WJEC GCSE Chemistry: Combined Science

Your notes

Atomic Structure & The Periodic Table

Contents

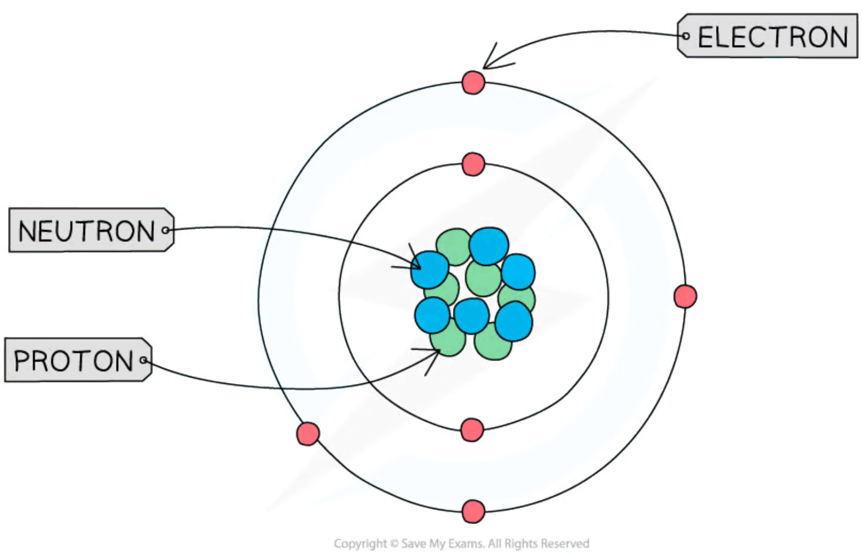
Atomic Structure

Your notes

Atomic Structure

- All substances are made of tiny particles of matter called **atoms** which are the building blocks of all matter
- Each atom is made of subatomic particles called **protons**, **neutrons,** and **electrons**
- The protons and neutrons are located at the centre of the atom, which is called the **nucleus**
  - The nucleus is positively charged
- The electrons move very fast around the nucleus in orbital paths called **shells**
- The mass of the electron is negligible, hence the mass of an atom is contained within the nucleus where the protons and neutrons are located

Atomic Structure Diagram



The structure of the carbon atom

Examiner Tip

The atom is the smallest part of an element that exists even though it can be divided into smaller particles.

This is because the atom is the smallest part of an element that still retains the properties of the element - the subatomic particles do not.

Relative Masses

Your notes

- The **protons**, **neutrons** and **electrons** that an atom is made up of are called **subatomic particles**
- These subatomic particles are so small that it is not practical to measure their masses and charges using **conventional units** (such as grams or coulombs)
- Instead, their masses and charges are compared to each other, and so are called **'relative atomic masses'** and '**relative atomic charges**'
- These are not actual charges and masses, but rather charges and masses of particles relative to each other
  - Protons and neutrons have a very similar mass, so each is assigned a relative mass of **1**
  - Electrons are 2000 times smaller than a proton and neutron, and so their mass is often described as being negligible
- The relative mass and charge of the subatomic particles are:

The Mass & Charge of Subatomic Particles Table

Particle	Relative Mass	Charge
Proton	1	+1
Neutron	1	0 (neutral)
Electron	1/2000	-1

- Atoms are electrically **neutral**
- This is achieved by having the **same number** of electrons as protons
- The **negative charge** of an electron exactly cancels out the **positive charge** of a proton
- An ion is formed when an atom loses or gains electrons to achieve a full outer shell
  - If an atom loses electrons, it forms positively charged ions
  - If an atom gains electrons, it forms negatively charged ions

Worked example

Explain why a magnesium ion has a 2+ charge.

Answer:

- A magnesium atom has 12 positively charged protons and 12 negatively charged electrons
- To gain a full outer shell magnesium loses two electrons
- It now has 12 protons but only 10 electrons
- The overall charge is 2+

Examiner Tip

The mass of an electron can just be stated as 'almost zero' in an exam.

Your notes

Your notes

Atomic Number, Mass Number & Isotopes

Atomic Number, Mass Number & Isotopes

Atomic Number

- The **atomic number** (or **proton number**) is the number of protons in the nucleus of an atom
  - The symbol for this number is ***Z***
- It is also the number of electrons present in an atom and determines the **position** of the element on the Periodic Table
- The proton number is **unique** to each element, so no two elements have the same number of protons
- Electrons can be lost, gained, or shared during chemical processes but the proton number of an atom does not change in a chemical reaction

Mass Number

- The **mass number** (or **nucleon number**) is the total number of protons **and** neutrons in the nucleus of an atom
  - The symbol for this number is ***A***
- The mass number **minus** the proton number gives you the number of **neutrons** of an atom
- Note that protons and neutrons can collectively be called **nucleons**
- The atomic number and mass number for every element is on the periodic table

Atomic Number & Mass Number diagram

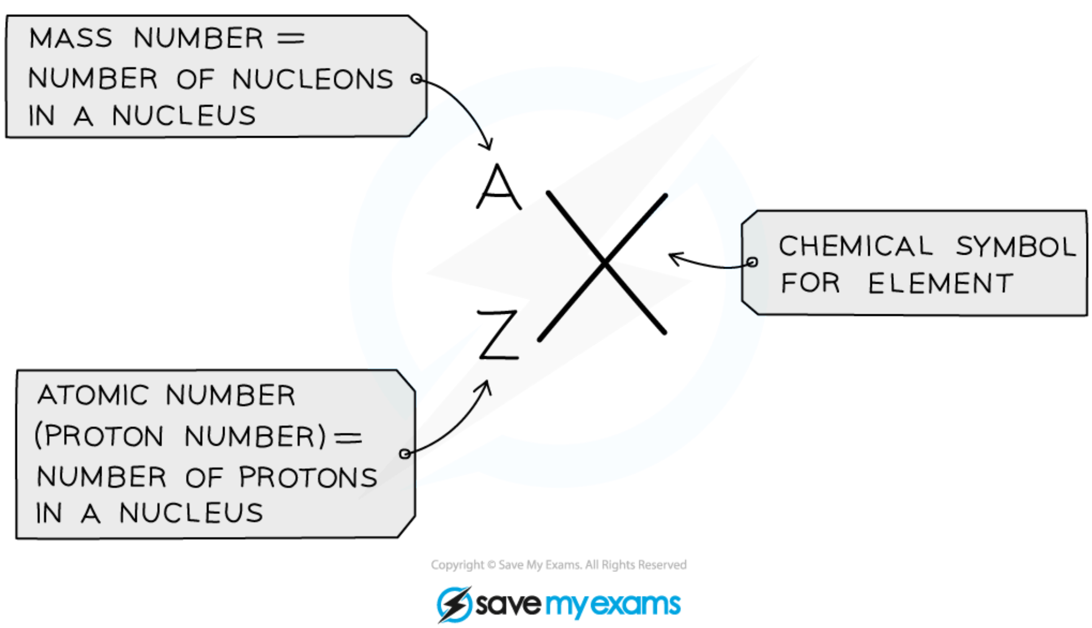


Diagram showing the notation used on the periodic table

Examiner Tip

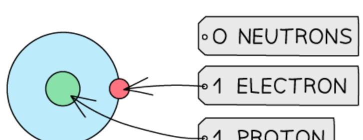
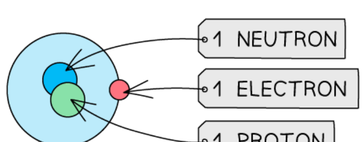
Both the atomic number and the mass number are given on the periodic table, but it can be easy to confuse them.

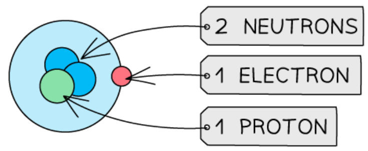
Think **MASS = MASSIVE**, as the mass number is **always** the bigger of the two numbers, the other smaller one is therefore the atomic number

Isotopes

- Isotopes are atoms of the **same element** that contain the same number of **protons** and electrons but a different number of **neutrons**
- The symbol for an isotope is the chemical symbol (or word) followed by a dash and then the mass number
- So, C-14 is the isotope of carbon which contains 6 protons and 6 electrons, but the 14 signifies that it has 8 neutrons (14 - 6 = 8)
  - It can also be written as ^{14}C
- Isotopes display the **same chemical characteristics**
- This is because they have the same number of electrons in their outer shells, and this is what determines their chemistry
- The difference between isotopes is the neutrons which are neutral particles within the nucleus and add mass only

Table to show the structures of isotopes of hydrogen

Isotope	Atomic Structure	Symbol
Hydrogen – 1		$^{1}_{1}H$
Hydrogen – 2		$^{2}_{1}H$

Hydrogen - 3		${}^{3}_{1}H$

Your notes

Examiner Tip

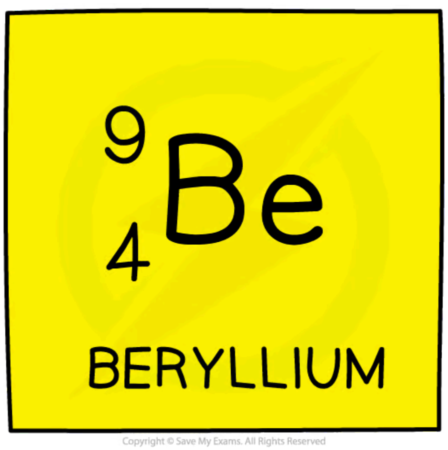
For atoms to be isotopes of each other, they must both be from the same element, hence they must have the same atomic number. E.g. C-13 and C-14 are isotopes whereas C-13 and H-2 are not

Your notes

PEN Numbers

- The **atomic** number is equal to the number of **protons (p)** in an atom
- Since atoms are **neutral**, then it is also the same as the number of **electrons (e)**
- The **mass** number is the number of protons **plus** neutrons
- The number of **neutrons (n)** can thus be calculated by subtracting the atomic number from the mass number
- For example, beryllium has an atomic number of 4, therefore it has 4 protons and 4 electrons
- The mass number of beryllium is 9, so it has 9 - 4 = 5 neutrons
- The PEN numbers for beryllium are thus:
  - p = 4
  - e = 4
  - n = (9 - 4 =) 5

Diagram showing an element from the Periodic Table



The symbol key for beryllium as represented on the Periodic Table

Examiner Tip

The PEN numbers refer to the numbers of protons, electrons, and neutrons in an atom. It is a good study habit to write down the PEN numbers first before answering calculation questions on Atomic Structure

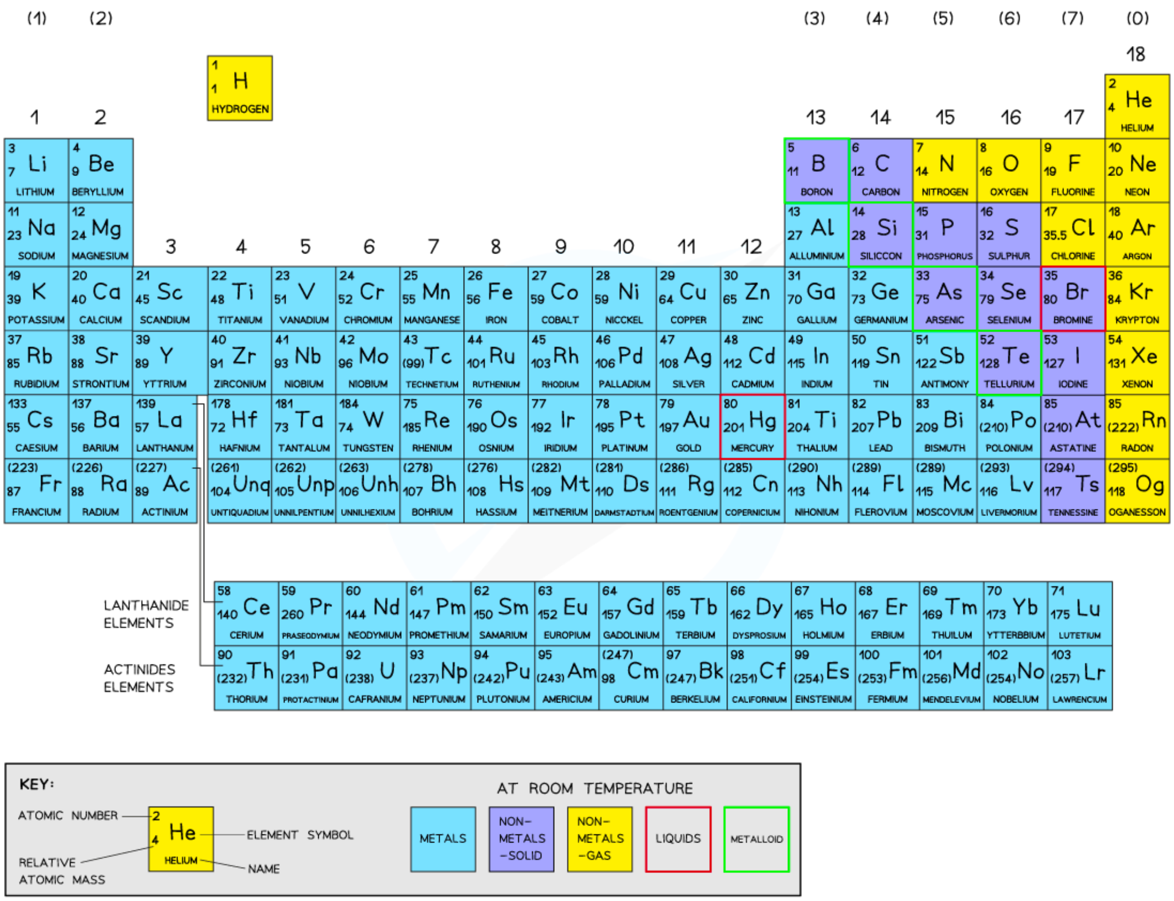
The Periodic Table

The Periodic Table

- 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 is arranged in groups (columns) and periods (rows)

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.

Metals in the Periodic Table

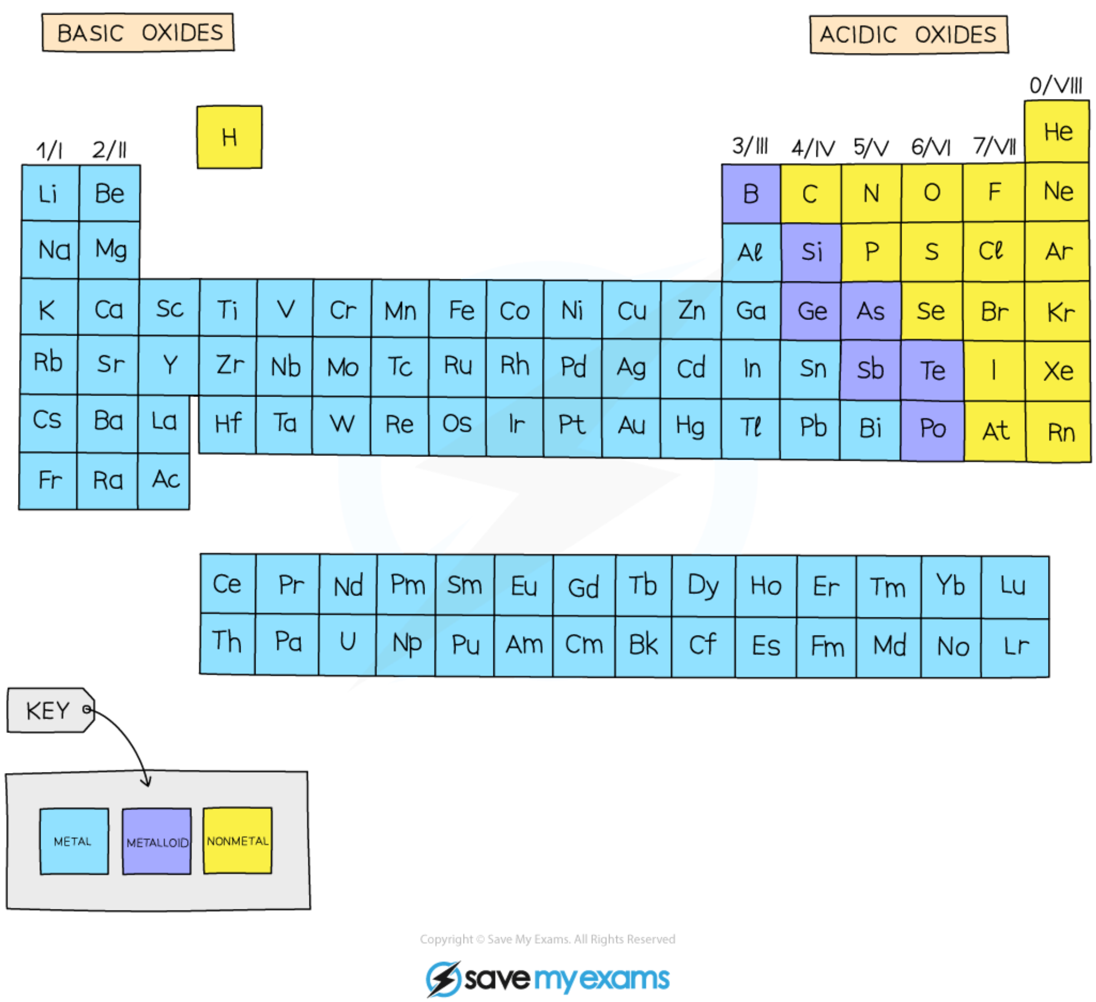
Your notes

- 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
  - Metals are located on the **left** and **centre** of the Periodic Table
- 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
  - Non-metals are located on the **right** hand side of the Periodic Table
- 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

The Periodic Table showing the location of metals and non-metals

Your notes



The metallic character diminishes moving left to right across the Periodic Table

Examiner Tip

In an exam, it is a good idea to draw a 'stair line' on the Periodic Table to separate the metals and non-metals.

This should start above aluminium and continue as if drawing a staircase down the Periodic Table.

This can be seen in the Periodic Table above- the metals are on the left, and the non-metals on the right.

Your notes

Electronic Structure

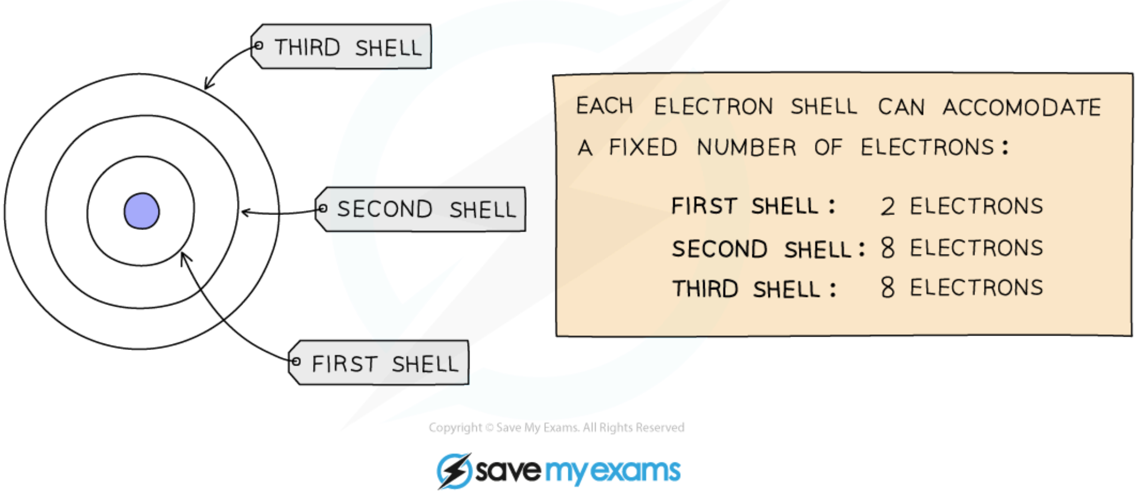
Electronic Structure

Electronic configuration

- We can represent the structure of the atom in two ways: using diagrams called **electron shell diagrams** or by writing out a special notation called the **electronic configuration** (or electronic structure or electron distribution)

Electron shell diagrams

- Electrons orbit the nucleus in **shells** (or **energy levels**) and each shell has a different amount of energy associated with it
- The further away from the nucleus, the more energy a shell has
- Electrons fill the shell closest to the nucleus
- When a shell becomes full of electrons, additional electrons have to be added to the next shell
  - The first shell can hold 2 electrons
  - The second shell can hold 8 electrons
- For this course, a simplified model is used that suggests that the third shell can hold 8 electrons
  - For the first 20 elements, once the third shell has 8 electrons, the fourth shell begins to fill
- The outermost shell of an atom is called the **valence** shell and an atom is much more stable if it can manage to completely fill this shell with electrons



A simplified model showing the electron shells

- The arrangement of electrons in shells can also be explained using numbers
- Instead of drawing electron shell diagrams, the number of electrons in each electron shell can be written down, separated by commas
- This notation is called the electronic configuration (or electronic structure)

Your notes

- E.g. Carbon has 6 electrons, 2 in the 1st shell and 4 in the 2nd shell
  - Its electronic configuration is 2,4
- Electronic configurations can also be written for ions
  - E.g. A sodium atom has 11 electrons, a sodium ion has lost one electron, therefore has 10 electrons; 2 in the first shell and 8 in the 2nd shell
    - Its electronic configuration is 2,8

The Electronic Configuration of the First Twenty Elements

Element	Atomic Number	Electronic Configuration
hydrogen	1	1
helium	2	2
lithium	3	2,1
berylium	4	2,2
boron	5	2,3
carbon	6	2,4
nitrogen	7	2,5
oxygen	8	2,6
fluorine	9	2,7
neon	10	2,8
sodium	11	2,8,1
magnesium	12	2,8,2
aluminium	13	2,8,3
silicon	14	2,8,4

Your notes

phosphorus	15	2,8,5
sulfur	16	2,8,6
chlorine	17	2,8,7
argon	18	2,8,8
potassium	19	2,8,8,1
calcium	20	2,8,8,2

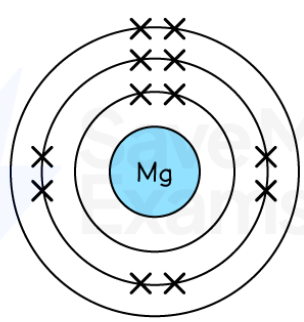
Note: although the third shell can hold up to 18 electrons, the filling of the shells follows a more complicated pattern after potassium and calcium. For these two elements, the third shell holds 8 and the remaining electrons (for reasons of stability) occupy the fourth shell first before filling the third shell.

Worked example

Draw and write the electronic structure of magnesium.

Answer:

- Magnesium has 12 electrons in total.
- A maximum of two can fit in the first shell and eight in the second shell.
- The remaining two will occupy the third shell.



- The written form of this electronic structure is 2,8,2

Your notes

Examiner Tip

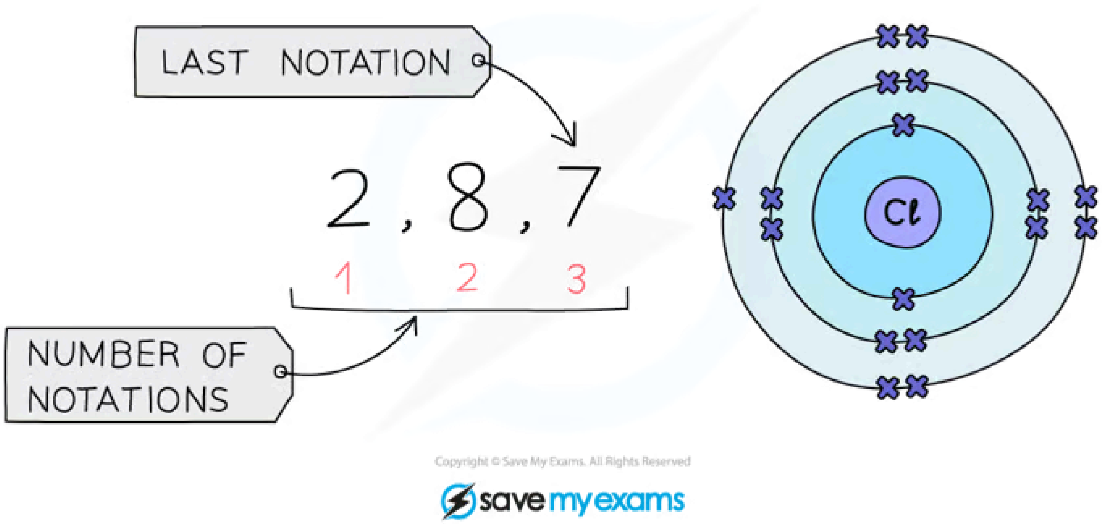
It is a good idea to draw the electrons in their shells in pairs.

You will still score the marks if they aren't, as long as you have the correct number in each shell, but this makes it easier for the examiner to count.

How does the electronic structure of an element relate to its location in the Periodic Table?

- There is a clear relationship between the electronic configuration and how the Periodic Table is designed
- The number of notations in the electronic configuration tells us the number of occupied shells
  - This tells us what **period** an element is in
- The last notation shows the number of outer electrons the atom has
  - This tells us the **group** an element is in
- Elements in the same group have the same number of outer shell electrons

Diagram showing the relationship between the electronic configurations

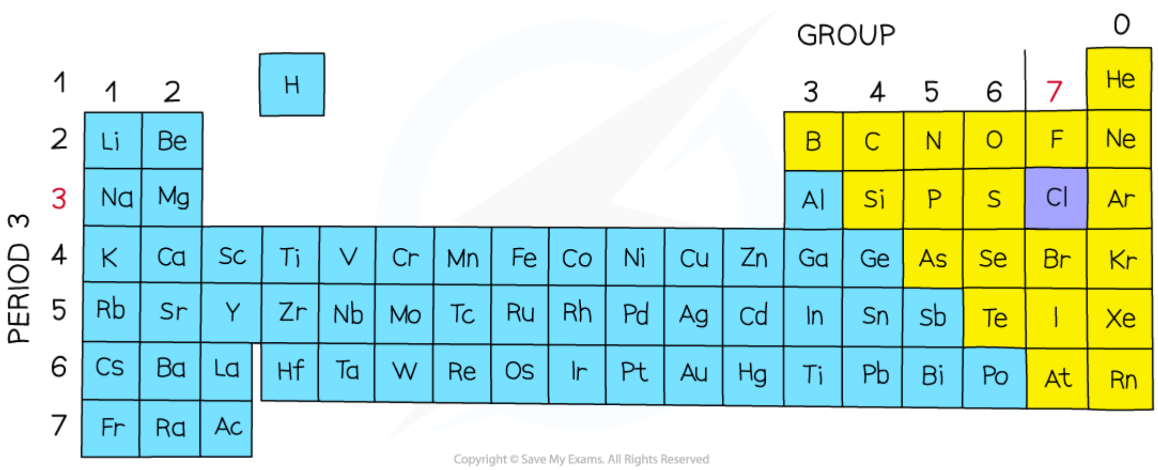


The electronic configuration for chlorine

- **Period:** The red numbers at the bottom show the number of notations
  - The number of notations is 3
  - Therefore chlorine has 3 occupied shells
- **Group:** The last notation, in this case 7
  - This means that chlorine has 7 electrons in its outer shell
  - Chlorine is therefore in Group 7

Your notes

The Periodic Table showing the location of chlorine



Chlorine is in Group 7, Period 3

Examiner Tip

The group number will be labelled on the Periodic Table you are given in your exam, but the period number isn't so it is a good idea to write this on yourself at the beginning.

Trends in the Periodic Table

Your notes

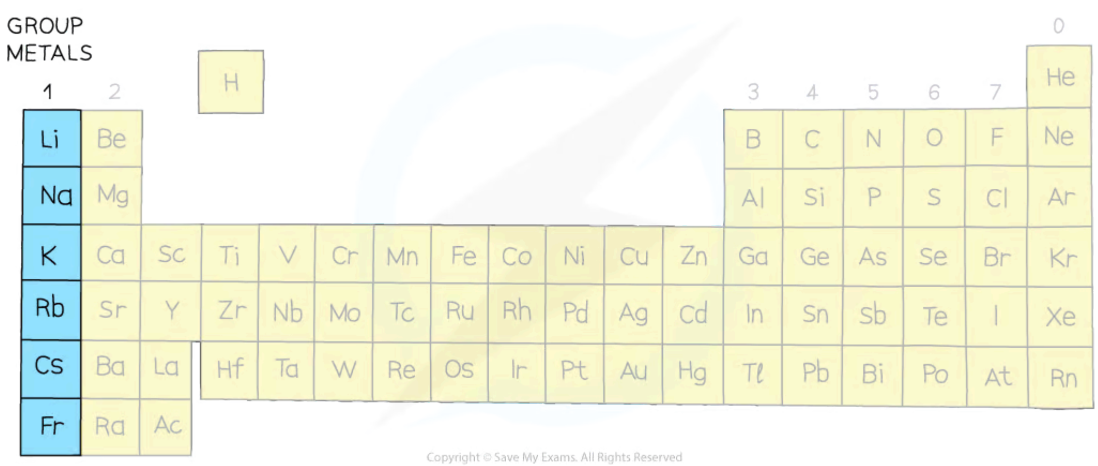
Trends in Groups 1 & 7

- Elements in the same group have similar chemical and physical properties
- We can observe these for Group 1 and Group 7 elements

Group 1 elements

- The Group 1 metals are located in the first column of the Periodic Table and are known as the alkali metals
  - They form **alkaline solutions** when they react with water
- Group 1 metals all share the following properties:
  - 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)
- The alkali metals share similar chemical properties because they each have one electron in their outermost shell

The Periodic Table showing the location of Group 1



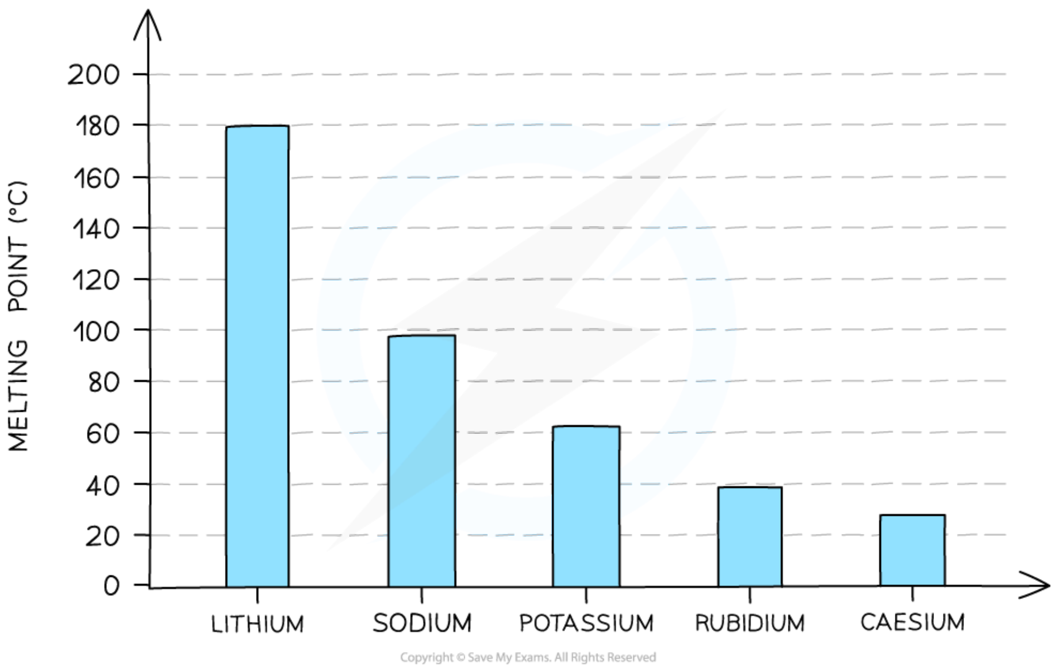
The alkali metals lie on the far left of the Periodic Table, in the very first group

Trends in Group 1 properties

- 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
  - These **decrease** as you move down the group, due to **decreasing attractive forces** between outer electrons and positive ions

Your notes

Diagram to show the trend in melting points going down Group 1



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

Group 7 elements

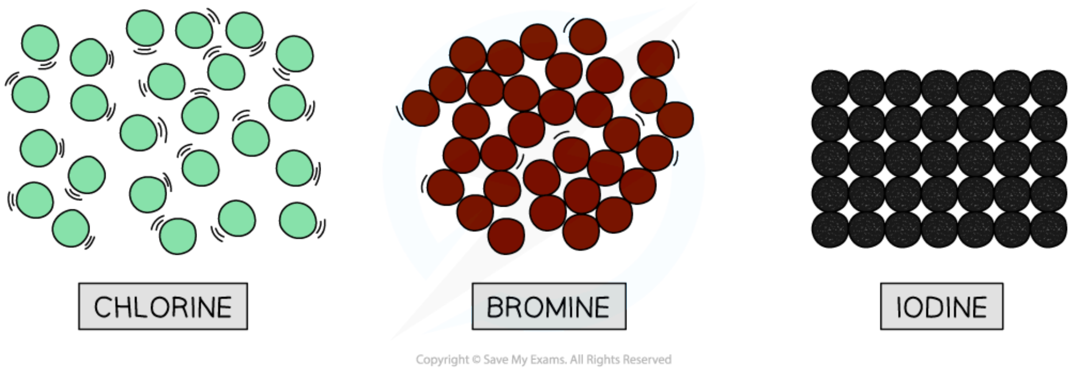
- The elements in Group 7 are known as the halogens
- These are fluorine, chlorine, bromine, iodine and astatine
- These elements are non-metals that are **poisonous**
- Halogens are **diatomic**, meaning they form molecules made of pairs of atoms sharing electrons (forming a single covalent bond between the two halogen atoms) such as F_2, Cl_2, etc

Trends in Group 7 properties

- At room temperature (20 °C), the physical state of the halogens changes as you go down the group
  - Chlorine is a **pale yellow-green gas**, bromine is a **red-brown liquid** and iodine is a **grey-black solid**
  - This demonstrates that the **density** of the halogens **increases** as you go **down the group**
- The colours of the halogens also change as you descend the group - they become darker

Diagram to show the change in state, density and colour of the halogens

Your notes

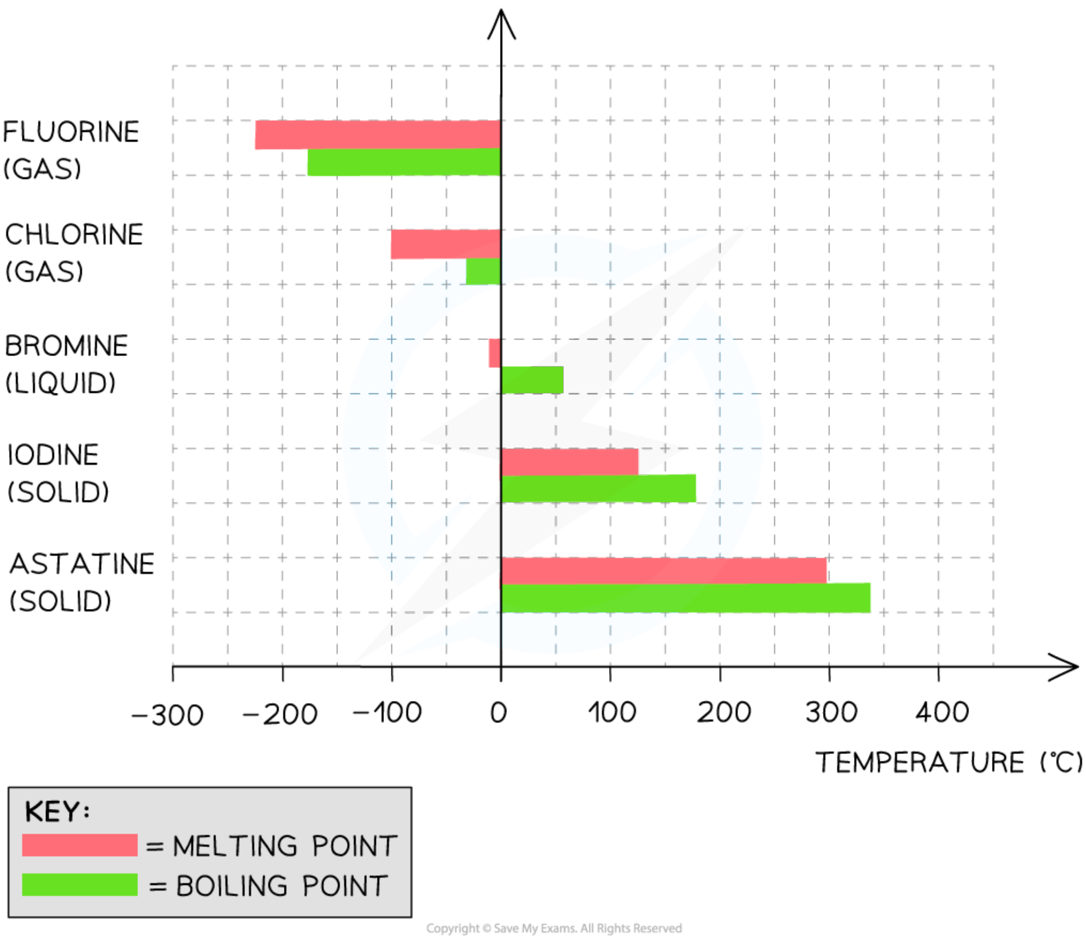


The halogens get denser and change colour moving down the group

- 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

Diagram to show the trend in melting/boiling points going down Group 7

Your notes



Melting and boiling points increase going down Group 7

Examiner Tip

Make sure you learn these properties for each group.

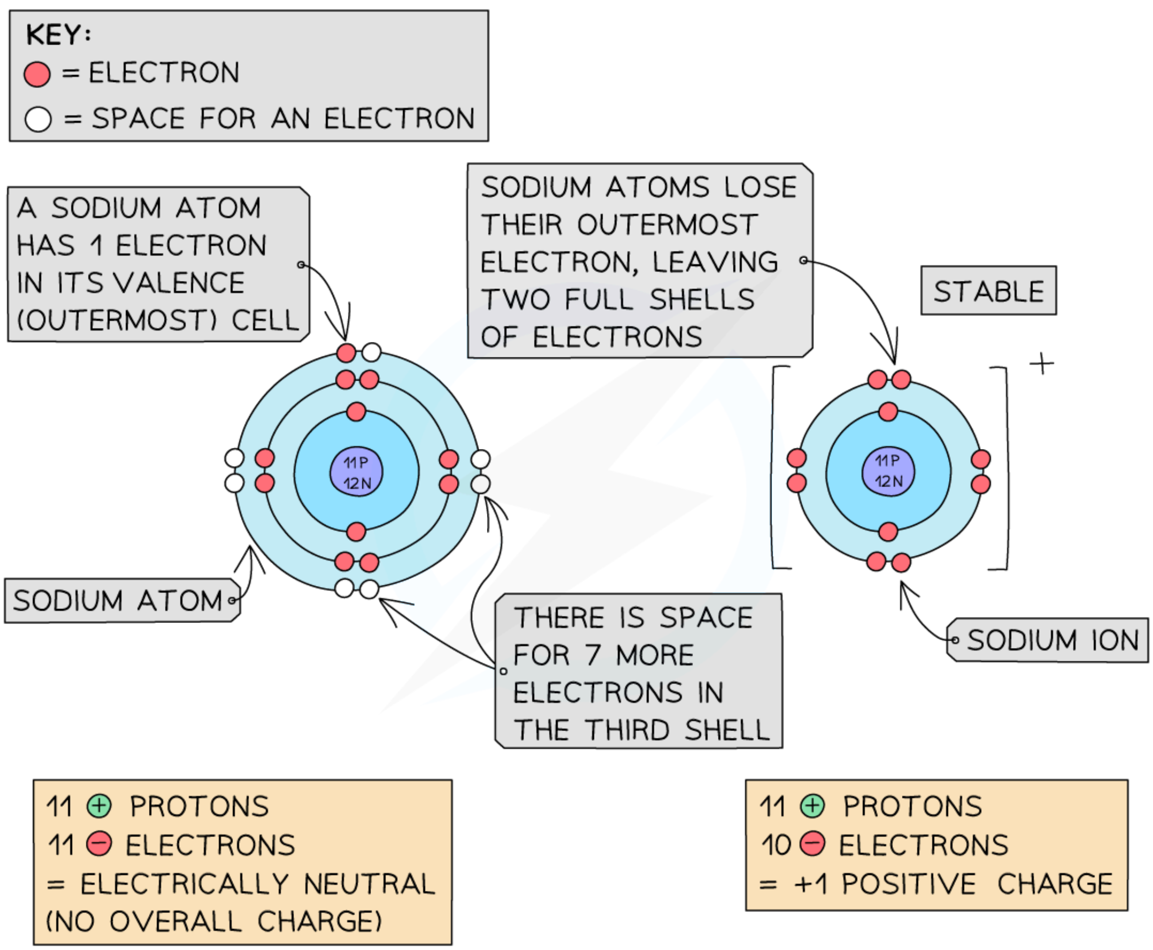
You could be asked to recall these specific properties, or predict the properties of other elements further down the group.

Your notes

Groups 1 & 7 Forming Ions

- Elements react in order to obtain a full outer shell of electrons and become chemically stable
  - To do this, they either lose or gain electrons
- Noble gases are chemically unreactive due to already having a full outer shell of electrons
- Therefore the reactions of Group 1 and Group 7 elements involves the loss or gain of electrons
- Group 1 atoms lose one electron from their outer shell to form positively charged ions
  - `The ions formed will have a 1+ charge

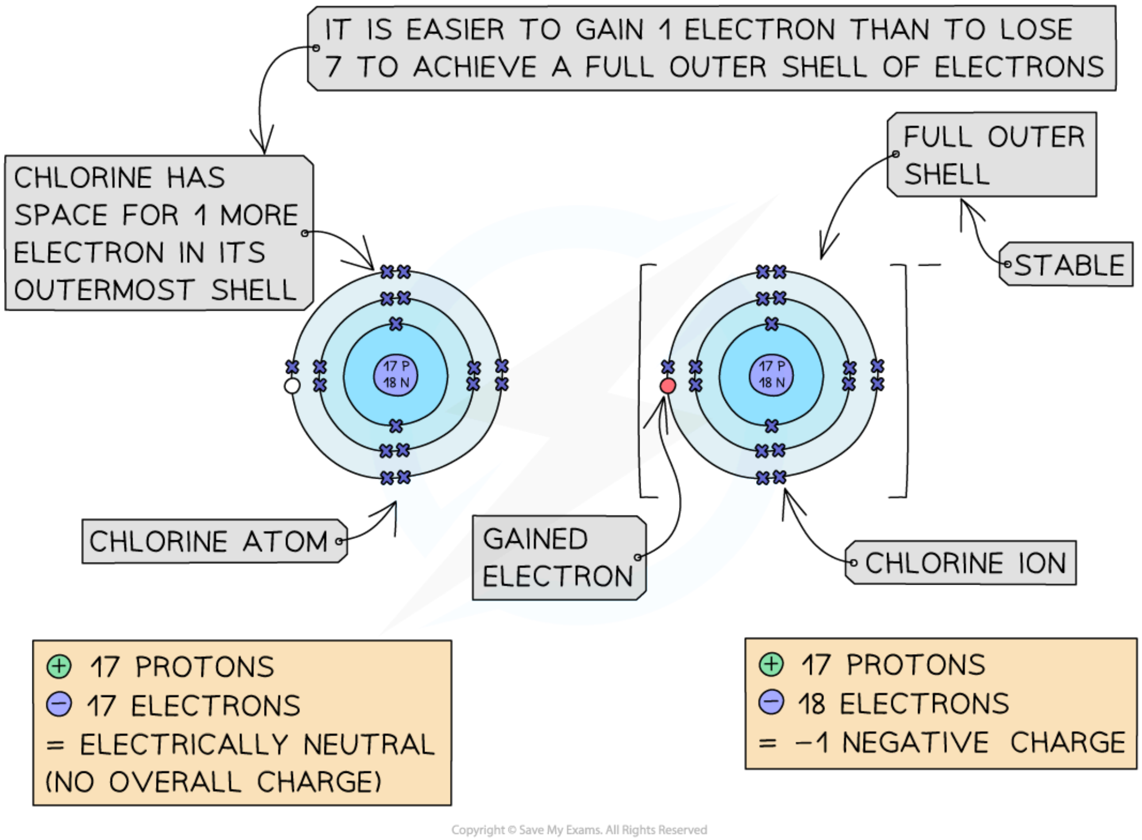
Diagram to show how sodium loses an electron to form an ion



Group 1 elements form ions with a 1+ charge

- Group 7 atoms gain an electron to form negatively charged ions
  - The ions formed will have a 1– charge

Diagram to show how chlorine loses an electron to form an ion



Group 7 elements form ions with a 1– charge

Explaining Trends

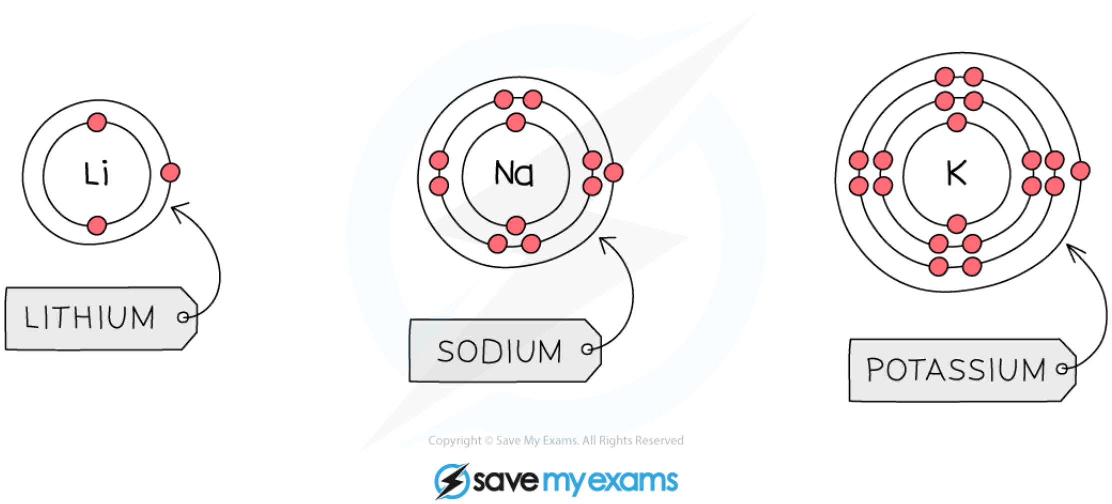
Higher Tier

- The reactivity of an element is determined by its readiness to lose or gain electrons and form ions
- The more readily an atom loses or gains an electron, the more reactive it is
- This explains the trend in reactivity down Groups 1 and 7

Why does reactivity increase going down Group 1?

- When a Group 1 element reacts, its atoms only need to lose one electron, as there is only one 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 ion obtains **noble gas configuration**
- As you go down Group 1 reactivity increases because:
  - The number of electron shells increases
  - 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
  - The outer electron is lost **more easily**

Diagram to show the electronic structure of Group 1 elements



Going down Group 1, the outer electron is further away from the nucleus

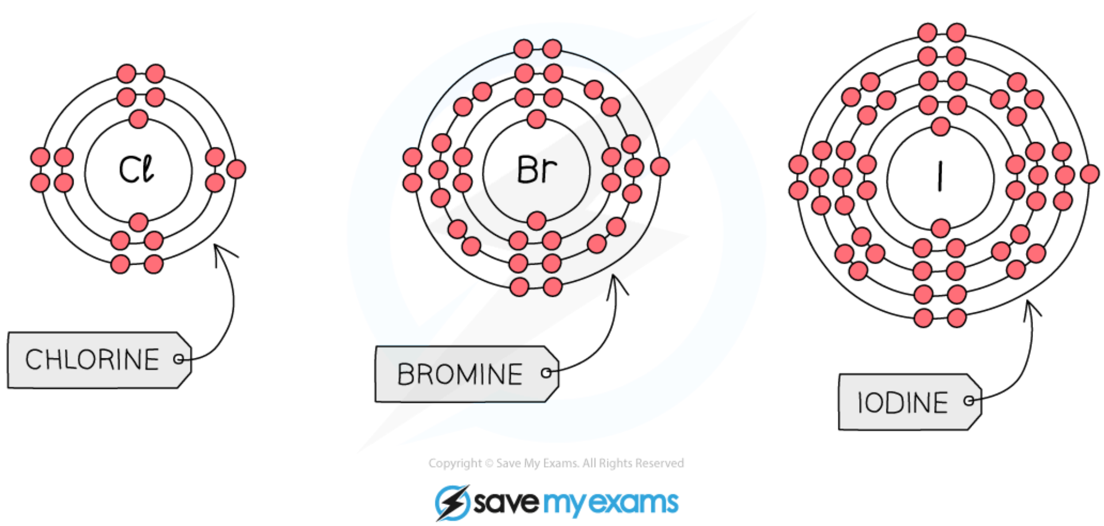
Why does reactivity decrease going down Group 7?

- When a Group 7 element reacts its atoms need to gain one electron as they have seven electrons in their outer shell
  - When this happens, ions with a 1– charge are formed

Your notes

- As you go down Group 7 reactivity decreases because:
  - The number of electron shells increases
  - The **distance** between the electron being gained and the nucleus **increases** so the **attraction** between the electron being gained and the nucleus **decreases**
  - It is **harder** for the atom to **gain** an electron

Diagram to show the electronic structure of Group 1 elements



Going down Group 7, the electron being gained is less attracted to the positive nucleus

Examiner Tip

It would be harder for an atom to lose or gain more than one electron.

This is why Group 2 elements are less reactive than Group 1, and Group 6 elements are less reactive than Group 7.

Your notes

Reactions of Alkali Metals

Reactions of Alkali Metals

- You need to be able to describe the reactions of the first three alkali metals with water, oxygen and the halogens
  - 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

- Alkali metals react with water to form a **metal hydroxide** and **hydrogen** gas
- The reactions of the alkali metals with water get more vigorous as you descend the group, as with the other reactions

Summary table for the reactions of Group 1 metals and 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)
- All Group 1 metals produce **alkaline** solutions (> pH 7) when they react with water, hence the name alkali metals

- Lithium will produce a solution of lithium hydroxide; sodium will produce a solution of sodium hydroxide and so on.
- If universal indicator is added to the solution, it will turn **purple / blue**.

Your notes

Examiner Tip

A common mistake students make when they are asked to give observations of these reactions is that 'hydrogen is formed'.

This is not an observation. Observations would include fizzing / hissing etc.

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
- When placed in a glass jar of oxygen, the metals react vigorously with the oxide forming as white smoke

Summary table for the reactions of Group 1 metals and oxygen

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 table for the reactions of Group 1 metals and halogens

Element	Reaction	Observations

Your notes

Li	lithium + halogen → lithium halide $2Li\ (s) + X_2\ (g) \rightarrow 2LiX\ (s)$ Where X is Cl **or** Br **or** I	▪ A red flame ▪ A white solid is formed
Na	sodium + halogen → sodium halide $2Na\ (s) + X_2\ (g) \rightarrow 2NaX\ (s)$ Where X is Cl **or** Br **or** I	▪ A yellow flame ▪ A white solid is formed
K	potassium + halogen → potassium halide $2K\ (s) + X_2\ (g) \rightarrow 2KX\ (s)$ Where X is Cl **or** Br **or** I	▪ A lilac flame ▪ A white solid is formed

Examiner Tip

For each reaction of Group 1 metals, make sure you can write the word and symbol equations (with state symbols!)

Testing for Hydrogen Gas

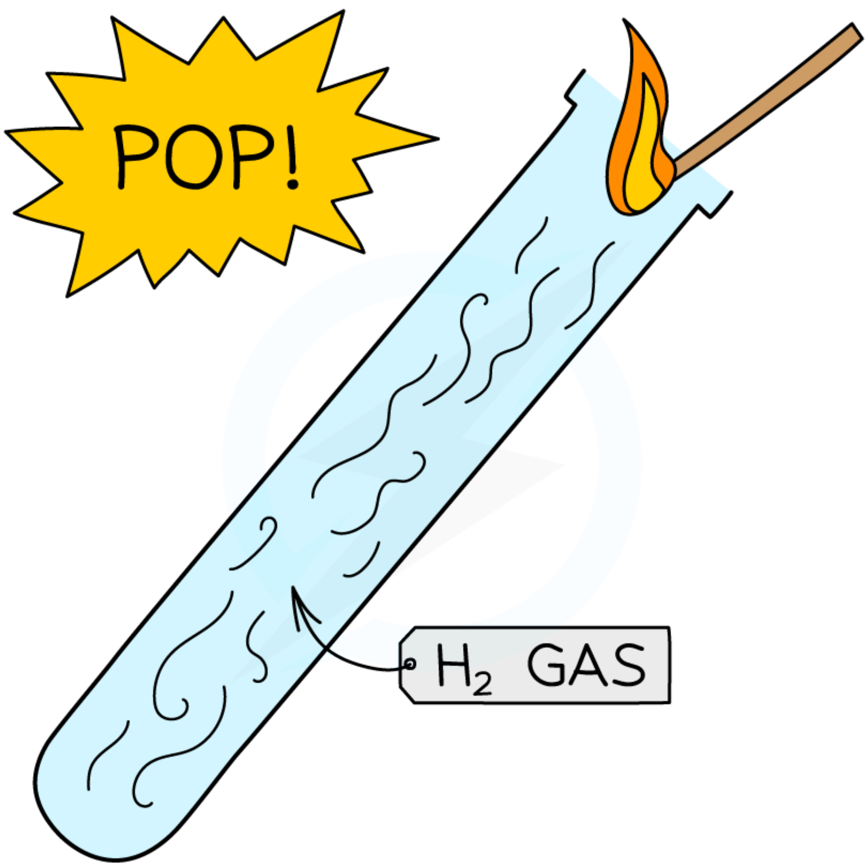
Your notes

Testing for Hydrogen Gas

How to test for hydrogen

- The test for hydrogen consists of holding a **lit splint** held at the open end of a test tube of gas
- If the gas is hydrogen it burns with a loud "**squeaky pop**" which is the result of the rapid combustion of hydrogen with oxygen to produce water
- Be sure not to insert the splint right into the tube, just at the mouth, as the gas needs air to burn

Diagram to show the squeaky pop test



A positive result gives a squeaky pop sound

Your notes

Examiner Tip

It is easy to confuse the tests for hydrogen and oxygen.

Try to remember that a lig**H**ted splint has an **H** for **H**ydrogen, while a gl**O**wing splint has an **O** for **O**xygen.

Your notes

Reactions of Halogens

Reactions of Halogens

The Reaction of Halogens with Alkali Metals

- The halogens react with Group 1 metals (the alkali metals) to form **ionic** compounds which are **metal halide** salts
- Some of these reactions are:
  - Lithium reacts with bromine to form sodium bromide:

lithium + bromide → lithium bromide

$$2Li + Br_2 \rightarrow 2LiBr$$

  - Chlorine will react with sodium to form sodium chloride:

sodium + chlorine → sodium chloride

$$2Na + Cl_2 \rightarrow 2NaCl$$

  - Potassium reacts with iodine to form potassium iodide

potassium + iodine → potassium iodide

$$2K + I_2 \rightarrow 2KI$$

- The word and symbol equations all have the same format, you just need to change the name or the symbol in each case.
- All of the reactions between lithium and halogens produce a **crimson** flame
- All of the reactions between sodium and halogens produce an **orange** flame
- All of the reactions between potassium and halogens produce a **lilac** flame
- The reactions must be carried out in a fume cupboard due to the toxic nature of the halogens
- The halogens **decrease** in reactivity moving down the group so the reactions become less vigorous going down the group

The Reaction of Halogens with Iron

- Group 7 elements react with iron to form an iron halide
- For example, chlorine will react with iron to form iron(III) chloride

iron + chlorine → iron(III) chloride

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

- Each reaction should be carried out in a fume cupboard
- Going down the group, the reaction between the halogen and iron becomes less vigorous

The Reaction of Iron with the Halogens

Halogen	Reaction with iron wool

Fluorine	Cold iron wool burns to produce white iron(III) fluoride
Chlorine	Hot iron wool burns vigorously to form an orange-brown solid of iron(III) chloride
Bromine	Hot iron wool burns vigorously to form an orange-brown solid of iron(III) bromide
Iodine	Hot iron wool reacts slowly with iodine vapour to form a grey solid of iron(III) iodide

Your notes

Worked example

Make sure you can write the word and symbol equations for each reaction.

Your notes

Relative Reactivity of Chlorine, Bromine & Iodine

Higher Tier

- Displacement reactions can show the relative reactivity of the Group 7 elements
- A **halogen displacement reaction** occurs when a more reactive halogen displaces a less reactive halogen from an aqueous solution of its halide
- Aqueous halide solutions are **colourless**
- 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 Bromides & Iodides

- 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 Iodides

- 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

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

Table to show the summary of displacement reactions

	Chlorine (Cl_2)	Bromine (Br_2)	Iodine (I_2)
Potassium chloride (KCl)	x	No reaction	No reaction

Your notes

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

- From this pattern of reaction we can predict that:
  - Fluorine will **displace** all other halogens from their compounds
  - Astatine will **be displaced** by all the halogens from its compounds

- Having said that, astatine is the rarest naturally occurring element so there is not enough around to actually test!
- These displacement reactions provide **stronger evidence** for the decreasing reactivity down Group 7 than that gained from the elements reactions with iron
  - The halogens having different states at room temperature can make it difficult to make a fair comparison when observing their reaction with iron whereas in displacement reactions they are competing directly against one another

Examiner Tip

Displacement reactions are sometimes known as single replacement reactions.

Properties & Uses of Chlorine & Iodine

- Chlorine and iodine are both located in Group 7 of the Periodic Table
- The both exist as diatomic molecules, so are made up of pairs of atoms
- Although both in Group 7, their appearance and state are different:
  - Chlorine is a **green-yellow gas** at room temperature which is toxic
  - Iodine is a **dark grey solid** which can form a purple vapour when it is warmed
- Chlorine is used to **sterilise** drinking water, making it safe for us to drink
- It is also used to kill the bacteria in **swimming pools**
  - The levels of chlorine in swimming pools have to be carefully monitored because chlorine is **toxic**
  - There needs to be enough chlorine present to kill bacteria and sterilise the water without causing any harm to us
- Iodine can be used as an **antiseptic** following hospital procedures, in **plasters** and **sterilising sprays**
  - Iodine is also toxic would levels are carefully monitored

Your notes

Your notes

Photo by Jason Hawke on Unsplash

Chlorine levels need to be carefully monitored in swimming pools

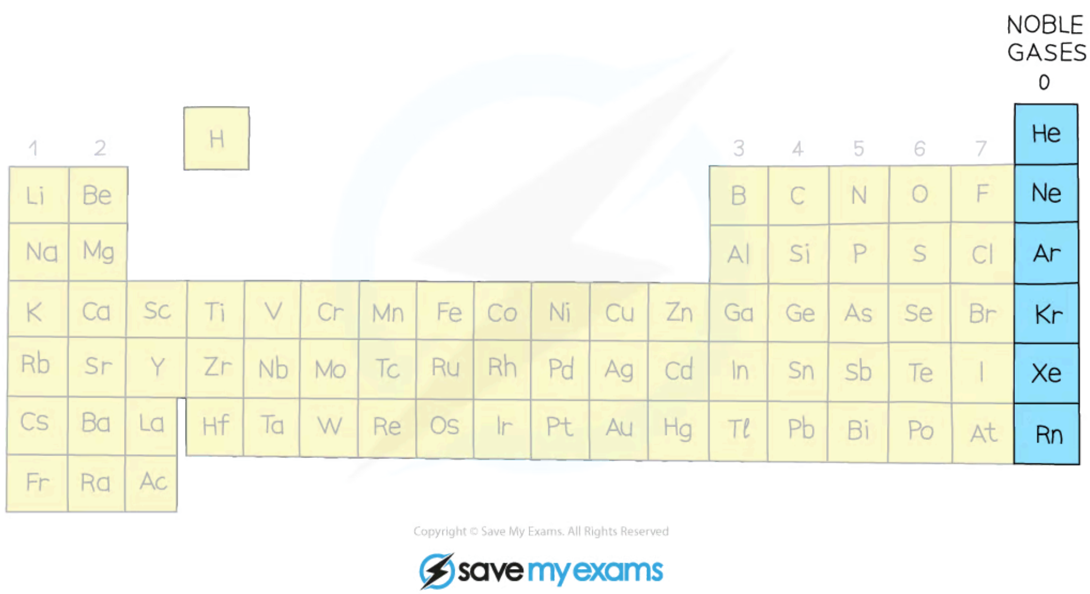
Your notes

Group 0 Gases

Group 0 Gases

- The noble gases are in Group VIII (or Group 0); they are non-metals and have very **low melting** and **boiling** points
- They are all **monoatomic**, **colourless** gases
- The Group 0 elements all have **full outer shells**
- This electronic configuration is **extremely stable** so these elements are unreactive and are inert
- 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

The Periodic Table showing the location of the noble gases



Noble gases are located in Group 0 / 8 of the Periodic Table

- Although they are all unreactive, noble gases have other properties that make them ideal for different uses
- **Helium** has a very low density and is unreactive so is used in airships and weather balloons
- **Neon** emits light when electric current passes through t so is used in advertising signs
- **Argon** is unreactive so is used in light bulbs and as an inert atmosphere for welding

Examiner Tip

Noble gases all have 8 electrons in the outer shell, except for helium which has 2.

Your notes

Testing Cations & Anions

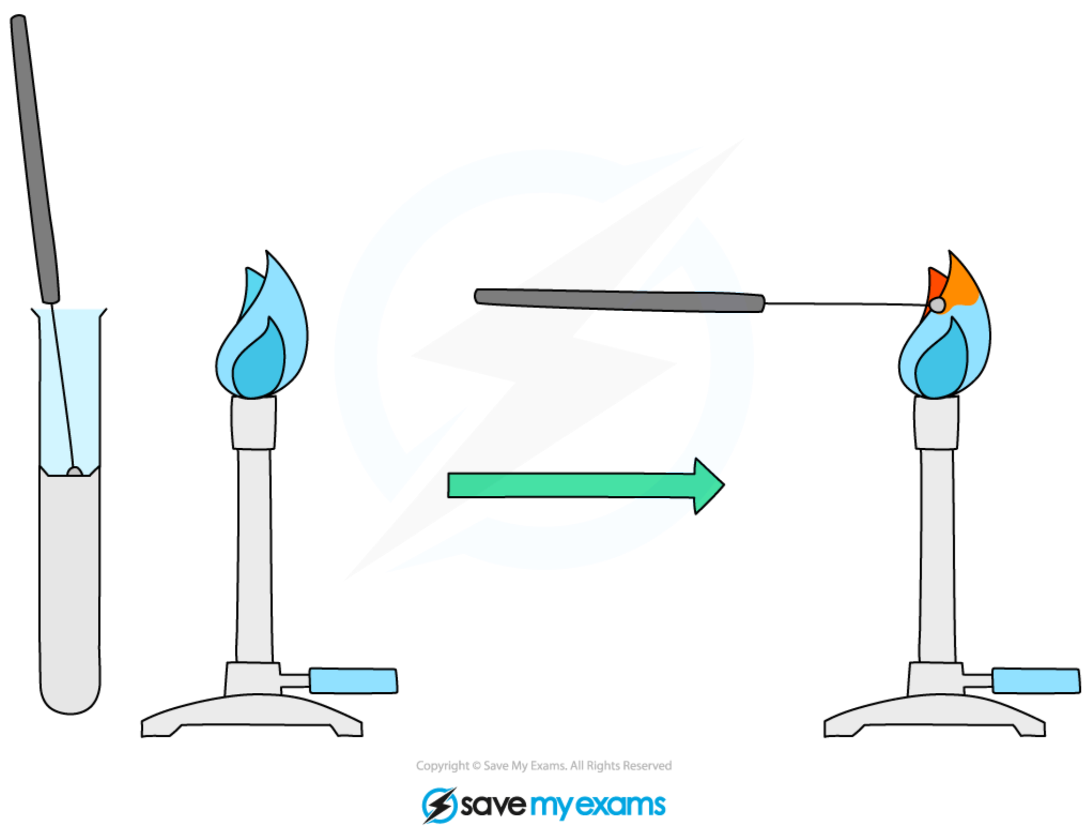
Your notes

Testing Cations & Anions

Testing for Cations

- The **flame test** is used to identify the positive metals ion (cations) by the colour of the flame they produce
  - Ions from different metals produce different colours
- To carry out a flame test:
  - Dip the loop of an unreactive metal wire such as **nichrome** or **platinum** in dilute acid
  - Hold it in the blue flame of a Bunsen burner until there is no colour change
  - Dip the loop into the solid sample / solution and place it in the edge of the **blue** Bunsen flame
- It is important that place the wire into acid first to prevent contamination
  - Not doing this might result in two or more ions being present on the wire meaning the colours will mix
  - One colour could mask another colour and you will not be able to identify the ion

How to carry out a flame test



The blue flame must be used to carry out a flame test

Your notes

- The colour of the flame is observed and used to identify the metal ion present:

Cation	Flame Colour
Li^{+}	Red
Na^{+}	Yellow-orange
K^{+}	Lilac
Ca^{2+}	Brick-red
Ba^{2+}	Apple-green

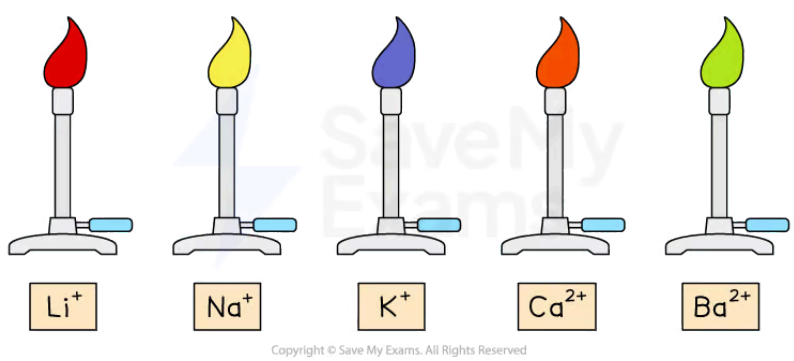


Diagram showing the colours formed in the flame test for metal ions

Examiner Tip

You must be specific when giving the colours.

You will not score the mark for stating that copper produces a 'green' flame. It must be 'apple-green.'

Testing for Anions

- Negatively charged non- metal ions are known as anions
- You must be able to test for **halide** ions
  - These are the ions formed by the elements in Group 7

Your notes

How do I test for halide ions?

- Add **silver nitrate solution,** $AgNO_3$
- If a halide is present it forms a silver halide **precipitate**
- For example, the following reaction occurs between aqueous potassium chloride and and silver nitrate solution:

potassium chloride + silver nitrate → potassium nitrate + silver chloride

$$KCl\,(aq) + AgNO_3\,(aq) \rightarrow KNO_3\,(aq) + AgCl\,(s)$$

- The silver halide formed, in this case silver chloride forms a precipitate
  - This is represented using the state symbol, s
- The ionic equation for the precipitation reaction occurring in this example is:

$$Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$$

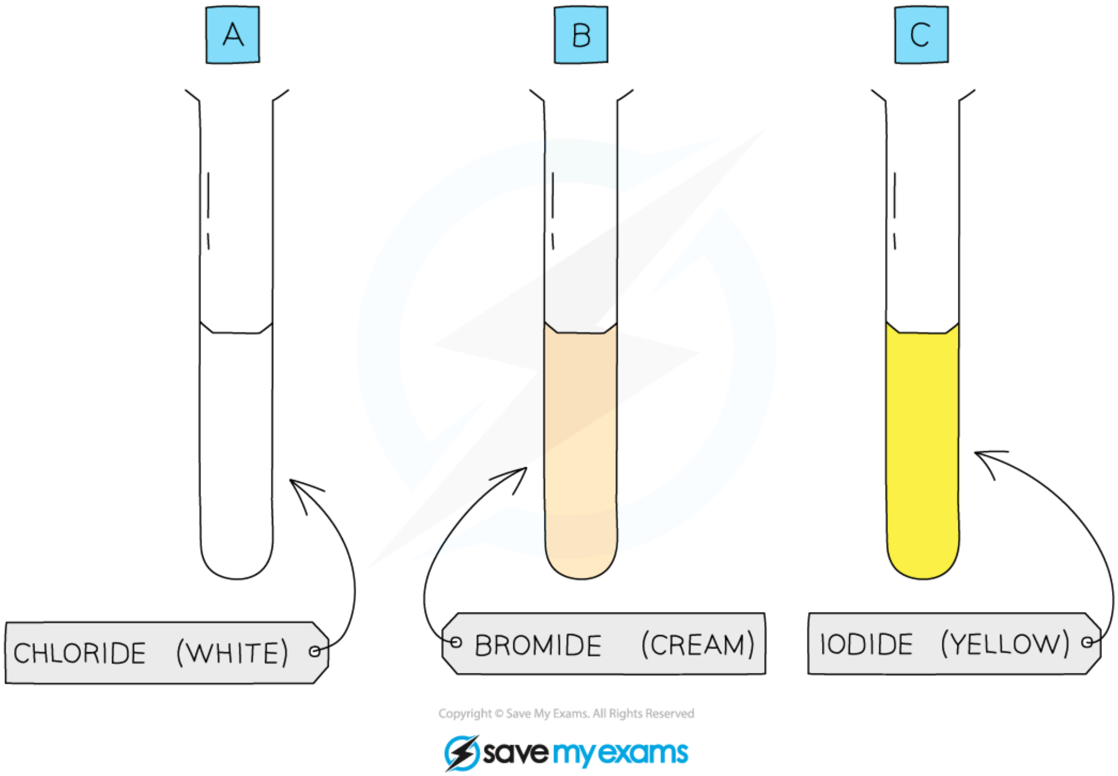
- The **general** ionic equation for the precipitation reaction that occurs when a silver halide is formed is:

$$Ag^+\,(aq) + X^-\,(aq) \rightarrow AgX\,(s)$$

- The potassium and nitrate ions do not participtate in the precipitate reaction so are known as **spectator ions**
- Depending on the halide present, a different coloured **precipitate** is formed, allowing for identification of the halide ion
- Silver chloride forms a **white** precipitate
  - $Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$
- Silver bromide forms a **cream** precipitate
  - $Ag^+\,(aq) + Br^-\,(aq) \rightarrow AgBr\,(s)$
- Silver iodide forms a **yellow** precipitate
  - $Ag^+\,(aq) + I^-\,(aq) \rightarrow AgI\,(s)$

Diagram to show the precipitates formed by different halide ions

Your notes



Each silver halide produces a precipitate of a different colour

Examiner Tip

It is important that you include state symbols when you are writing equations for precipitation reactions but you will be told to include these in your exam.