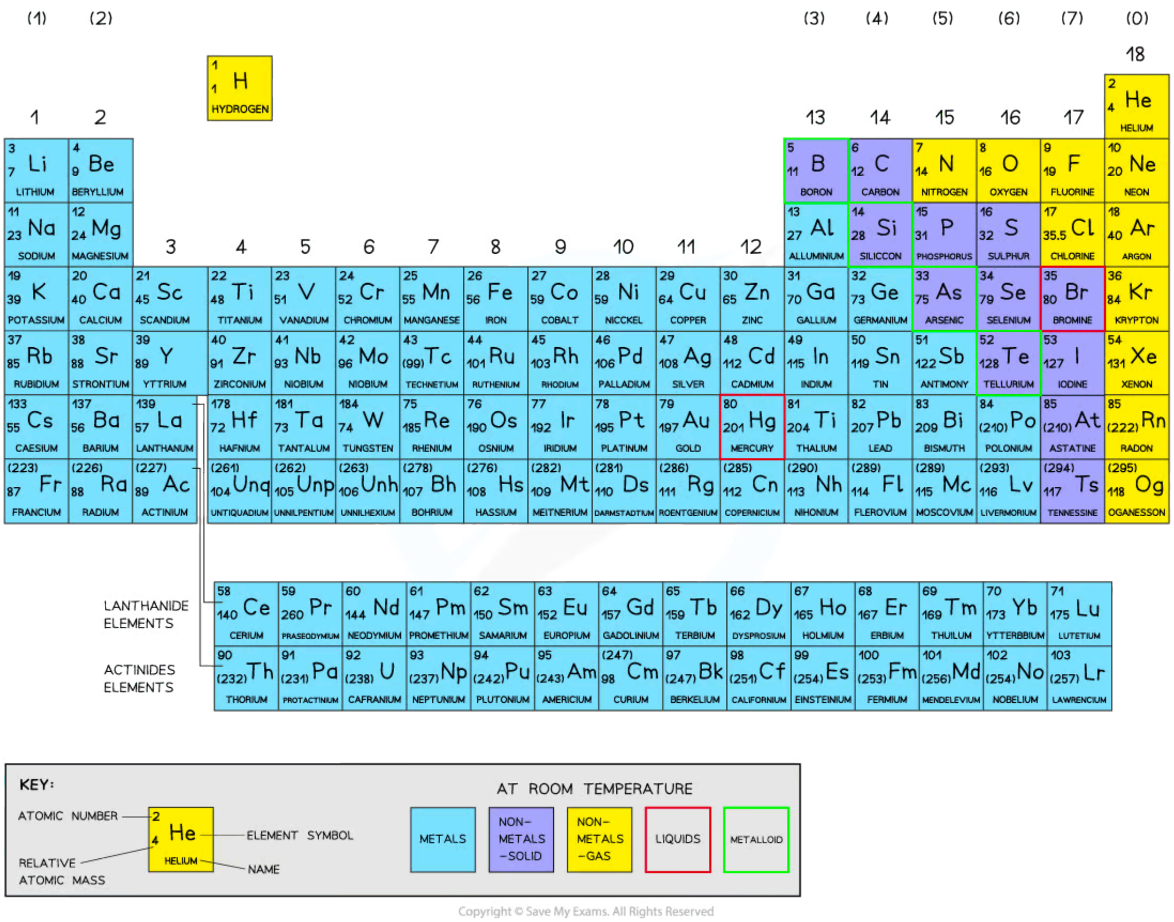
Arranging the Elements

How the elements are ordered

- There are over 100 chemical elements which have been isolated and identified
- Elements are arranged on the periodic table in order of **increasing atomic number**
  - Each element has one proton **more** than the element preceding it
  - This is done so that elements end up in columns with other elements which have similar properties
- The table is arranged in vertical columns called **groups** and in rows called **periods**
  - **Period:** These are the horizontal rows that show the number of shells of electrons an atom has and are numbered from 1 - 7
  - E.g. elements in Period 2 have two electron shells, elements in Period 3 have three electron shells
- **Group:** These are the vertical columns that show how many outer electrons each atom has and are numbered from 1 – 7, with a final group called Group 0 (instead of Group 8)
  - E.g. Group 4 elements have atoms with 4 electrons in the outermost shell, Group 6 elements have atoms with 6 electrons in the outermost shell and so on

The Periodic Table

Your notes



The Periodic Table of the Elements

Examiner Tip

The atomic number is unique to each element and could be considered as an element's "fingerprint".

The number of electrons changes during chemical reactions, but the atomic number does not change.

Predicting reactions

- The group number of an element which is given on the Periodic Table indicates the number of electrons in the outer shell (**valence** electrons)
  - This rule holds true for all elements except **helium**; although is in Group 0, it has only **one** shell, the first and innermost shell, which holds only 2 electrons

- We can use the group number to predict how elements will react as the number of valence shell electrons in an element **influences** how the element reacts.
- Therefore, elements in the same group react **similarly**
  - By observing the reaction of one element from a group, you can predict how the other elements in that group will react
  - By reacting two or more elements from the same group and observing what happens in those reactions you can make predictions about reactivity and establish **trends** in reactivity in that group

- For example, lithium, sodium and potassium are in Group 1 and can all react with elements in Group 7 to form an ionic compound
- The Group 1 metals become **more reactive** as you move down the group while the Group 7 halides show a **decrease** in reactivity moving down the group

Examiner Tip

The word "periodic" is used in the name of the Periodic Table as similar properties appear in elements placed at regular intervals throughout the table.

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History of the Periodic Table

Your notes

How was the early Periodic Table arranged?

- Before the discovery of the subatomic particles, scientists arranged the elements in order of their atomic **weight** and not their atomic **number**
- When the elements that were known at that time were sorted by mass into a table, patterns emerged at **regular periods** along the table, giving rise to the term periodic
- The earlier tables were incomplete as some elements were forced into a position to fill **gaps** which appeared during the sorting process
- Other elements were placed in the wrong group as they were sorted strictly on their atomic weight and had their chemical properties ignored
- There were many early versions of the tables as scientists in different countries grappled with the ordering of the elements

How did Mendeleev arrange the Periodic Table?

- In 1869, the Russian chemist Dmitri Mendeleev created his first draft of the periodic table
- He organised the elements into **vertical columns** based on their properties and the properties of their compounds
- He then started to arrange them **horizontally** in order of **increasing atomic weight** and as he worked, he found that a **pattern** began to appear in which chemically similar elements fell naturally into the same columns

Mendeleev's Periodic Table

ROW	GROUPS							
	I	II	III	IV	V	VI	VII	VIII
1	H	–	–	–	–	–	–	–
2	Li	Be	B	C	N	O	F	–
3	Na	Mg	Al	Si	P	S	Cl	–
4	K	Ca	?	Ti	V	Cr	Mn	Fe, Co, Ni, Cu
5	(Cu)	Zn	?	?	As	So	Br	–
6	Rb	Sr	Yt	Zr	Nb	Mo	?	Ru, Rh, Pd, Ag

Mendeleev's Periodic Table showed gaps to allow for undiscovered elements

Why are there gaps in Mendeleev's Periodic Table?

- There were exceptions though as some elements didn't fit the pattern when arranged by atomic weight
- Mendeleev worked to include all the elements, but he didn't force an element to fit the pattern, rather he left **gaps** in the table that he thought would best be filled by elements that had not yet been

Your notes

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discovered

- He also **switched the order** of the elements to maintain consistency down the columns
- He quickly realised that elements with the **same properties** should be placed in the **same column**
- He realised that gaps in the table must correspond to elements that had not yet been discovered or isolated
- He used the properties and trends of other elements in the group with the gap to **predict** the properties of these undiscovered elements
- When these elements were later discovered and found to fit the pattern developed by Mendeleev, it served to confirm his theories
- The existence and properties of "eka-silicon" for example, which we now know as germanium, was predicted by Mendeleev

Examiner Tip

Mendeleev's table had gaps into which he didn't force an element, rather he left them empty to be filled at a later date when the correct element was isolated. In this way, his version of the table allowed him to predict the existence and properties of then-unknown elements.

Isotopes & the Periodic Table

- Once he was finished, Mendeleev thought he had organised the elements systematically but there were still some elements which didn't quite fit in as neatly as he wanted
- This is because **isotopes** were not known in Mendeleev's time, and he made no provisions for them in his table
- This meant that there was always going to be some level of inaccuracy in Mendeleev's work even though he did also consider the element's chemical properties as well as their atomic mass when sorting them
- As soon as the subatomic particles were discovered, the **atomic number** was calculated for each element
- This number is used to arrange the elements in the modern-day Periodic Table which fits with Mendeleev's patterns

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Metals & Non-Metals

Metals & non metals in the Periodic Table

- The elements can be divided into two broad types: **metals** and **non-metals**
- Atoms of different elements which do not have a full outer shell of electrons, can try to achieve a full outer shell by gaining or losing electrons in chemical reactions
  - Elements that react by losing electrons to form positive ions are metals
  - Elements that do not form positive ions are non-metals; this includes elements that react by gaining electrons to form negative ions and Group 0 elements
- Most of the elements are metals and a small number of elements display properties of both types
  - These elements are called **metalloids** or **semi-metals**
- The metallic character of the elements **decreases** as you move across a period on the periodic table, from **left** to **right**, and it **increases** as you move down a group
- This trend occurs due to atoms more **readily accepting** electrons to fill their valence shells

Examiner Tip

An ion is an atom or molecule which has become charged through the loss or gain of one or more electron(s).

Metals will form positive ions when they react – when they lose electrons, the atom ends up with more positively charged protons than negatively charged electrons, which leaves it with an overall positive charge.

Non-metals will form negative ions when they react – when they gain electrons, the atom ends up with more negatively charged electrons than positively charged protons, which leaves it with an overall negative charge.

Positive ions are called cations and negative ions are called anions.

Atomic structure & position on the Periodic Table

Your notes

- The metals that are further to the left on the periodic table do not have many electrons to remove from their outer shells
- As you descend the groups, the outer shell electrons become **further away** from the nucleus due to increasing atomic size
  - This **weakens their attraction** to the nucleus
- The further down the group an element is, the more easily it can react and lose its outer electron(s)
- For the non-metals which are placed on the right-hand side, the opposite is the case
- These elements have a lot of outer electrons, and it is more feasible for them to gain (or share electrons) to obtain a full outer shell
  - This is a key difference between metals and non-metals and influences their chemical behaviour
- It also clearly illustrates the important link between an element's atomic number and how it reacts as well as its position on the periodic table

Examiner Tip

Atoms react to gain full outer shells by losing, gaining or sharing electrons.

The number of outer electrons an atom has determines its chemical properties and how it reacts.

Properties of metals & non-metals

Your notes

- The general properties of most metals and non-metals are summarised below

A summary of the general properties of metals & non-metals

Property	Metals	Non-metals
Electron arrangement	1–3 outer shell electrons	4–7 outer shell electrons
Bonding	Metallic bonding due to loss of electrons	Covalent by sharing of outer shell electrons
Electrical conductivity	Good conductor of electricity	Poor conductors of electricity
Type of oxide	Basic oxides	Acidic oxides (some are neutral)
Reaction with acids	Many react with acids	Usually do not react with acids
Physical characteristics	▪ Usually lustrous (shiny) ▪ Solid at room temperature (excluding mercury) ▪ Malleable, can be bent and shaped ▪ High melting and boiling point	▪ Dull, non-reflective ▪ Different states at room temperature ▪ Flaky, brittle ▪ Low melting and boiling points

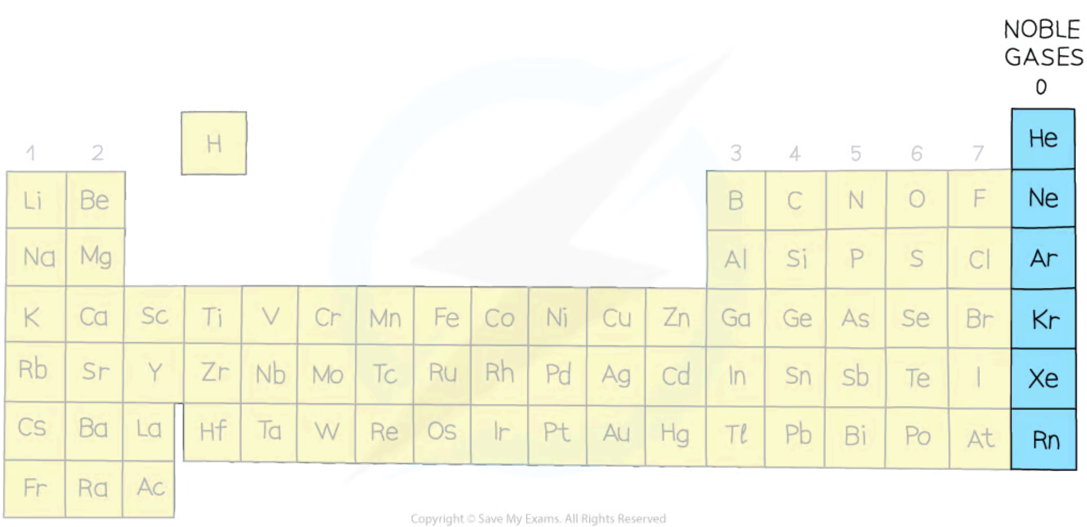
Group 0: The Noble Gases

Your notes

The noble gases

- The elements in Group 0 of the Periodic Table are called the noble gases
- Noble gases are:
  - Non-metals
  - Monatomic (exist as single atoms)
  - Colourless and non-flammable gases at room temperature
- Most elements participate in reactions to complete their outer shells by losing, gaining, or sharing electrons
- Group 0 elements do **not** do this because they have **full outer shells of electrons**
  - They are therefore **unreactive (inert)** and do not form molecules easily
- Most noble gases have 8 electrons in their outer shell, except helium which has 2
- Electronic configurations of the noble gases:
  - He = 2
  - Ne = 2, 8
  - Ar = 2, 8, 8
  - Kr = 2, 8, 18, 8
  - Xe = 2, 8, 18, 18, 8
- Being chemically inert makes noble gases useful for many applications
  - Argon is used to provide an inert atmosphere for welding
  - Argon is also used to fill light bulbs

The noble gases



Noble gases are located in the last group on the right hand side of the Periodic Table

Your notes

Examiner Tip

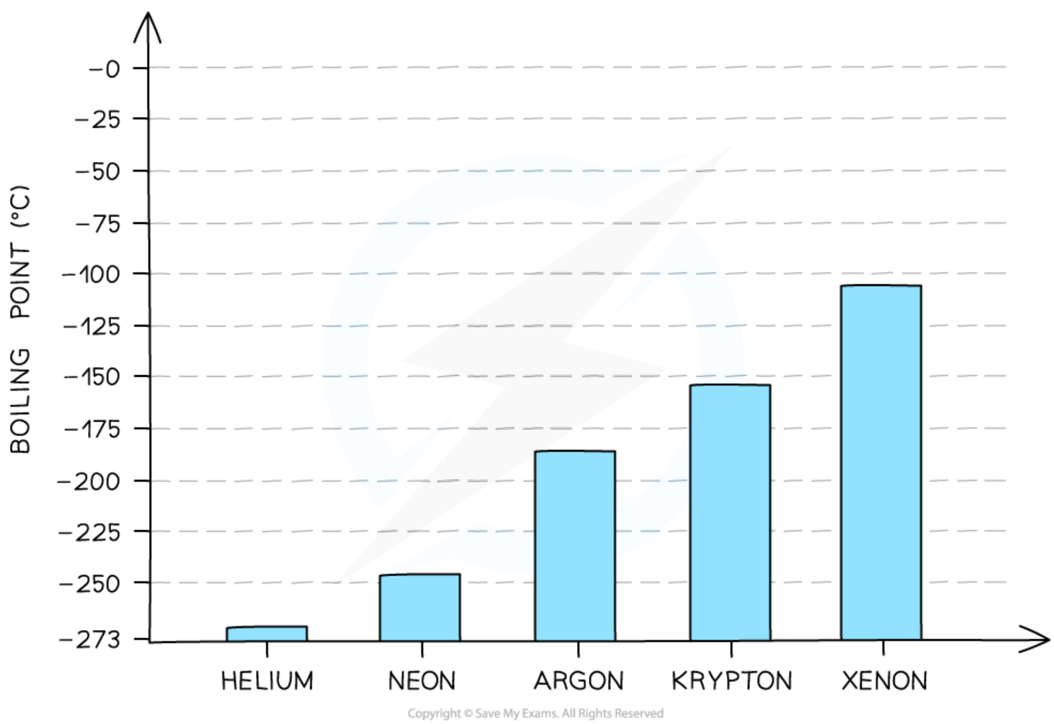
You do not need to know specific uses of noble gases but be aware that they are useful in these applications due to their inertness.

Physical properties of the noble gases

Your notes

- There are trends in the physical properties of the noble gases
- The noble gases have very low melting and boiling points
- Going down the group melting and boiling point **increases** because:
  - The atoms get larger as you move down the group and their **relative atomic mass** increases
  - This leads to an increase in intermolecular forces between atoms
  - Therefore more energy is needed to overcome these
- Although elements further down the group have **higher** boiling points, they still lie **below 0 °C**.
- Helium has the lowest boiling point of all known elements at -269 °C, while radon boils at around -60 °C.

Boiling points of the noble gases



This graph shows the trend in boiling point of the noble gases

Examiner Tip

Exam questions often give you the boiling point of a noble gas and ask you to estimate the value of another one, so it is important to remember the **general** trends in the Group 0 elements. You do not need to learn these values exactly!

Your notes

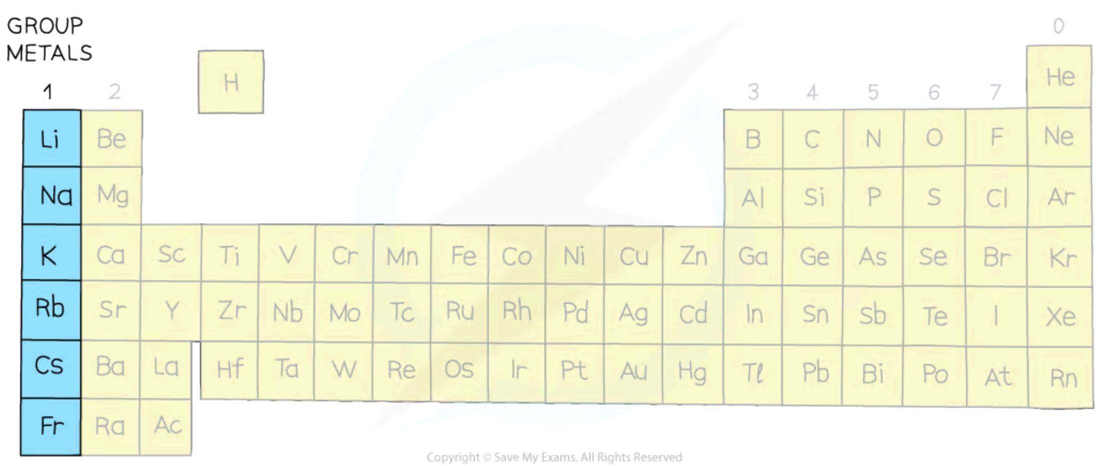
Group 1: The Alkali Metals

Group 1 elements

- The Group 1 elements are known as the alkali metals
  - They form **alkaline solutions** when they react with water

- The Group 1 metals are lithium, sodium, potassium, rubidium, caesium and francium and they are found in the first column of the periodic table
- The alkali metals share similar characteristic chemical properties because they each have one electron in their outermost shell
- Some of these properties are:
  - They are all **soft** metals which can easily be cut with a knife
  - They have relatively **low** densities and **low** melting points
  - They are **very reactive** (they only need to lose one electron to become highly stable)

Group 1 elements in the Periodic Table



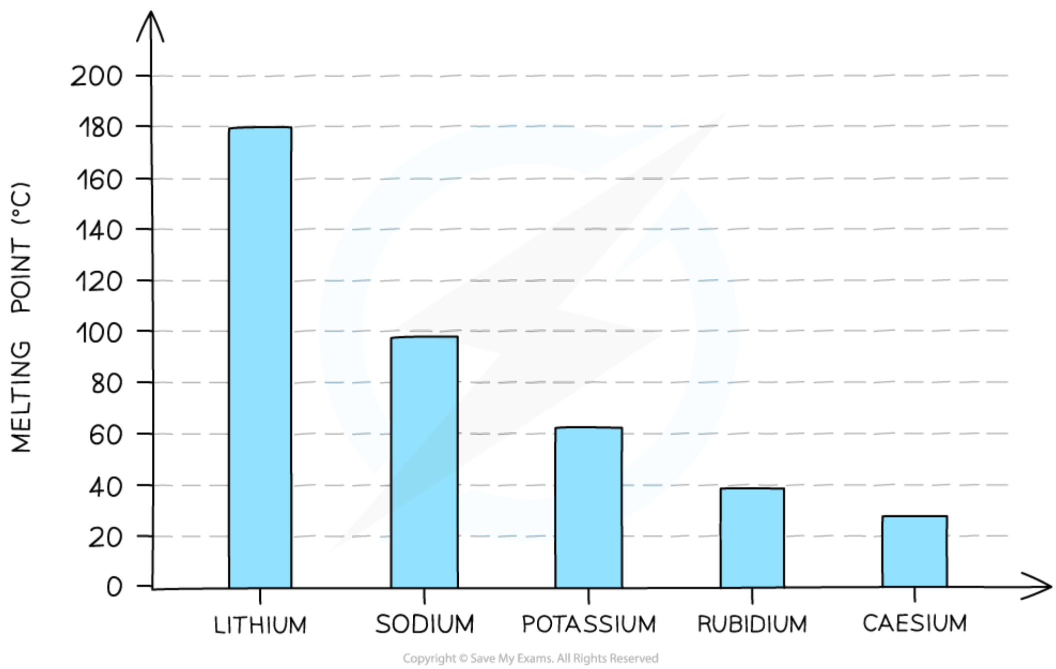
The alkali metals lie on the far left of the periodic table, in the very first group

Physical properties of the Group 1 elements

Your notes

- The alkali metals are **soft** and easy to cut, getting softer as you move down the group
  - Potassium is the exception; it has a lower density than sodium
- The first three alkali metals are less dense than water
- They all have relatively **low** melting points which **decrease** as you move down the group, due to **decreasing attractive forces** between outer electrons and positive ions

Melting points of the Group 1 elements

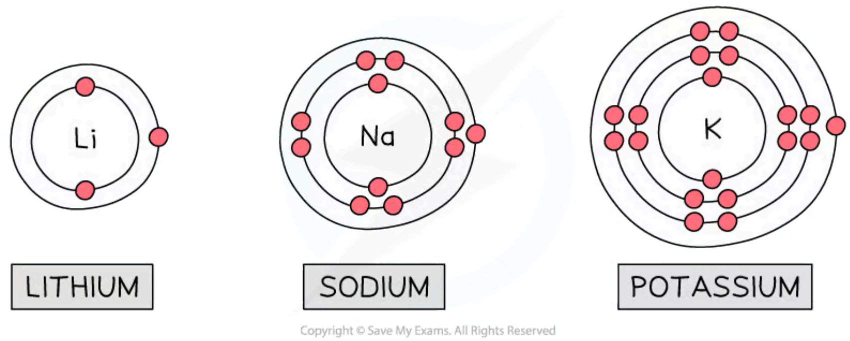


The melting point of the Group 1 metals decreases as you descend the group

- The reactivity of the Group 1 metals **increases** as you go down the group
- When a Group 1 element reacts, its atoms only need to lose the 1 electron in the outer shell
  - When this happens, 1+ ions are formed
- The next shell down automatically becomes the outermost shell and since it is **already full,** a Group 1 atom obtains **noble gas configuration**
- As you go down Group 1, the number of shells of electrons increases by 1
  - This means that the outermost electron gets **further** away from the nucleus, so there are **weaker forces of attraction** between the outermost electron and the nucleus
  - **Less energy** is required to overcome the force of attraction as it gets weaker, so the outer electron is lost **more easily**
  - So, the alkali metals get more reactive as you descend the group

Your notes

The electronic structure of Group 1 elements



These electron shell diagrams of the first 3 alkali metals show that the group 1 metals have 1 electron in their outer shell

Examiner Tip

In your exams, you could be asked to explain the trend in reactivity of the alkali metals - make sure you answer this question using their electronic configuration to support your answer.

Your notes

Group 1 reactivity

- You need to be able to describe the reactions of the first three alkali metals with water, oxygen and chlorine
  - This includes providing reaction equations to show what is happening

- Alkali metals react readily with oxygen and water vapour in air, so they are usually stored in **oil** to stop them from reacting

Reactions with Water

- The reactions of the alkali metals with water get more vigorous as you descend the group, as with the other reactions
- You could be asked to describe and explain the reactions of the alkali metals with water

Summary of the Reactions of the First Three Alkali Metals with Water

Element	Reaction	Observations
Li	lithium + water → lithium hydroxide + hydrogen $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$	▪ Relatively slow reaction ▪ Fizzing ▪ Lithium moves on the surface of the water
Na	sodium + water → sodium hydroxide + hydrogen $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$	▪ More vigorous fizzing ▪ Moves rapidly on the surface of the water ▪ Dissolves quickly
K	potassium + water → potassium hydroxide + hydrogen $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$	▪ Reacts more vigorously than sodium ▪ Burns with a lilac flame ▪ Moves very rapidly on the surface ▪ Dissolves very quickly

- Rubidium, caesium and francium will react even more vigorously with air and water than the first three alkali metals
- Of the alkali metals, lithium is the **least** reactive (as it is at the top of group 1) and francium would be the **most** reactive (as it's at the bottom of group 1)

Reactions with Oxygen

- The alkali metals react with oxygen in the air forming **metal oxides,** which is why the alkali metals tarnish when exposed to the air
- The metal oxide produced is a dull coating which covers the surface of the metal

Summary of the reactions of the first three alkali metals with oxygen

Your notes

Element	Reaction
Li	lithium + oxygen → lithium oxide $4Li\ (s) + O_2\ (g) \rightarrow 2Li_2O\ (s)$
Na	sodium + oxygen → sodium oxide $4Na\ (s) + O_2\ (g) \rightarrow 2Na_2O\ (s)$
K	potassium + oxygen → potassium oxide $4K\ (s) + O_2\ (g) \rightarrow 2K_2O\ (s)$

Reactions with Chlorine

- All the group 1 metals react vigorously when heated with chlorine gas to form salts called **metal chlorides**
- This reaction becomes more vigorous **moving down** the group, the same as with the reaction between the metals and water

Summary of the Reactions of the First Three Alkali Metals with Chlorine

Element	Reaction
Li	lithium + chlorine → lithium chloride $2Li\ (s) + Cl_2\ (g) \rightarrow 2LiCl\ (s)$
Na	sodium + chlorine → sodium chloride $2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$
K	potassium + chlorine → potassium chloride $2K\ (s) + Cl_2\ (g) \rightarrow 2KCl\ (s)$

Examiner Tip

Remember: All Group 1 metals produce alkaline solutions (> pH 7) when they react with water. Lithium will produce a solution of lithium hydroxide; sodium will produce a solution of sodium hydroxide and so on. Make sure you can give the reaction equations with the correct state symbols to show what is happening during the reactions!

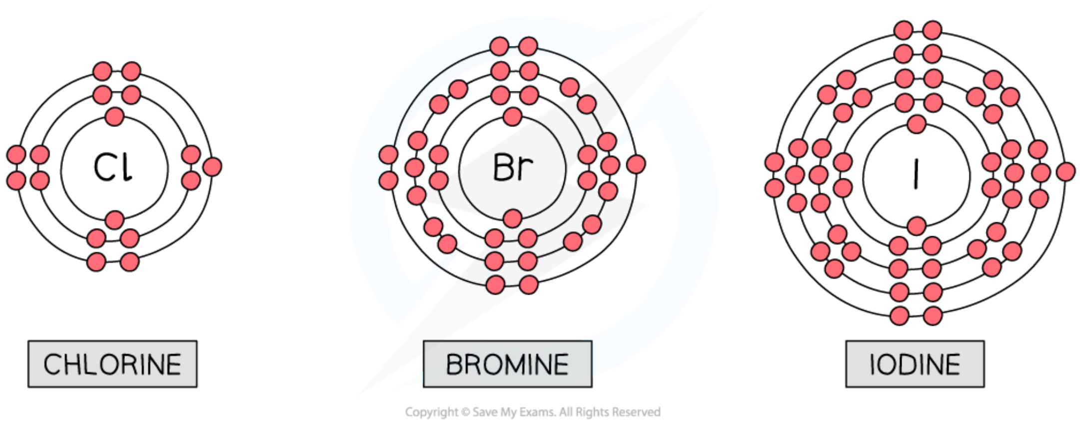
Your notes

Group 7: The Halogens

Atomic structure of Group 7 elements

- The elements in Group 7 are known as the halogens
  - These are fluorine, chlorine, bromine, iodine and astatine
- These elements are non-metals that are **poisonous**
- All halogens have similar reactions as they each have seven electrons in their outermost shell
- Halogens are **diatomic**, meaning they form molecules made of pairs of atoms sharing electrons (forming a single covalent bond between the two halogen atoms) such as F_2, Cl_2, etc
- When halogen atoms gain an electron during reactions, they form -1 ions called halide ions

Atomic structure of Group 7 elements



The atoms of the elements of Group 7 all have 7 electrons in their outer shell

Examiner Tip

You will only be expected to draw the electron configurations for fluorine and chlorine. This is because bromine, iodine and astatine are beyond the GCSE model and specification.

Properties of Group 7 Elements

- At room temperature, the halogens exist in different states and colours, with different characteristics

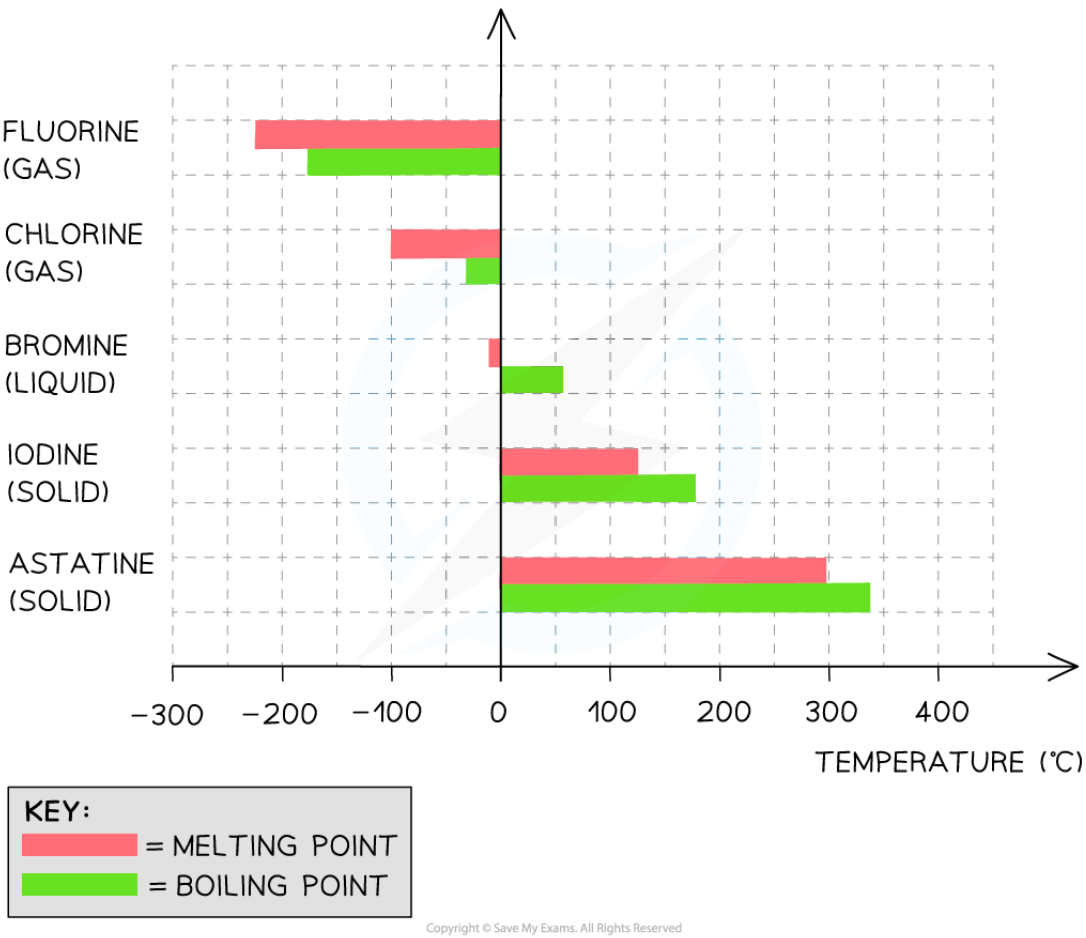
The properties of the Group 7 elements

Halogen	State & appearance at room temperature	Characteristics	Colour in solution
Fluorine	Yellow gas	Very reactive, poisonous gas	-
Chlorine	Pale yellow-green gas	Reactive, poisonous and dense gas	Pale green
Bromine	Red-brown liquid	Dense red-brown volatile liquid	Orange
Iodine	Grey solid	Shimmery, crystalline solid that sublimes to form a purple vapour	Dark brown

- The melting and boiling points of the halogens **increase** as you go down the group
- This is due to increasing intermolecular forces as the atoms become larger, so more energy is required to overcome these forces

Boiling points of the Group 7 elements

Your notes



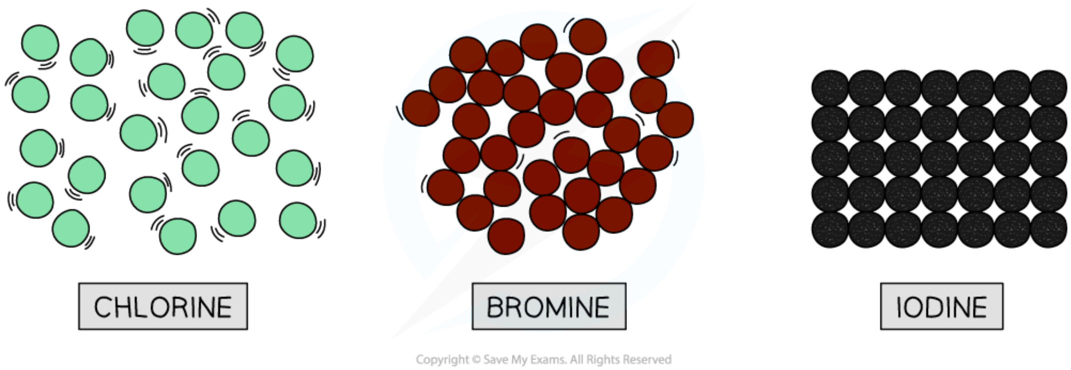
This graph shows the melting and boiling points of the Group 7 elements

Group 7 elements state at room temperature

- At room temperature (20 °C), the physical state of the halogens changes as you go down the group
  - Fluorine and chlorine are **gases**, bromine is a **liquid** and iodine is crumbly **solid**
- The colours of the halogens also change as you descend the group - they become darker

The appearance and state of the Group 7 elements

Your notes



The physical states and colours of chlorine, bromine and iodine at room temperature

Examiner Tip

Exam questions on this topic occur often so make sure you know and can explain the trends of the group 7 elements in detail, using their electron configurations.

Reactivity of the Halogens

Your notes

- Reactivity of Group 7 non-metals **decreases** as you go down the group
  - As you go down Group 7, the number of shells of electrons **increases**, the same as with all other groups
- However, halogen atoms form negative ions when they **gain an electron** to obtain a full outer shell
  - This means that the increased distance from the outer shell to the nucleus as you go down a group makes the halogens become **less reactive**
- Fluorine is the smallest halogen, which means its outermost shell is the **closest** to the positive nucleus of all the halogen
  - Therefore, the ability to **attract an electron** is strongest in fluorine making it the most reactive
  - As you move down the group, the forces of **attraction** between the nucleus and the outermost shell **decreases**
  - This makes it **harder** for the atoms to gain electrons as you descend the group
  - Therefore, the halogens are less reactive the further down the group you go

Displacement Reactions

- A **halogen displacement reaction** occurs when a more reactive halogen displaces a less reactive halogen from an aqueous solution of its halide
- The reactivity of Group 7 elements decreases as you move down the group
- You only need to learn the displacement reactions with chlorine, bromine and iodine
  - Chlorine is the most reactive and iodine is the least reactive

Chlorine with Bromine & Iodine

- If you add chlorine solution to colourless potassium bromide or potassium iodide solution a displacement reaction occurs:
  - The solution becomes orange as bromine is formed or
  - The solution becomes brown as iodine is formed
- Chlorine is **above** bromine and iodine in Group 7 so it is more reactive
- Chlorine will **displace** bromine or iodine from an aqueous solution of the metal halide

chlorine + potassium bromide → potassium chloride + bromine

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$$

chlorine + potassium iodide → potassium chloride + iodine

$$Cl_2 + 2KI \rightarrow 2KCl + I_2$$

Bromine with iodine

Your notes

- Bromine is above iodine in Group 7 so it is **more** reactive
- Bromine will displace iodine from an aqueous solution of the metal iodide

bromine + potassium iodide → potassium bromide + iodine

$$Br_2 + 2KI \rightarrow 2KBr + I_2$$

Summary table of the displacement reactions of the halogens: chlorine, bromine and iodine

	Chlorine (Cl_2)	Bromine (Br_2)	Iodine (I_2)
Potassium chloride (KCl)	x	No reaction	No reaction
Potassium bromide (KBr)	Chlorine displaces the bromide ions Yellow-orange colour of bromine seen	x	No reaction
Potassium Iodide (KI)	Chlorine displaces the iodide ions Brown colour of iodine is seen	Bromine displaces the iodide ions Brown colour of iodine is seen	x

- You could be asked at Higher Tier to provide ionic equations to show what is happening during displacement reactions of the halogens
- These are:

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

$$Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$$

$$Br_2 + 2I^- \rightarrow 2Br^- + I_2$$

Reactions of the Halogens

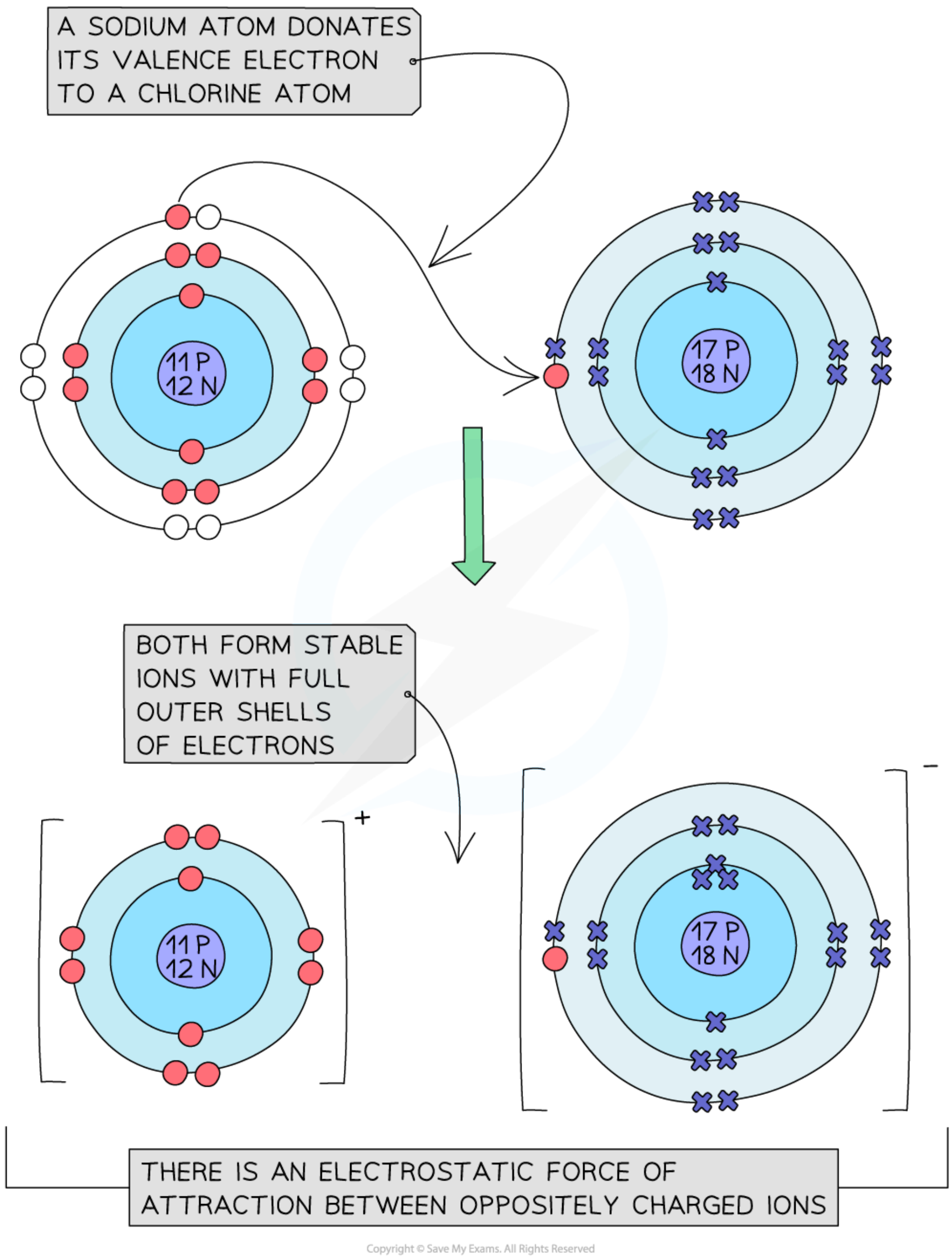
Metal Halides

- Chlorine, bromine and iodine react with metals and non-metals to form compounds
- The halogens react with some metals to form **ionic compounds** which are **metal halide salts**
- The halide ion carries a 1– charge so the ionic compound formed will have different numbers of halogen atoms, depending on the **valency** of the metal
  - E.g., sodium is a group 1 metal:
    - $2Na + Cl_2 \rightarrow 2\,NaCl$
- Calcium is a group 2 metal:
  - $Ca + Br_2 \rightarrow CaBr_2$
- The halogens **decrease** in **reactivity** moving down the group, but they still form halide salts with some metals including iron
- The rate of reaction is **slower** for halogens which are **further** down the group such as bromine and iodine

Formation of sodium chloride

Your notes

Your notes



Sodium donates its single outer electron to a chlorine atom and an ionic bond is formed between the positive sodium ion and the negative chloride ion

Non-metal halides

- The halogens react with non-metals to form simple molecular covalent structures
- For example, the halogens react with hydrogen to form **hydrogen halides** (e.g., hydrogen chloride)
- Reactivity decreases down the group, so iodine reacts less vigorously with hydrogen than chlorine (which requires light or a high temperature to react with hydrogen)
- Fluorine is the most reactive (reacting with hydrogen at low temperatures in the absence of light)
- Other compounds formed include CCl_4, HF and PCl_5

Your notes