

A Level Chemistry AQA

2. Inorganic Chemistry

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2.1 Periodicity

2.1.1 Classification of an Element

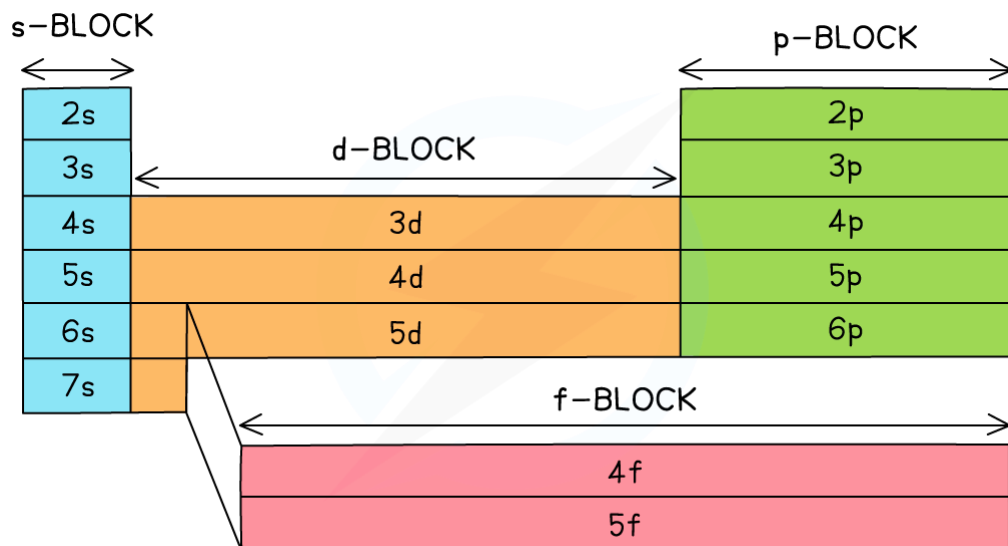
The Periodic Table: Structure & Classification

- The **periodic table** is a list of all known elements arranged in order of **increasing atomic number**, from 1 to 118.
- In addition to this, the elements are arranged in such a way that atoms with the same number of shells are placed together, and atoms with similar electronic configurations in the outer shell are also placed together. This is achieved as follows:
 - The elements are arranged in rows and columns.

-
- Elements with **one** shell are placed in the **first row** (i.e. H and He)
- Elements with **two** shells are placed in the **second row** (Li to Ne) and so on.
- A row of elements thus arranged is called a **period**. The period number, **n**, is the outer energy level that is occupied by electrons.
- In addition, the elements are aligned vertically (in columns) with other elements in different rows, if they share the same outer-shell electronic configuration. The outer electrons are known as the **valence** electrons.
- A column of elements thus arranged is called a **group**

PERIODIC TABLE OF THE ELEMENTS

| | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | | |
|-------|----------------------|--|--|--|--|--|--|--|--|--|--|--|--|--|--|--|-----------------------|--|--------------------|--|--|--|--|--|--|--|--|-------------------|----------------------------|----------------------------|-----------------------------|----------------------------|-------------------------------|-----------------------|-----------------------|-----------------------|----------------------|---------------------|---------------------|
| | ALKALI METALS | | | | | | | | | | | | | | | | ALKALINE EARTH METALS | | 1 H HYDROGEN | | | | | | | | | | | NOBLE GASES 0/VIII | | | | | | | | | |
| 1/I | | | | | | | | | | | | | | | | | | | | | | | | | | | | 4 He HELIUM | | | | | | | | | | | |
| 2/II | 3 Li LITHIUM | | | | | | | | | | | | | | | | 4 Be BERYLLIUM | | | | | | | | | | | | 13/XIII 5 B BORON | 14/XIV 6 C CARBON | 15/XV 7 N NITROGEN | 16/XVI 8 O OXYGEN | 17/XVII 9 F FLUORINE | 20 Ne NEON | | | | | |
| 3/III | 11 Na SODIUM | | | | | | | | | | | | | | | | 12 Mg MAGNESIUM | | TRANSITION METALS | | | | | | | | | | 13 Al ALUMINIUM | 14 Si SILICON | 15 P PHOSPHORUS | 16 S SULPHUR | 17 Cl CHLORINE | 18 Ar ARGON | | | | | |
| 4/IV | 19 K POTASSIUM | | | | | | | | | | | | | | | | 20 Ca CALCIUM | | | | | | | | | | | | 27 Fe IRON | 28 Co COBALT | 29 Ni NICKEL | 30 Cu COPPER | 31 Zn ZINC | 32 Ga GALLIUM | 33 Ge GERMANIUM | 34 As ARSENIC | 35 Se SELENIUM | 36 Br BROMINE | 37 Kr KRYPTON |
| 5/V | 37 Rb RUBIDIUM | | | | | | | | | | | | | | | | 38 Sr STRONTIUM | | | | | | | | | | | | 45 Rh RHODIUM | 46 Pd PALLADIUM | 47 Ag SILVER | 48 Cd CADMIUM | 49 In INDIUM | 50 Sn TIN | 51 Sb ANTIMONY | 52 Te TELLURIUM | 53 I IODINE | 54 Xe XENON | |
| 6/VI | 55 Cs CAESIUM | | | | | | | | | | | | | | | | 56 Ba BARIUM | | | | | | | | | | | | 77 Ir IRIDIUM | 78 Pt PLATINUM | 79 Au GOLD | 80 Hg MERCURY | 81 Tl THALIUM | 82 Pb LEAD | 83 Bi BISMUTH | 84 Po POLONIUM | 85 At ASTATINE | 86 Rn RADON | |
| 7/VII | 87 Fr FRANCIUM | | | | | | | | | | | | | | | | 88 Ra RADIUM | | | | | | | | | | | | 101 Pb LEAD | 102 Bi BISMUTH | 103 Po POLONIUM | 104 At ASTATINE | 105 Rn RADON | | | | | | |
| | | | | | | | | | | | | | | | | | | | | | | | | | | | | | 108 Hg MERCURY | 109 Tl THALIUM | 110 Pb LEAD | 111 Bi BISMUTH | 112 Po POLONIUM | 113 At ASTATINE | 114 Rn RADON | | | | |
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| | | | | | | | | | | | | | | | | | | | | | | | | | | | | | 172 Pb LEAD | 173 Bi BISMUTH | 174 Po POLONIUM | 175 At ASTATINE | 176 Rn RADON | | | | | | |
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The blocks of the periodic table

- All elements belong to one of **four** main blocks: the s-block, the p-block, the d-block and the f-block
 - **The s-block elements are all those with only s electrons in the outer shell**
 - **The p-block elements are all those with at least one p-electron in the outer shell**
 - **The d-block elements are all those with at least one d-electron and at least one s-electron but no f or p electrons in the outer shell (up to 5d)**
 - **The f-block elements are all those with at least one f-electron and at least one s-electron but no d or p electrons in the outer shell**
- The physical and chemical properties of elements in the periodic table show clear patterns related to the position of each element in the table
 - Elements in the same group show similar properties, and properties change gradually as you go across a period
- As atomic number increases, the properties of the elements show trends which repeat themselves in each period of the periodic table
 - These trends are known as **periodic trends** and the study of these trends is known as **periodicity**

Reactions of Period 3 Elements

Reactions with oxygen

- The reactions of period 3 elements with oxygen can be summarised as follows:

| | Chemical Equation | Reaction Conditions | Reaction | Flame | Product |
|----|--|------------------------------|------------|-----------------------|--------------|
| Na | $4\text{Na(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{Na}_2\text{O(s)}$ | Heated | Vigorously | Bright yellow flame | White solid |
| Mg | $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$ | Heated | Vigorously | Bright white flame | White solid |
| Al | $4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)}$ | Powdered Al | Fast | Bright white flame | White powder |
| Si | $\text{Si(s)} + \text{O}_2\text{(g)} \rightarrow \text{SiO}_2\text{(s)}$ | Powdered Si Heat strongly | Slowly | Bright white sparkles | White powder |
| P | $4\text{P(s)} + 5\text{O}_2\text{(g)} \rightarrow \text{P}_4\text{O}_{10}\text{(s)}$ | Heated | Vigorously | Yellow or white flame | White clouds |
| S | $\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)}$ | Powdered S is heated | Gently | Blue flame | Toxic fumes |

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Reactions with chlorine

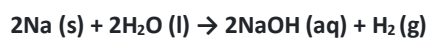
- The reactions of period 3 elements with chlorine can be summarised as follows:

| | Chemical Equation | Reaction Conditions | Reaction |
|----|---|---------------------------|------------|
| Na | $2\text{Na(s)} + \text{Cl}_2\text{(g)} \rightarrow 2\text{NaCl(s)}$ | Heated | Vigorously |
| Mg | $\text{Mg(s)} + \text{Cl}_2\text{(g)} \rightarrow \text{MgCl}_2\text{(s)}$ | Heated | Vigorously |
| Al | $2\text{Al(s)} + 3\text{Cl}_2\text{(g)} \rightarrow \text{Al}_2\text{Cl}_6\text{(s)}$ | Heated | Vigorously |
| Si | $\text{Si(s)} + 2\text{Cl}_2\text{(g)} \rightarrow \text{SiCl}_4\text{(l)}$ | Heated | Slowly |
| P | $2\text{P(s)} + 5\text{Cl}_2\text{(g)} \rightarrow 2\text{PCl}_5\text{(s)}$ | Heated Excess chlorine | Slowly |

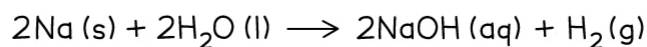
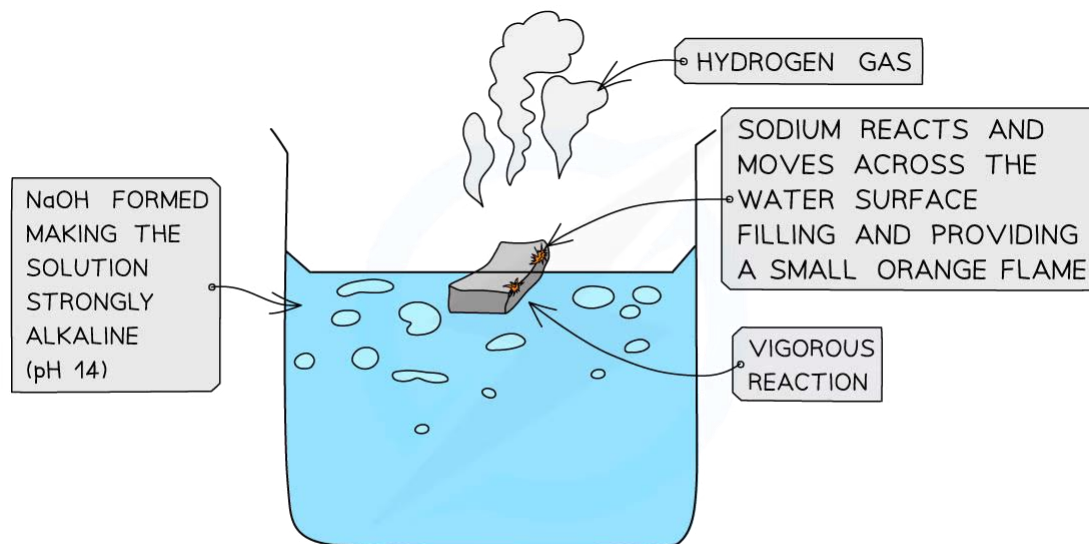
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Reaction of sodium & magnesium with water

- Sodium reacts **vigorously** with **cold water**:



- The sodium **melts** into a ball and moves across the water surface until it disappears
- **Hydrogen gas** is given off
- The solution formed is **strongly alkaline** (pH 14) due to the **sodium hydroxide** which is formed



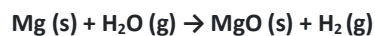
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The diagram shows the reaction of sodium with cold water

- Magnesium reacts **extremely slowly** with cold water:



- The solution formed is **weakly alkaline** (pH 9-10) as **magnesium hydroxide** is only **slightly soluble**
- However, when magnesium is **heated in steam**, it reacts **vigorously** with steam to make magnesium oxide and hydrogen gas:

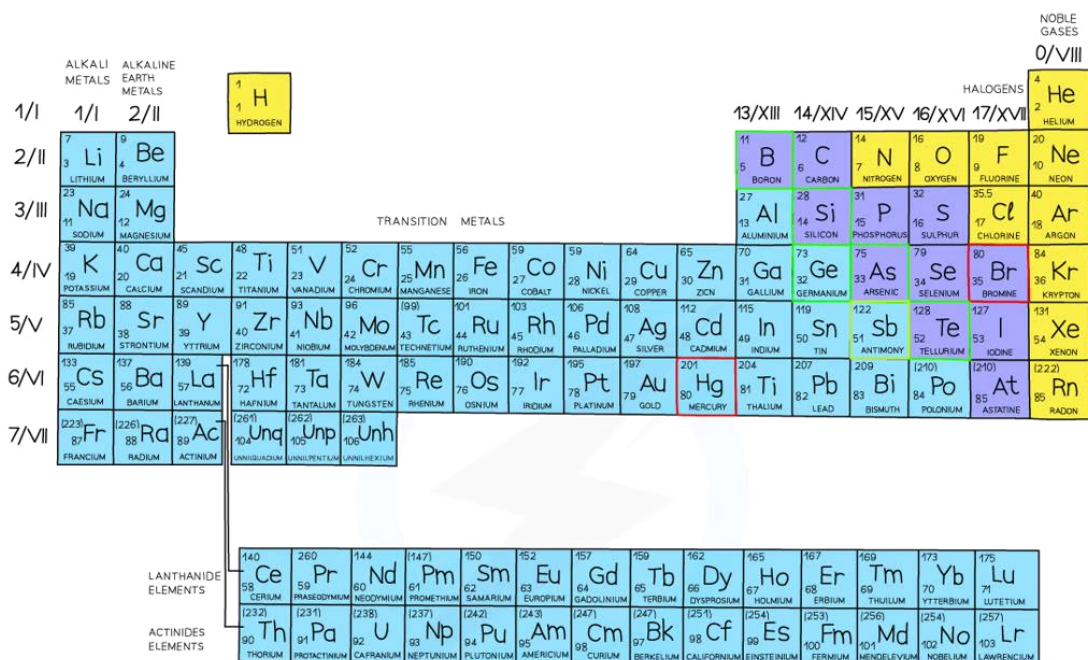


2.1.2 Trends of Period 3 Elements: Atomic Radius

Trend: Atomic Radius

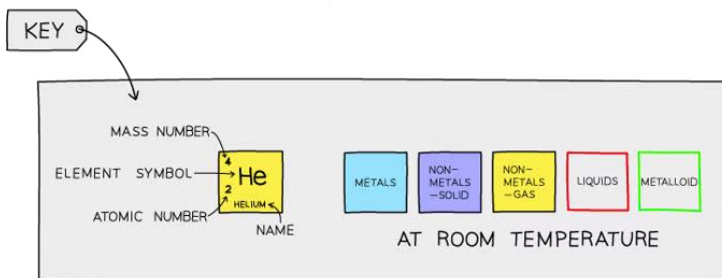
- Elements in the periodic table are arranged in order of increasing atomic number and placed in vertical columns (**groups**) and horizontal rows (**periods**)
- The elements across the periods show **repeating patterns** in chemical and physical properties
- This is called **periodicity**

PERIODIC TABLE OF THE ELEMENTS



The periodic table shows elements arranged by increasing atomic number. Key features include:

- Alkali Metals:** Group 1 (Li, Na, K, Rb, Cs, Fr)
- Alkaline Earth Metals:** Group 2 (Be, Mg, Ca, Sr, Ba, Ra)
- Transition Metals:** Groups 3-10 (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr)
- Halogens:** Group 17 (F, Cl, Br, I, At)
- Noble Gases:** Group 18 (He, Ne, Ar, Kr, Xe, Rn)
- Lanthanide Elements:** Period 6, elements 57-71 (Ce to Lu)
- Actinide Elements:** Period 7, elements 89-103 (Th to Lr)



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All elements are arranged in the order of increasing atomic number from left to right

Atomic radius

- The **atomic radius** is the distance between the nucleus and the outermost electron of an atom

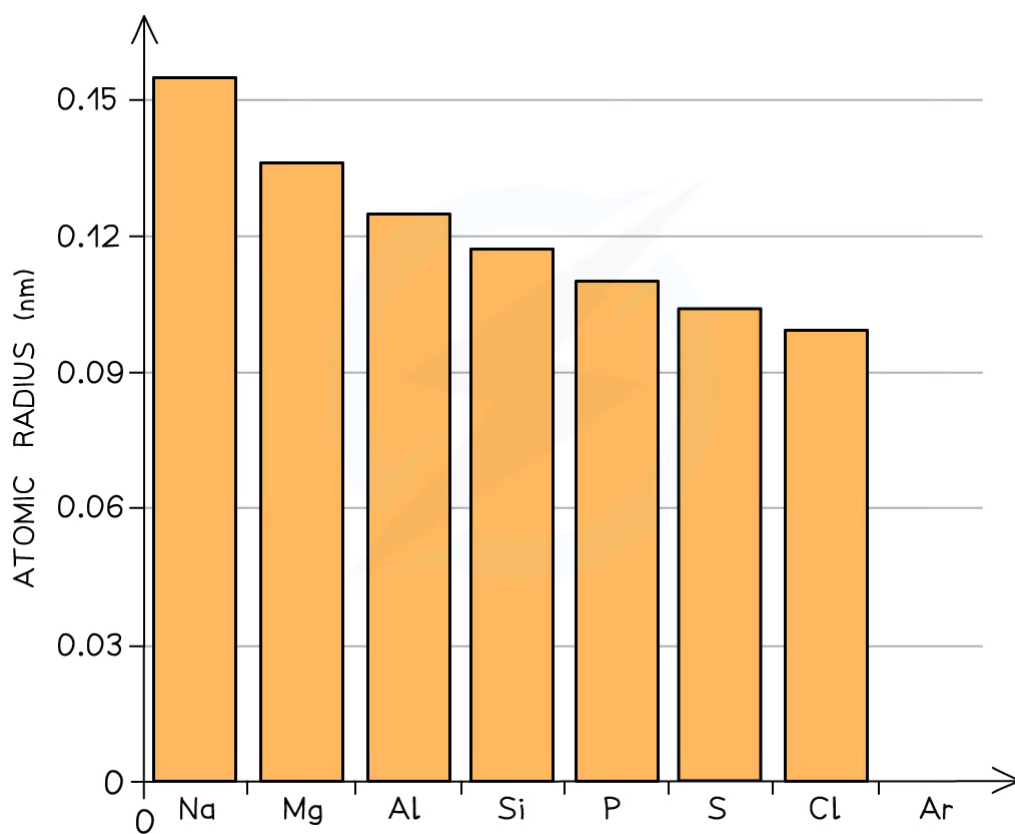
- The atomic radius is measured by taking two atoms of the same element, measuring the distance between their nuclei and then halving this distance In **metals** this is also called the **metallic radius** and in **non-metals**, the **covalent radius**

Atomic radii of period 3 elements

- You can see a clear trend across the period:

| Period 3 Element | Na | Mg | Al | Si | P | S | Cl | Ar |
|--------------------|-------|-------|-------|-------|-------|-------|-------|---------|
| Atomic Radius (nm) | 0.157 | 0.136 | 0.125 | 0.117 | 0.110 | 0.104 | 0.099 | No data |

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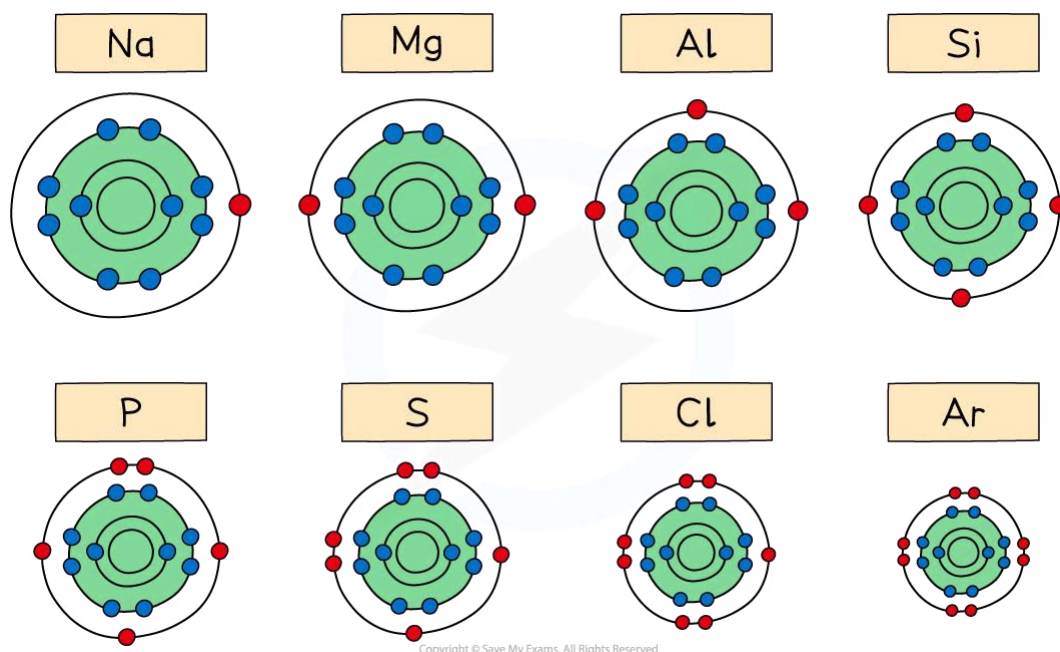


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The graph shows a decrease in atomic radii of period 3 elements across the period

- Across the period, the atomic radii decrease
- This is because the number of protons (**the nuclear charge**) and the number of **electrons** increases by one every time you go an element to the right

- The elements in a period all have the same number of shells (so the **shielding effect** is the same)
- This means that as you go across the period the nucleus attracts the electrons more **strongly pulling them closer** to the nucleus
- Because of this, the atomic radius (and thus the size of the atoms) **decreases** across the period



The diagram shows that across period 3, the elements gain extra electrons in the same principal quantum shell

Ionic radius

- The **ionic radius** is the distance between the nucleus and the outermost electron of an ion
- Metals produce positively charged ions (**cations**) whereas nonmetals produce negatively charged ions (**anions**)
- The **cations** have lost their valence electrons which causes them to be much **smaller** than their parent atoms
- Because there are fewer electrons, this also means that there is less **shielding** of the outer electrons
- Going across the period from Na^+ to Si^{4+} the ions get **smaller** due to the **increasing nuclear charge** attracting the outer electrons in the **second principal quantum shell** nucleus (which has an increasing atomic number)
- The **anions** are larger than their original parent atoms because each atom has gained one or more electrons in their **third principal quantum shell**
- This increases the repulsion between electrons, while the **nuclear charge** is still the same, causing the electron cloud to spread out
- Going across P^{3-} to Cl^- the ionic radii **decreases** as the nuclear charge increases across the period and fewer electrons are gained by the atoms (P gains 3 electrons,

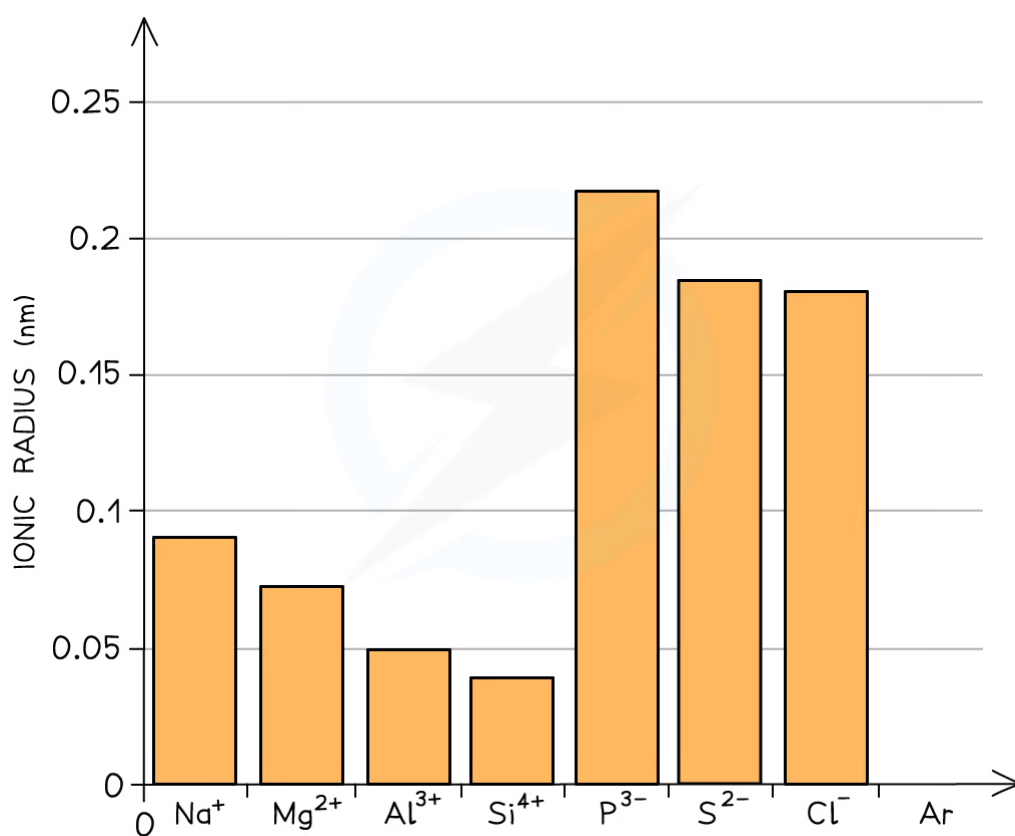
S 2 electrons and Cl 1 electron)

Ionic radii of ions of period 3 elements

- The pattern in ionic radius is less clear cut:

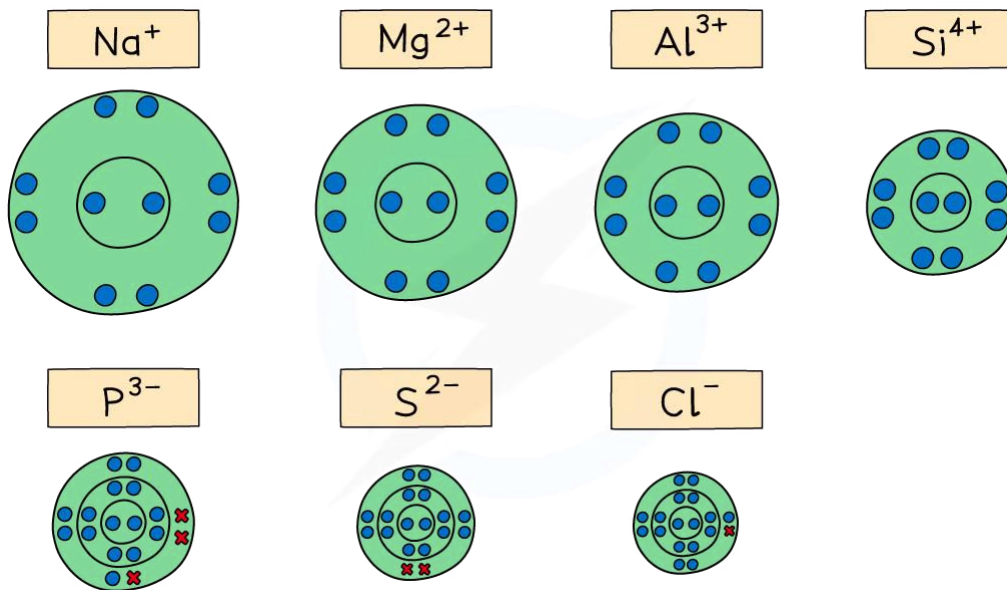
| Ions of Period 3 Elements | Na ⁺ | Mg ²⁺ | Al ³⁺ | Si ⁴⁺ | P ³⁻ | S ²⁻ | Cl ⁻ | Ar |
|---------------------------|-----------------|------------------|------------------|------------------|-----------------|-----------------|-----------------|---------|
| Ionic Radius (nm) | 0.095 | 0.065 | 0.050 | 0.041 | 0.212 | 0.184 | 0.181 | No data |

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Ions of period 3 elements with increasing positive charge (metals) and increasing of outer electrons across the period



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The diagram shows the electron configuration of the ions of period 3 elements and their relative sizes

2.1.3 Trends of Period 3 Elements: First Ionisation Energy

Trend First Ionisation Energy

- The **first ionisation energy** (IE_1) is the energy required to **remove** one mole of electrons from one mole of atoms of an element in the **gaseous state** to form one mole of **gaseous ions**

◦ Eg. the first ionisation energy of Na is:



IE_1 Values of the Period 3 Elements

| Period 3 Element | IE_1 (kJ mol^{-1}) |
|------------------|---------------------------------|
| Na | 494 |
| Mg | 736 |
| Al | 577 |
| Si | 786 |
| P | 1060 |
| S | 1000 |
| Cl | 1260 |
| Ar | 1520 |

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- There is a general **increase** in IE_1 across a period
 - The nuclear charge **increases**
 - The atomic radius **decreases**

- There are **stronger** attractive forces between the nucleus and outer electrons
- It therefore gets harder to remove any electrons
- Small 'dips' are observed between Mg - Al and P - S

2.1.4 Trends of Period 3 Elements: Melting Point

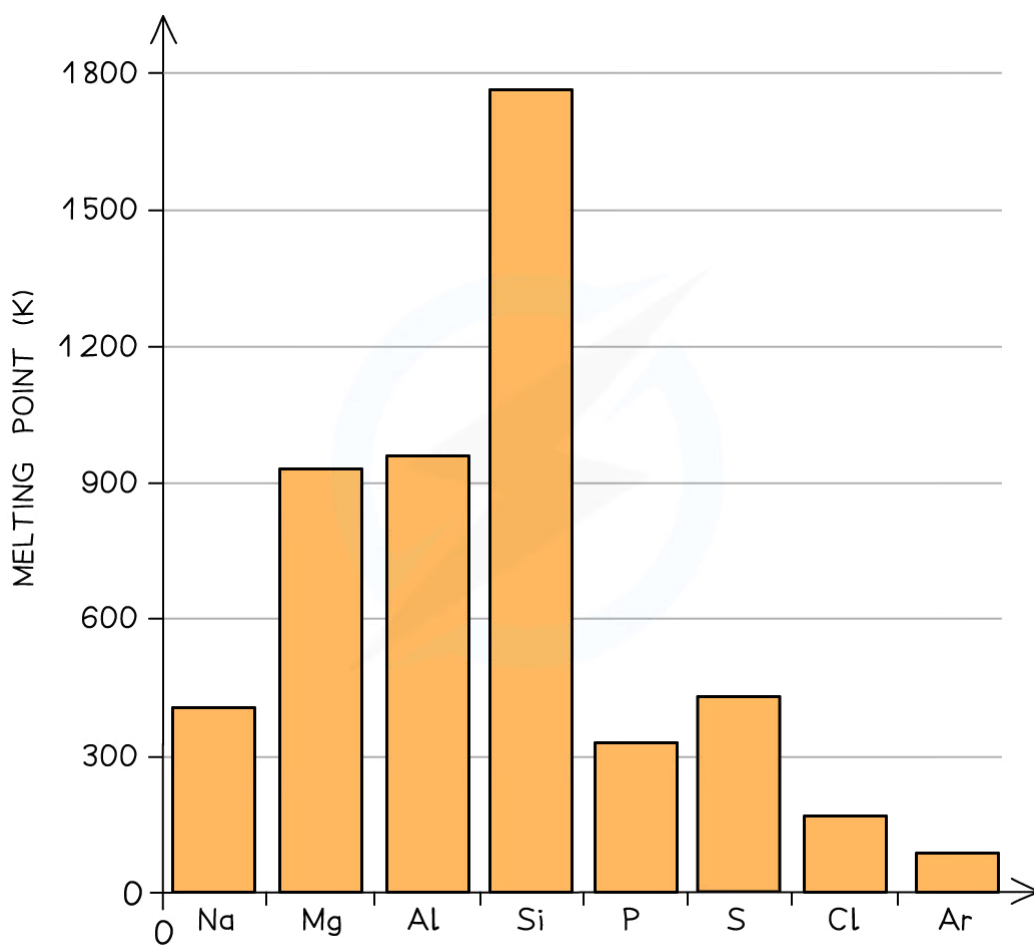
Trend: Melting Point

Melting points of the elements across period 3

falls:

| Period 3 Element | Na | Mg | Al | Si | P | S | Cl | Ar |
|-------------------|-----|-----|-----|------|-----|-----|-----|----|
| Melting Point (K) | 371 | 923 | 932 | 1683 | 317 | 392 | 172 | 84 |

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A pattern is a little harder to see from the data, but you can see that it rises and

Melting points of the period 3 elements

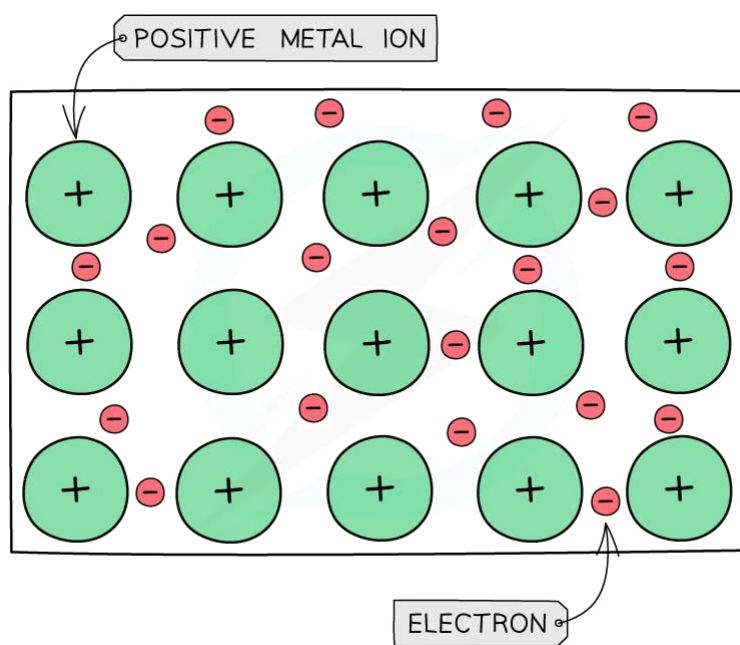
- The trends in melting point can be explained by looking at the bonding and structure of the elements

Bonding & Structure of the Period 3 Elements

| Period 3 Element | Na | Mg | Al | Si | P | S | Cl | Ar |
|------------------|----------------|----------------|----------------|-----------------|------------------|------------------|------------------|------------------|
| Bonding | Metallic | Metallic | Metallic | Covalent | Covalent | Covalent | Covalent | – |
| Structure | Giant metallic | Giant metallic | Giant metallic | Giant molecular | Simple molecular | Simple molecular | Simple molecular | Simple molecular |

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- The table shows that **Na**, **Mg** and **Al** are metallic elements which form positive ions arranged in a **giant lattice** in which the ions are held together by a 'sea' of delocalised electrons around them.



Metal cations form a giant lattice held together by electrons that can move around freely

- The electrons in the 'sea' of delocalised electrons are those from the **valence shell** of the atoms
- Na** will donate one electron into the 'sea' of delocalised electrons, **Mg** will donate two and **Al** three electrons
- As a result of this, the metallic bonding in **Al** is stronger than in **Na**
This is because the electrostatic forces between a **3+ ion** and the larger number of negatively charged delocalised electrons is much larger compared to a **1+ ion** and the smaller number of delocalised electrons in Na
- Because of this, the **melting points increase** going from **Na** to **Al**
- Si** has the highest melting point due to its giant molecular structure in which each

Si atom is held to its neighboring Si atoms by four **strong covalent bonds**

- **P, S, Cl** and **Ar** are non-metallic elements and exist as **simple molecules** (P_4 , S_8 , Cl_2 and Ar as a single atom)
- The **covalent bonds within** the molecules are strong, however, **between** the molecules, there are only weak **instantaneous dipole-induced dipole forces**
- It doesn't take much energy to break these **intermolecular** forces
- Therefore, the melting points decrease going from **P** to **Ar** (note that the melting point of S is higher than that of P as sulphur exists as larger S_8 molecules compared to the smaller P_4 molecule)

Electrical conductivity

- The **electrical conductivity decreases** going across the Period 3 elements

Electrical Conductivity across Period 3 Elements

| Period 3 Element | Na | Mg | Al | Si | P | S | Cl | Ar |
|---|-------|-------|-------|---------------------|------------|------------|----|----|
| Electrical Conductivity ($S\ m^{-1}$) | 0.218 | 0.224 | 0.382 | 2×10^{-10} | 10^{-17} | 10^{-23} | - | - |

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- Going from **Na** to **Al**, there is an increase in the number of valence electrons that are donated to the 'sea' of delocalised electrons
- Because of this, in **Al** there are more electrons available to move around through the structure when it conducts electricity, making **Al** a better electrical conductor than **Na**
- Due to the **giant molecular structure** of **Si**, there are no delocalised electrons that can freely move around within the structure
- **Si** is therefore not a good electrical conductor and is classified as a **semimetal (metalloid)**
- The lack of delocalised electrons is also why **P** and **S** cannot conduct electricity

Exam Tip

The SI unit of electrical conductivity is siemens per metre named after Ernest von Siemens. Previously, the unit used to be known as the mho, where not only is 1 mho the reciprocal of 1 ohm (the unit of resistance), it is also ohm spelled backwards!

2.2 Group 2, the Alkaline Earth Metals

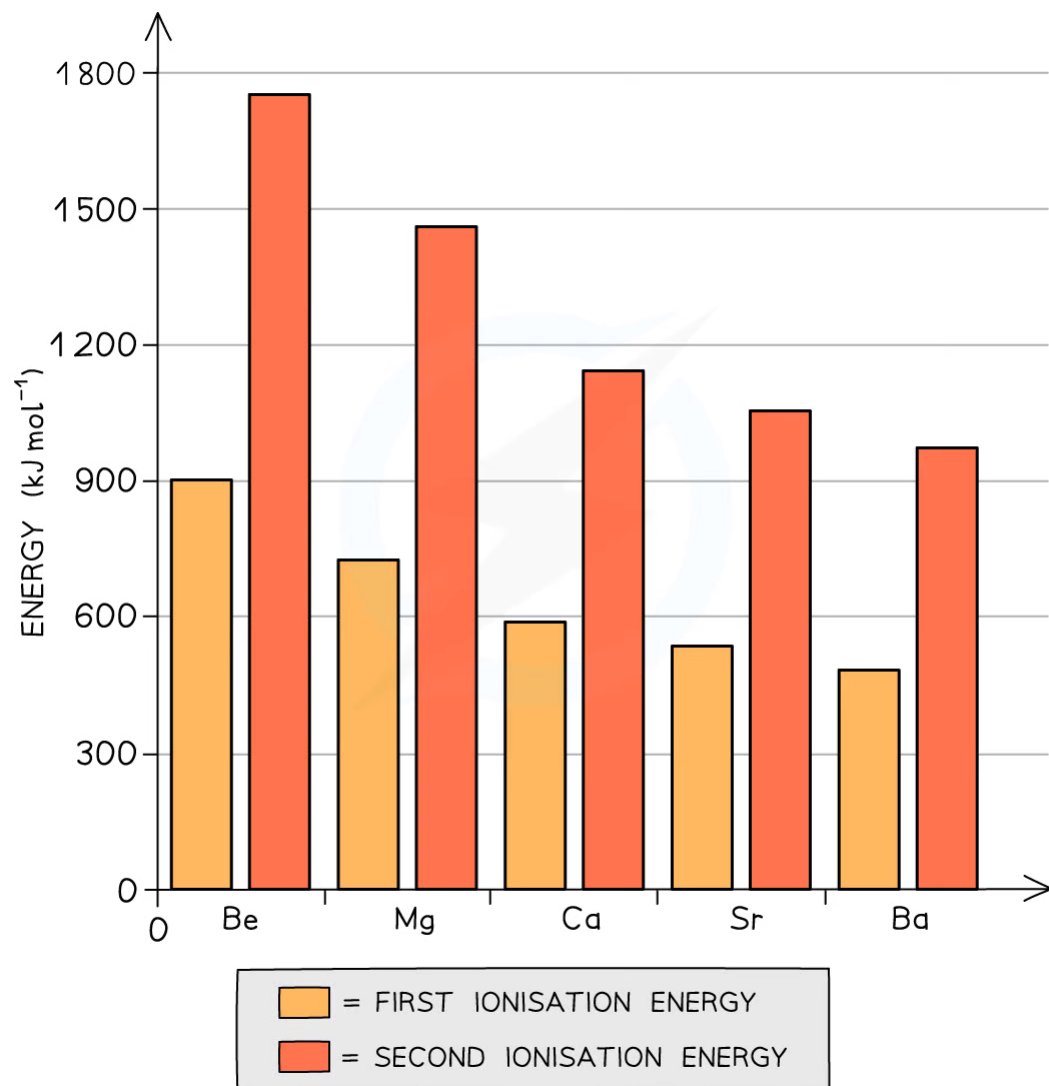
2.2.1 Trends in Group 2: The Alkaline Earth Metals

Group 2 Trends

Chemical trends

- All elements in Group 2 (also called **alkali earth metals**) have two electrons in their **outermost principal quantum shell**
- All Group 2 metals can form **ionic compounds** in which they donate these **two outermost electrons** (so they act as **reducing agents**) to become an ion with +2 charge (so they themselves become **oxidised**)

- Going down the group, the metals become more **reactive**



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This can be explained by looking at the Group 2 ionisation energies:

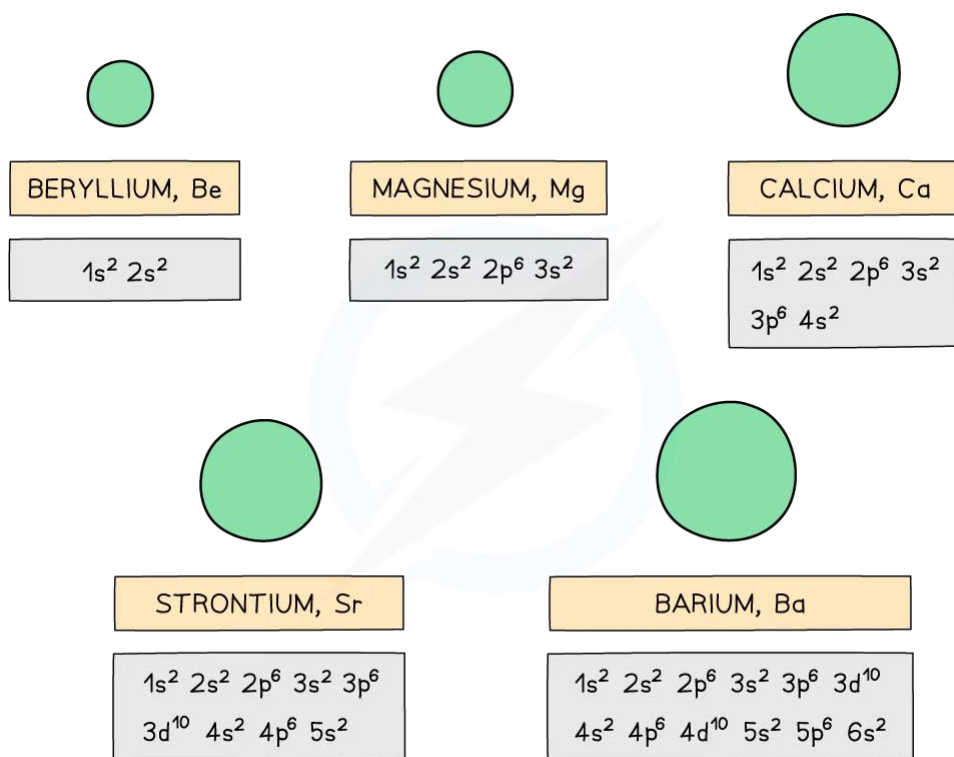


The graph shows that both the first and second ionization energies decrease going down the group

- The **first ionisation energy** is the energy needed to remove the first outer electron of an atom
The **second ionisation energy** is the energy needed to remove the second outer electron of an atom
- The graph above shows that going down the group, it becomes easier to remove the outer two electrons of the metals
- Though the **nuclear charge** increases going down the group (because there are more protons), factors such as an **increased shielding effect** and a **larger distance** between the outermost electrons and nucleus outweigh the attraction of the higher nuclear charge
- As a result of this, the elements become more reactive going down the group as it gets easier for the atoms to lose two electrons and become 2+ ions This trend is shown by looking at reactions of the Group 2 metals:
 - With dilute hydrochloric acid: **bubbles of hydrogen gas** are given off much faster indicating that the reactions become more vigorous
 - With oxygen: the metals get more reactive with oxygen down the group (Ba is so reactive, that it must be stored in oil to prevent it from reacting with oxygen in air)

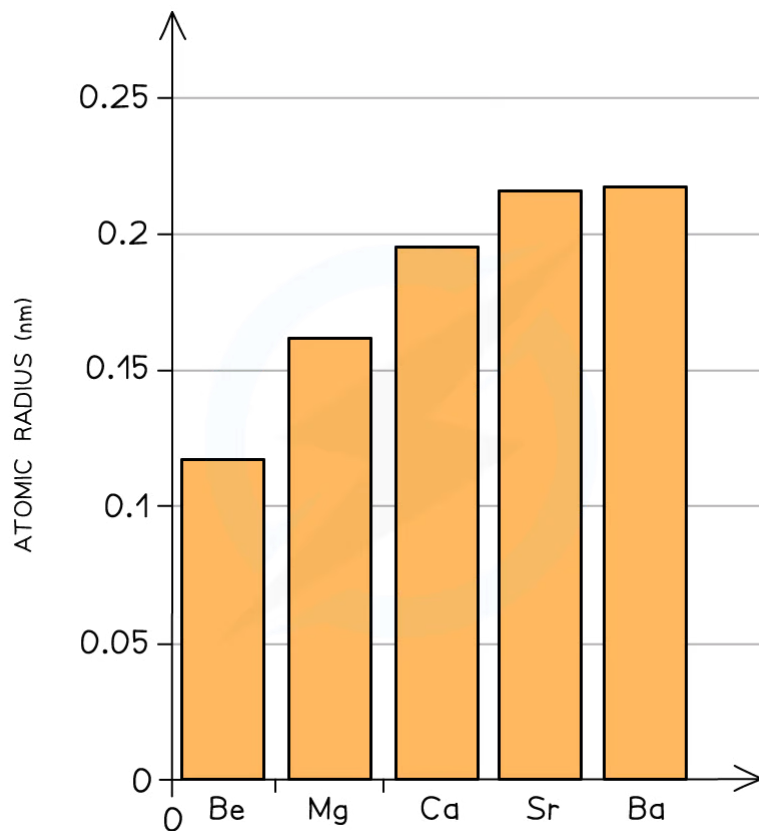
Physical trends

- Going down the group, the elements become larger as the outer two electrons occupy a new **principal quantum shell** which is further away from the nucleus



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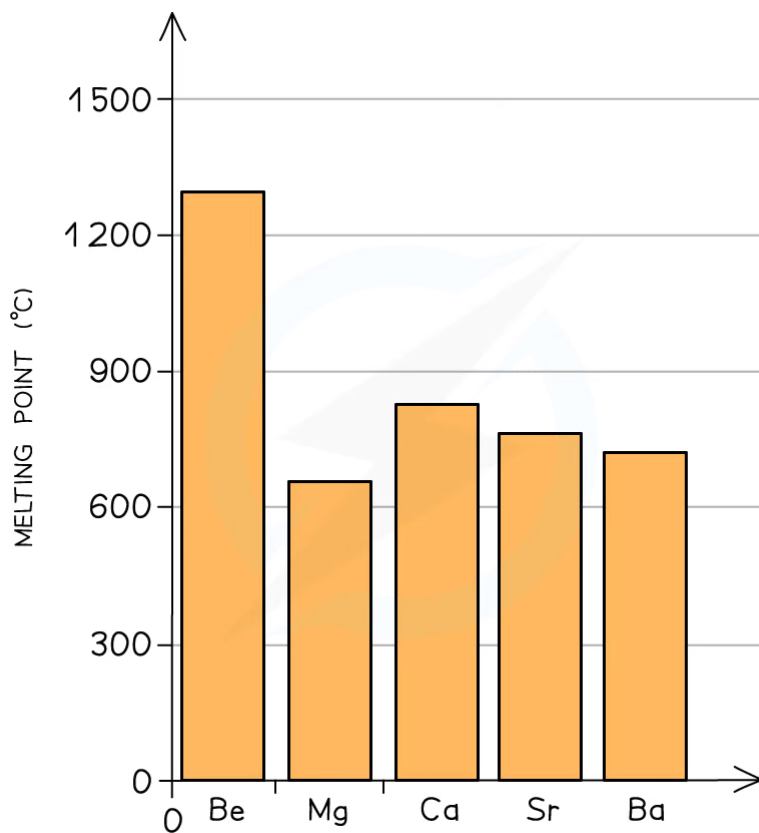
The atomic radius of the Group 2 elements increases going down the group due to the addition of an extra principal quantum shell



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The graph shows a **increase in atomic radius going down the group**

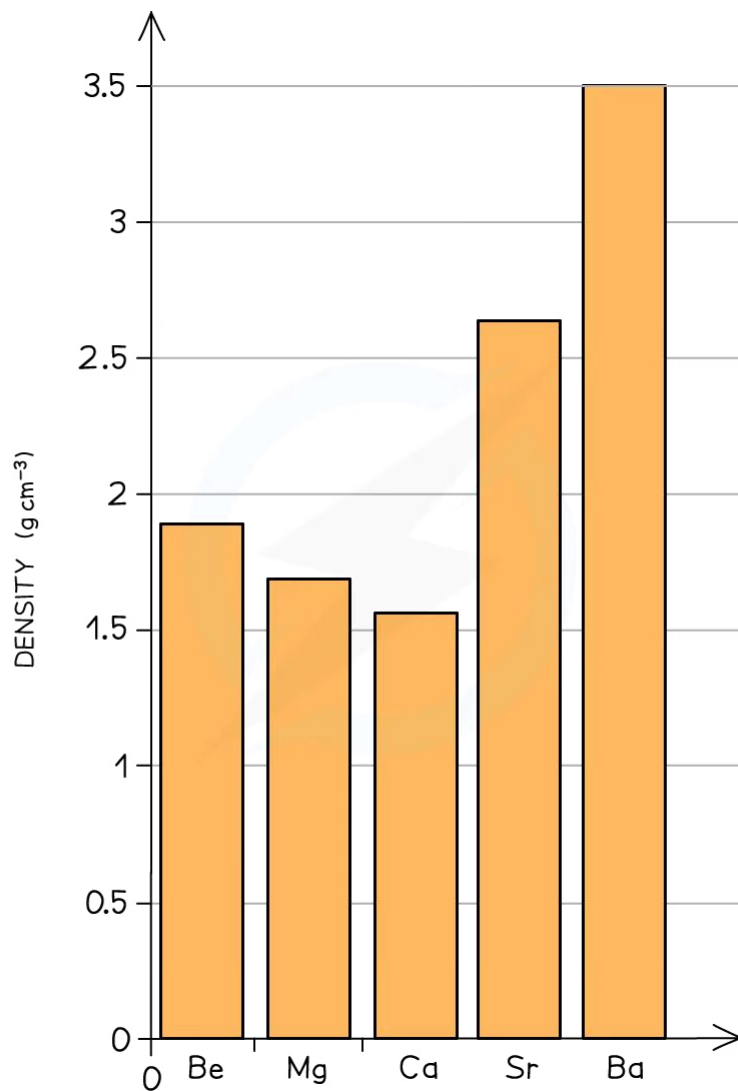
- The **melting point** of the elements decreases going down the group as the outer electrons get further away from the nucleus
- This means that the **attraction** between the **nucleus** and the **bonding electrons** decreases causing a decrease in melting point



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The graph shows a general decrease in melting point going down the group with a slight anomaly in magnesium

- As you go down the group, the **density** of the alkali earth metals drops and then increases
Density is also affected by the packing structure of the metals, not just the atomic radius - no trend
- is perfect!



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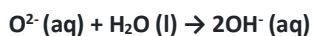
The graph show the broad trend in density going down the group

2.2.2 Solubility of Group 2 Compounds: Hydroxides & Sulfates

Solubility of Hydroxides & Sulfates

Group 2 hydroxides

- Going down the group, the solutions formed from the reaction of group 2 oxides with water become more **alkaline**
- When the oxides are dissolved in water, the following ionic reaction takes place:



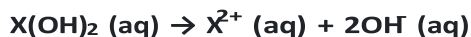
- The higher the **concentration** of OH^{-} ions formed, the more **alkaline** the solution. The **alkalinity** of the solution formed can therefore be explained by the **solubility** of the Group 2 **hydroxides**

Solubility of the Group 2 Hydroxides

| Group 2 Hydroxide | Solubility at 298K (mol / 100 g of water) |
|--------------------------|---|
| $\text{Mg}(\text{OH})_2$ | 2.0×10^{-5} |
| $\text{Ca}(\text{OH})_2$ | 1.5×10^{-3} |
| $\text{Sr}(\text{OH})_2$ | 3.4×10^{-3} |
| $\text{Ba}(\text{OH})_2$ | 1.5×10^{-2} |

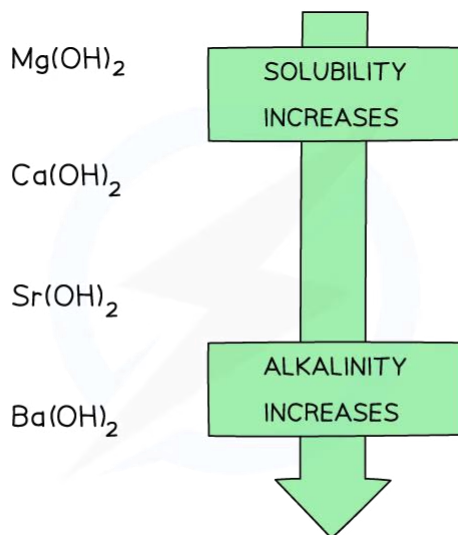
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- The hydroxides dissolve in water as follows:



Where X is the Group 2 element

- When the metal oxides react with water, a group 2 hydroxide is formed
- Going down the group, the **solubility** of these hydroxides **increases**
- This means that the **concentration** of OH^{-} ions **increases**, increasing the pH of the solution. As a result, going down the group, the **alkalinity** of the solution formed increases when Group 2 oxides react with water

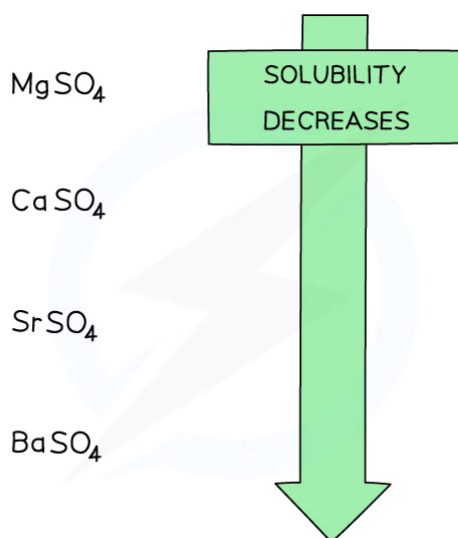


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Going down the group, the solubility of the hydroxides increases which means that the solutions formed from the reactions of the group 2 metal oxides and water become more alkaline going down the group

Group 2 sulfates

- The solubility of the group 2 sulfates decreasing going down the group



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Going down the group, the solubility of the sulfates decreases

? Worked Example

Predicting properties of radium Radium (Ra) is a radioactive element found below barium at the bottom of group 2. Applying your knowledge of the group 2 elements, predict:

Property 1: The formula of the ion formed by Ra.

Property 2: The formulae of its oxide and hydroxide

Property 3: Its first ionisation energy

Property 4: Its reactivity compared with barium

Property 5: The relative pH of its saturated hydroxide solution compared with a saturated solution of calcium hydroxide

Property 6: The solubility of its sulfate compared to strontium sulfate

Property 7: The equation for the reaction of its solid oxide with dilute hydrochloric acid

Property 8: What would you expect to see if you mixed radium hydroxide solution with dilute sulfuric acid?

Answer

Property 1:

Since, Ra is in group 2, it will form an ion with +2 charge to give Ra^{2+}

Property 2:

The group 2 oxides and hydroxides have general formula XO and X(OH)_2 respectively where X is the group 2 element.

Therefore, radium oxide is RaO and radium hydroxide is Ra(OH)_2

Property 3:

Radium is below barium so its atomic radius is larger than the atomic radius of barium.

This means that radium's outermost electrons are even further away and are therefore even more easily removed than barium's outermost electron pair. The first ionization energy is between $450\text{-}480 \text{ kJ mol}^{-1}$

Property 4:

Radium's outermost electrons are even further away than in barium and are therefore more easily removed making radium more reactive than barium.

Property 5:

The group 2 hydroxides become more soluble going down the group.

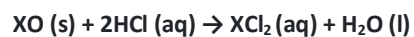
Radium hydroxide will therefore be more soluble than calcium hydroxide. Since more hydroxide ions will be present in solution the pH should be higher than the pH of calcium hydroxide

Property 6:

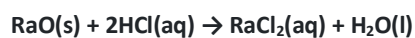
The Group 2 sulfates become less soluble going down the group.
Radium sulfate will therefore be less soluble than strontium sulfate.

Property 7:

The general equation for the reaction of group 2 oxides with dilute hydrochloric acid is:



The reaction of radium oxide with dilute hydrochloric acid is therefore:



Property 8:

Radium sulfate will be formed in this reaction, however the solubility of group 2 sulfates decreases going down the group, therefore a white precipitate of radium sulfate will be formed in this reaction

2.2.3 Reactions of Group 2

Reactions of Group 2

- The group 2 elements react with oxygen, water and dilute acids

Group 2 Reactions - Observations

| Group 2 | Reaction with Oxygen | Reaction with Water | Reaction with dilute HCl | Reaction with dilute H ₂ SO ₄ |
|---------|----------------------------------|--|--------------------------|---|
| Be | Reluctant to burn, white flame | No reaction | Reacts rapidly | Reacts vigorously |
| Mg | Burns easily, bright white flame | Vigorous reaction with steam, no reaction with water | Reacts vigorously | Reacts vigorously |
| Ca | Difficult to ignite, red flame | Reacts moderately, hydroxide formed | Reacts vigorously | Reaction slowed down by the formation of a sparingly soluble sulfate layer that covers the metal, stopping hydrogen bubbles from rising |
| Sr | Difficult to ignite, red flame | Reacts rapidly, hydroxide formed | Reacts vigorously | Reaction is quickly stopped by the formation of an insoluble sulfate layer on the metal surface |
| Ba | Difficult to ignite, green flame | Reacts vigorously, hydroxide formed | Reacts vigorously | Reaction is quickly stopped by the formation of an insoluble sulfate layer on the metal surface |

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Reactions with water and oxygen

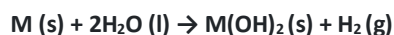
- The reaction of group 2 metals with oxygen follows the following general equation:



Where M is any metal in group 2

Remember that Sr and Ba **also** form a peroxide, MO₂

- The reaction of all metals with water follows the following general equation:



For more help, please visit www.exampaperspractice.co.uk

Except for, Be which does not react with water

Group 2 Metals reacting with Water and with Oxygen - Equations

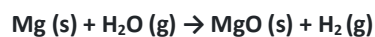
| Group 2 Element | Reaction with Oxygen | Reaction with Water |
|-----------------|---|--|
| Be | $2\text{Be(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{BeO(s)}$ | – |
| Mg | $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$ | $\text{Mg(s)} + \text{H}_2\text{O(g)} \rightarrow \text{MgO(s)} + \text{H}_2\text{(g)}$ |
| Ca | $2\text{Ca(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{CaO(s)}$ | $\text{Ca(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(s)} + \text{H}_2\text{(g)}$ |
| Sr | $2\text{Sr(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{SrO(s)}$ $\text{Sr(s)} + \text{O}_2\text{(g)} \rightarrow \text{SrO}_2\text{(s)}$ | $\text{Sr(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Sr(OH)}_2\text{(s)} + \text{H}_2\text{(g)}$ |
| Ba | $2\text{Ba(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{BaO(s)}$ $\text{Ba(s)} + \text{O}_2\text{(g)} \rightarrow \text{BaO}_2\text{(s)}$ | $\text{Ba(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ba(OH)}_2\text{(s)} + \text{H}_2\text{(g)}$ |

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- Magnesium reacts **extremely slowly** with cold water:



- The solution formed is **weakly alkaline** (pH 9-10) as **magnesium hydroxide** is only **slightly soluble**
- However, when magnesium is **heated in steam**, it reacts **vigorously** with steam to make magnesium oxide and hydrogen gas:



Reactions of Group 2 metals with acid

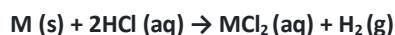
- The Group 2 metals will react with dilute acids to form **colourless solutions of metal salts**
For example, they will form colourless solutions of metal chlorides if reacted with hydrochloric acid
- When metals react with an acid, the by-product of this reaction is hydrogen gas

Group 2 Reactions with Dilute Acids - Equations

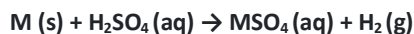
| Group 2 Element | Reaction with dilute HCl | Reaction with dilute H ₂ SO ₄ |
|-----------------|--|---|
| Be | $\text{Be(s)} + 2\text{HCl(aq)} \rightarrow \text{BeCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ | $\text{Be(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{BeSO}_4\text{(aq)} + \text{H}_2\text{(g)}$ |
| Mg | $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ | $\text{Mg(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{MgSO}_4\text{(aq)} + \text{H}_2\text{(g)}$ |
| Ca | $\text{Ca(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ | $\text{Ca(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{CaSO}_4\text{(s)} + \text{H}_2\text{(g)}$ |
| Sr | $\text{Sr(s)} + 2\text{HCl(aq)} \rightarrow \text{SrCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ | $\text{Sr(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{SrSO}_4\text{(s)} + \text{H}_2\text{(g)}$ |
| Ba | $\text{Ba(s)} + 2\text{HCl(aq)} \rightarrow \text{BaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ | $\text{Ba(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)} + \text{H}_2\text{(g)}$ |

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- When some of Group 2 metals react with sulfuric acid rather than hydrochloric, an insoluble sulfate forms
- Going down the group, the Group 2 sulfates become less and less soluble
 - Calcium sulfate is sparingly soluble, but strontium sulfate and barium sulfate are insoluble
- The reaction of the metals with dilute HCl follows the following general equation:



- The reaction of the metals with dilute H₂SO₄ follows the following general equation:



Remember that SrSO₄ and BaSO₄ are insoluble

Exam Tip

Learn the general equation for the reaction with oxygen, water and dilute HCl/H₂SO₄ and the exceptions instead of memorizing the entire table!

Group 2 Oxides, Hydroxides & Carbonates

Reactions of group 2 oxides with water

- All group 2 oxides are **basic**, except for BeO which is **amphoteric** (it can act both as an acid and base)
- Group 2 oxides react water to form **alkaline** solutions which get more **alkaline** going down the group

Group 2 Oxides reacting with Water

| Group 2 Oxide | Reaction with Water | Observations |
|---------------|--|---|
| BeO | $\text{BeO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Be(OH)}_2\text{(aq)}$ | |
| MgO | $\text{MgO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Mg(OH)}_2\text{(aq)}$ | MgO is only slightly soluble in water, therefore a weakly alkaline solution is formed (pH 10) |
| CaO | $\text{CaO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(aq)}$ | Vigorous reaction releases a lot of energy which cause some of the water to boil off as the CaO solid lump seems to expand and crack open (pH 11) |
| SrO | $\text{SrO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Sr(OH)}_2\text{(aq)}$ | |
| BaO | $\text{BaO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Ba(OH)}_2\text{(aq)}$ | |

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- Remember that:

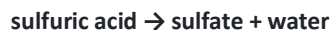
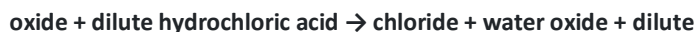


- You should know that calcium hydroxide, when in solution, is also called **limewater**

Reactions of Group 2 oxides with acid

- Group 2 sulfates also form when a group 2 oxide is reacted with sulfuric acid
- The **insoluble sulfates** form at the **surface** of the oxide, which means that the solid oxide beneath it can't react with the acid
- This can be prevented to an extent by using the oxide in **powder** form and **stirring** in which case neutralisation can take place

Remember that:



Reactions of group 2 hydroxides

- The group 2 metal hydroxides form **colourless solutions of metal salts** when they react with a dilute acid
- The sulfates decrease in **solubility** going down the group (barium sulfate is an insoluble white precipitate)

Group 2 Hydroxide Reactions with Dilute Acids

| Group 2 Hydroxide | Reaction with dilute HCl | Reaction with dilute H ₂ SO ₄ |
|-------------------|---|--|
| BeOH | $\text{Be}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{BeCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ | $\text{Be}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{BeSO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ |
| MgOH | $\text{Mg}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ | $\text{Mg}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{MgSO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ |
| CaOH | $\text{Ca}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ | $\text{Ca}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ |
| SrOH | $\text{Sr}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{SrCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ | $\text{Sr}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{SrSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ |
| BaOH | $\text{Ba}(\text{OH})_2(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow \text{BaCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$ | $\text{Ba}(\text{OH})_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ |

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- Remember that:

hydroxide + dilute hydrochloric acid → chloride + water

hydroxide + dilute sulfuric acid → sulfate + water

Reactions of group 2 carbonates

- All group 2 carbonates (except for **BeCO₃**) are **insoluble in water**
- All group 2 carbonates will form **soluble chloride salts**, water and carbon dioxide gas when reacted with dilute hydrochloric acid
- The carbonates of Ca, Sr and Ba form as an **insoluble sulfate layer** on their solid carbonates which

Group 2 Carbonates reacting with Dilute Acids

| Group 2 Carbonate | Reaction with dilute HCl | Reaction with dilute H ₂ SO ₄ |
|-------------------|--|--|
| BeCO ₃ | $\text{BeCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{BeCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ | $\text{BeCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{BeSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ |
| MgCO ₃ | $\text{MgCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ | $\text{MgCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{MgSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ |
| CaCO ₃ | $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ | $\text{CaCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ |
| SrCO ₃ | $\text{SrCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{SrCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ | $\text{SrCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{SrSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ |
| BaCO ₃ | $\text{BaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{BaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ | $\text{BaCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ |

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- Remember that:
stops any further reaction after the initial bubbling

(effervescence) of carbon dioxide gas is seen **carbonate + dilute hydrochloric acid → chloride**

+ water + carbon dioxide **carbonate + dilute sulfuric acid → sulfate + water + carbon**

dioxide

2.2.4 Uses of Group 2 Elements

Uses of Group 2

- Three different calcium compounds are used commonly in agriculture, construction and iron extraction:
 - **Limestone** - this is impure calcium carbonate
 - **Quicklime** - this is calcium oxide, formed by the thermal decomposition of calcium carbonate
 - **Slaked lime** - this is calcium hydroxide formed when water is added to quicklime
- All three materials are used in **agriculture** to raise the pH of the soil
 - Over time, the soil becomes more **acidic** while the **optimum pH** for many crops to grow is at around
- 6-6.5
 - The compounds are all bases and react with the acids in the soil and raise the pH of the soil
- Calcium carbonate is more commonly used in **agricultural lime** as it is **cheaper** and **safer** to handle
 - However, due to calcium carbonate being largely insoluble, it acts more slowly than calcium
 - hydroxide

2.3 Group 7 (17), the Halogens

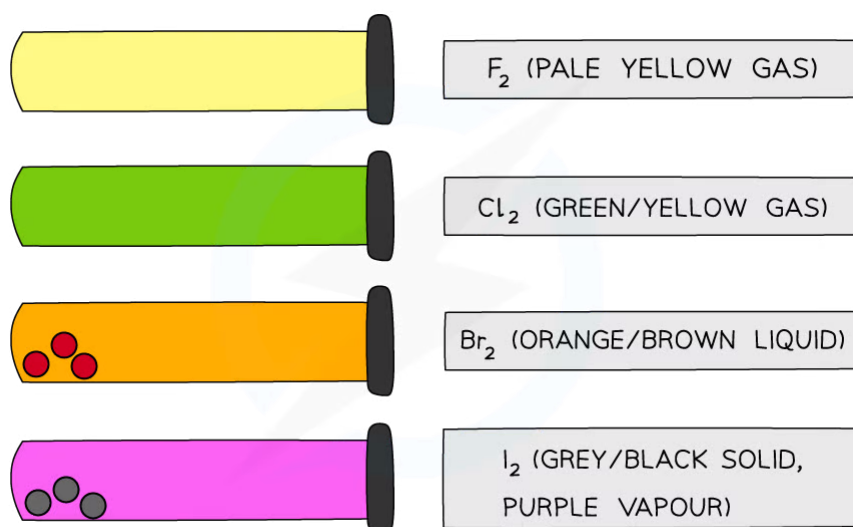
2.3.1 Physical Properties of Group 7

Group 7 Physical Properties

- The group 7 elements are called **halogens**
- The halogens have uses in water purification and as bleaching agents (chlorine), as flame-retardants and fire extinguishers (bromine) and as antiseptic and disinfectant agents (iodine)

Colours

- All halogens have distinct **colours** which get **darker** going down the group

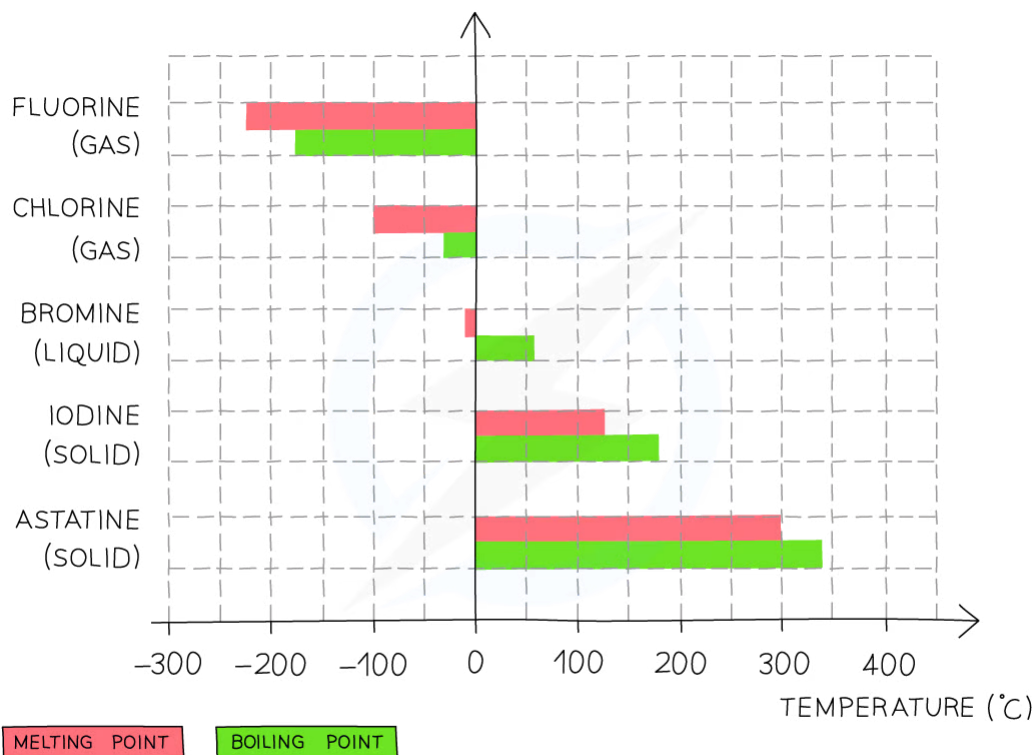


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The colours of the Group 7 elements get darker going down the group

Volatility

- **Volatility** refers to how easily a substance can evaporate
 - A volatile substance will have a low boiling point

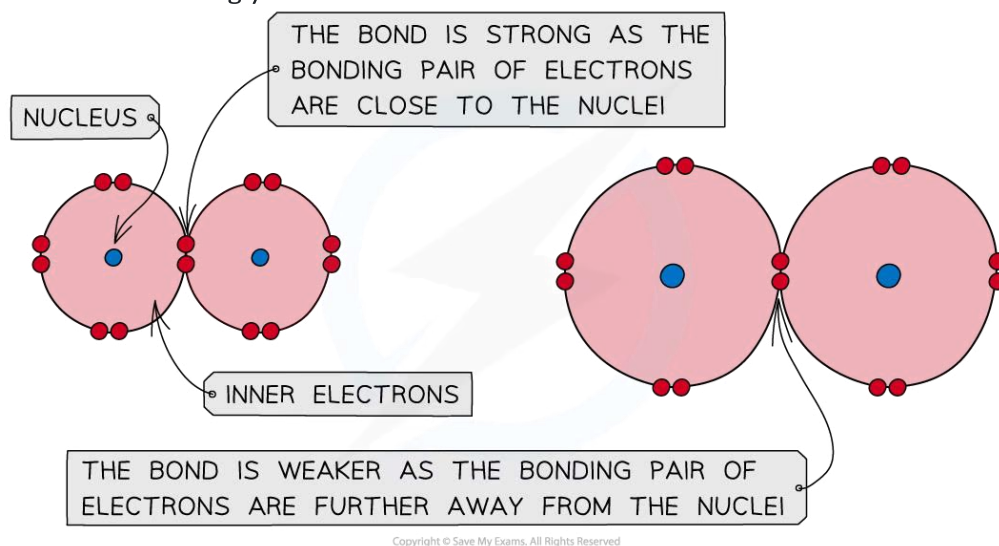


The melting and boiling points of the Group 7 elements increase going down the group which indicates that the elements become less volatile

- Going down the group, the **boiling point** of the elements increases which means that the **volatility** of the halogens decreases
 - This means that fluorine is the most volatile and iodine the least volatile

Bond Strength

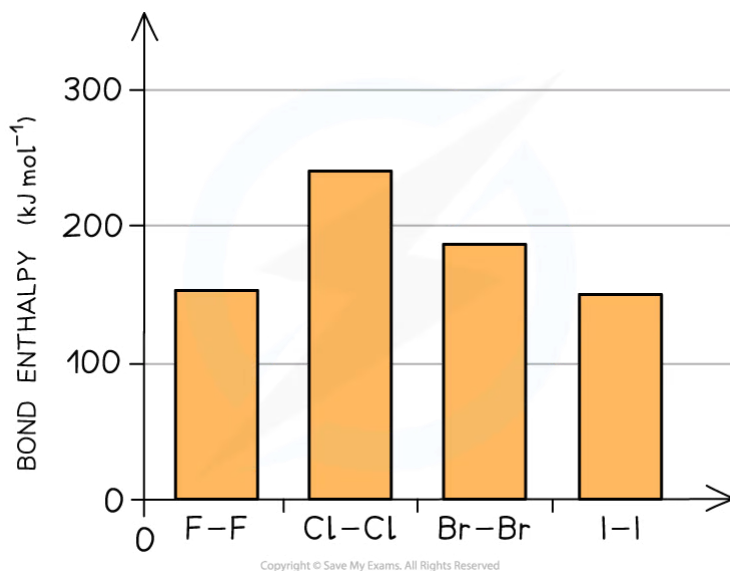
- Halogens are **diatomic** molecules in which **covalent bonds** are formed by overlapping their orbitals
- In a covalent bond, the **bonding pair** of electrons is attracted to the nuclei on either side and it is this attraction that holds the molecule together
- Going down the group, the atomic size of the halogens **increases**
- The bonding pair of electrons get further away from the halogen nucleus and are therefore less strongly attracted towards it



A covalent bond is formed by the orbital overlap of two atoms and the attraction of electrons towards the nuclei; the bigger the atom, the weaker the covalent bond

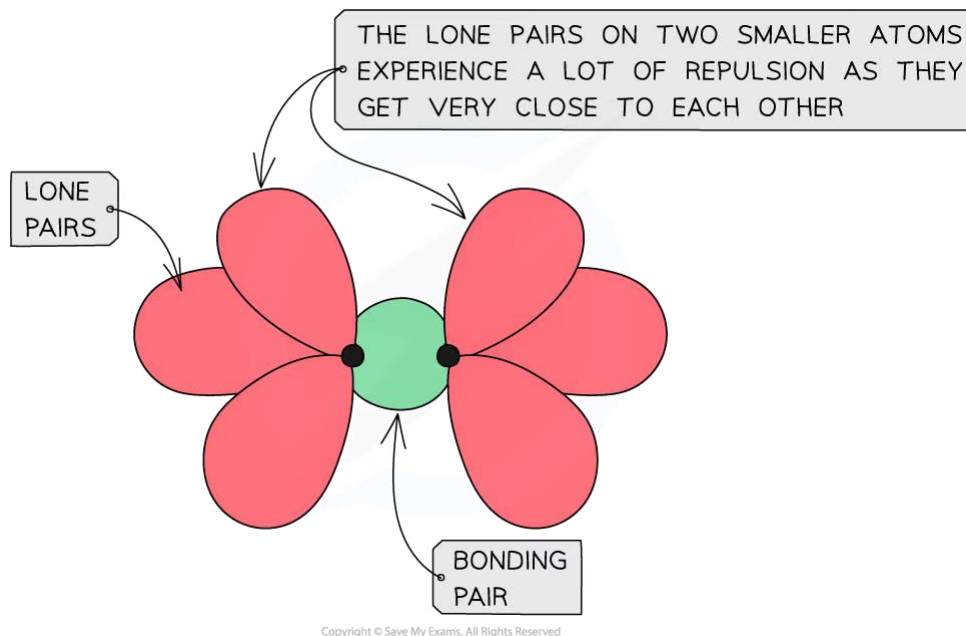
Bond enthalpy data

- The bond strength of the halogen molecules decreases going down the group



The bond enthalpies decrease indicating that the bond strengths decrease going down the group

- Bond enthalpy is the energy needed to break one mole of covalent bonds
- The higher the bond enthalpy, the stronger the bond
- An exception to this is **fluorine** which has a smaller bond enthalpy than chlorine and bromine
- Fluorine is so small that when two atoms of fluorine get together their **lone pairs** get so close that
- they cause significant **repulsion** counteracting the attraction between the bonding pair of electrons and two nuclei

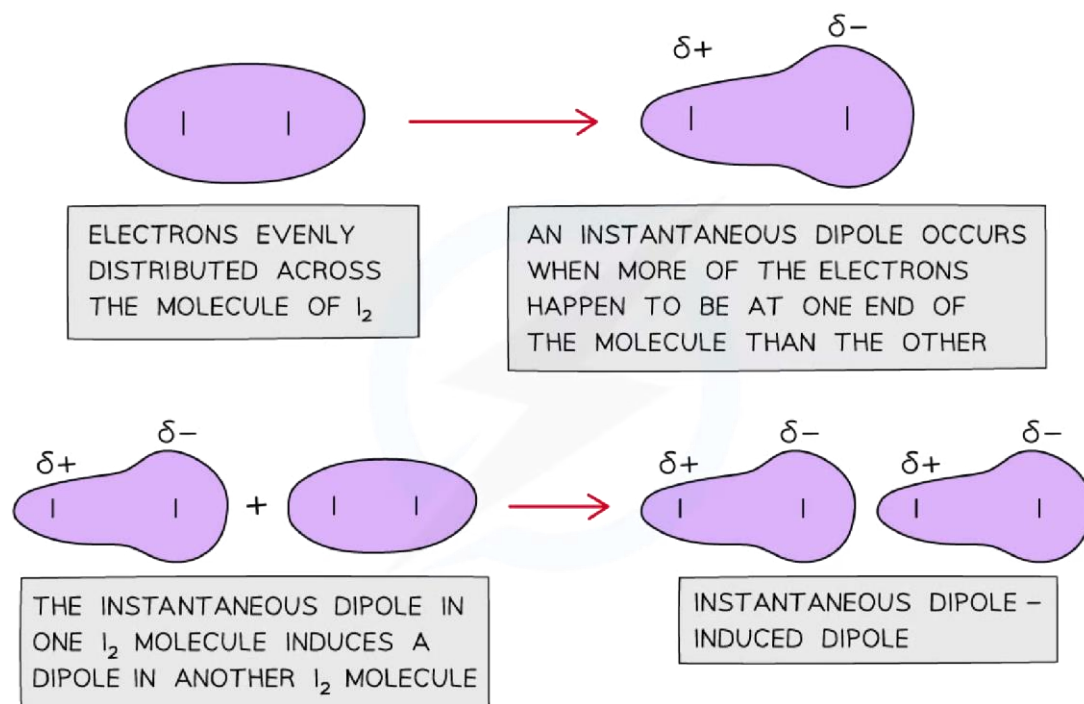


The lone pairs on fluorine get so close to each other in a fluorine molecule that they cause repulsion which decreases the bond strength

Dipole Forces & Volatility

- Halogens are non-metals and are **diatomic molecules** at room temperature This means that they
 - exist as molecules which are made up of two similar atoms, such as F_2

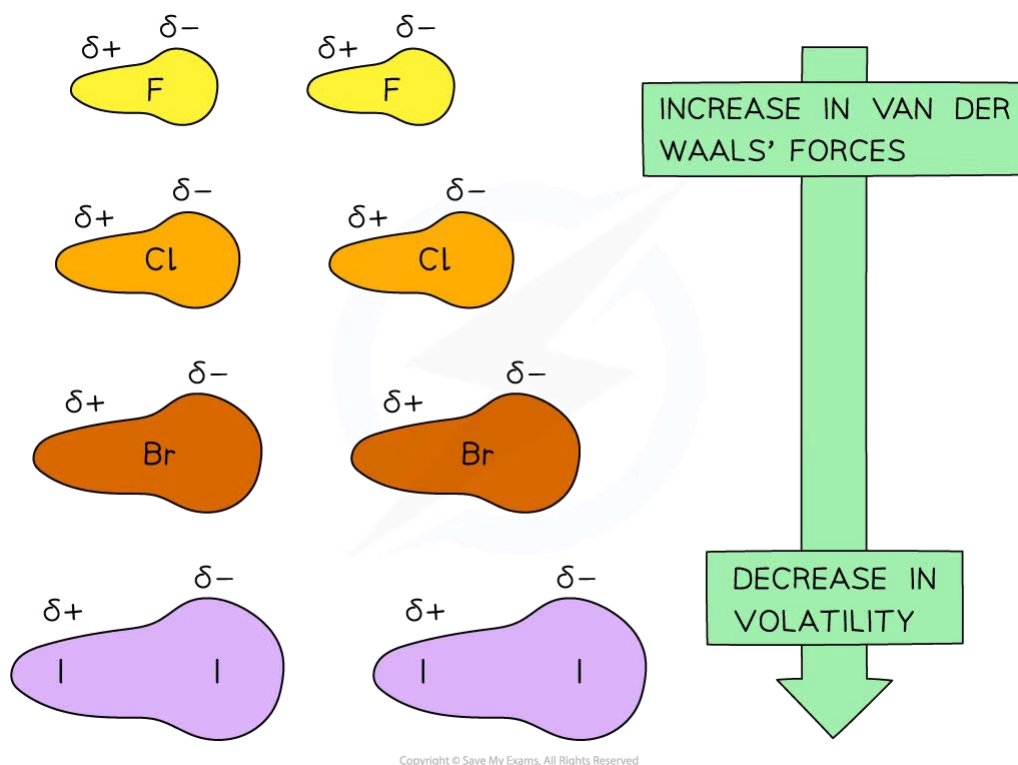
- The halogens are **simple molecular structures** with **weak** van der Waals' forces between the diatomic molecules caused by instantaneous dipole-induced dipole



The diagram shows that a sudden imbalance of electrons in a nonpolar molecule can cause an instantaneous dipole. When this molecule gets close to another non-polar molecule it can induce a dipole as the cloud of electrons repel the electrons in the neighbouring molecule to the other side

- The more **electrons** there are in a molecule, the greater the instantaneous dipole-induced dipole forces
- Therefore, the **larger** the molecule the **stronger** the van der Waals' forces between molecules
- This is why as you go down the group, it gets more difficult to separate the molecules and the **melting** and **boiling points** increase
- As it gets more difficult to separate the molecules, the **volatility** of the halogens **decreases** going down the group

forces



Going down the group, the van der Waals' forces increase due to an increased number of electrons in the molecules which means that the volatility decreases



Exam Tip

Instantaneous induced – induced dipole forces are a type of van der Waals' forces.

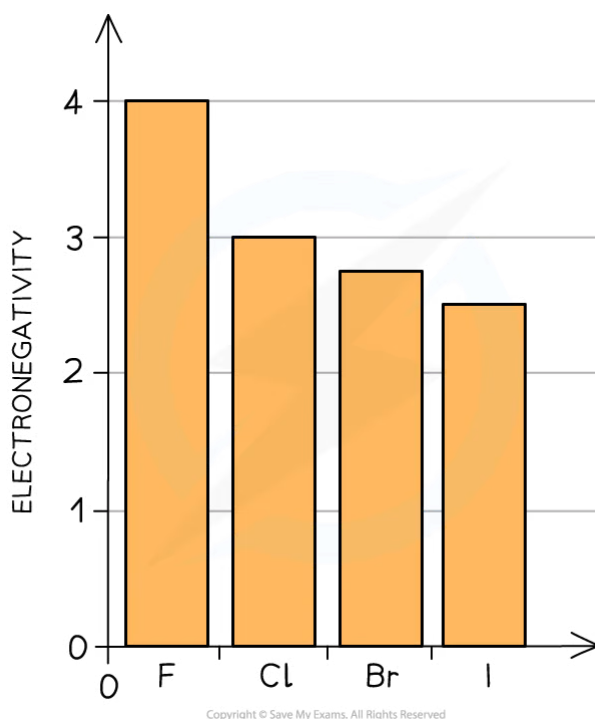
2.3.2 Chemical Properties of Group 7

Halogens Oxidising Agents

- Halogens react with metals by accepting an electron from the metal atom to become an ion with 1- charge

Eg. $\text{Ca (s)} + \text{Cl}_2\text{(g)} \rightarrow \text{CaCl}_2\text{(s)}$ consisting of Ca^{2+} and 2Cl^- ions

- Halogens are therefore **oxidising agents**:
 - Halogens **oxidise** the metal by removing an electron from the metal (the oxidation number of the metal increases)
 - Halogens become **reduced** as they gain an extra electron from the metal atom (the oxidation number of the halogen decreases) towards itself in a
- The **oxidising power** of the halogens **decreases going down the group** (the halogens get less reactive)
- This can be explained by looking at their electronegativities:



The electronegativity of the halogens decreases going down the group

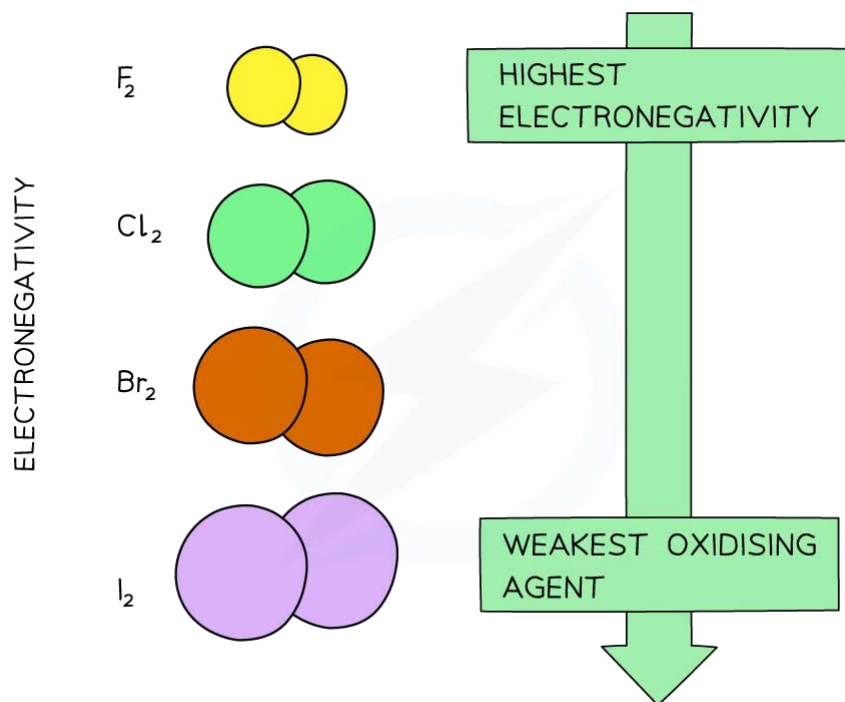
- The **electronegativity** of an atom refers to how strongly it attracts electrons
-
-
- covalent bond

The decrease in electronegativity is linked to the size of the halogens

Going down the group, the atomic radii of the elements increase which means that the outer shells get further away from the nucleus

An 'incoming' electron will therefore experience more **shielding** from the attraction of the positive nuclear charge

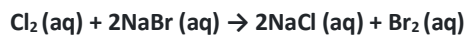
- The halogens' ability to accept an electron (their **oxidising power**) therefore decreases going down the group



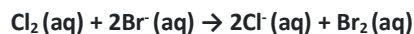
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With increasing atomic size of the halogens (going down the group) their electronegativity, and therefore oxidising power, decreases

- The reactivity of halogens is also shown by their **displacement reactions** with other halide ions in solutions
- A **more reactive** halogen can displace a **less reactive** halogen from a halide solution of the less reactive halogen
 - Eg. The addition of chlorine water to a solution of bromine water:



- The chlorine has displaced the bromine from solution as it is more reactive which can be summarised in the following ionic equation:



Reaction with Hydrogen

Halogens react with hydrogen gas to form **hydrogen halides**

- Due to the decrease in reactivity of the halogens going down the group, the reactions between halogen and hydrogen gas become less vigorous. The table below shows a summary of the reaction between the halogen and hydrogen gas.

Reaction between Halogen & Hydrogen Gas

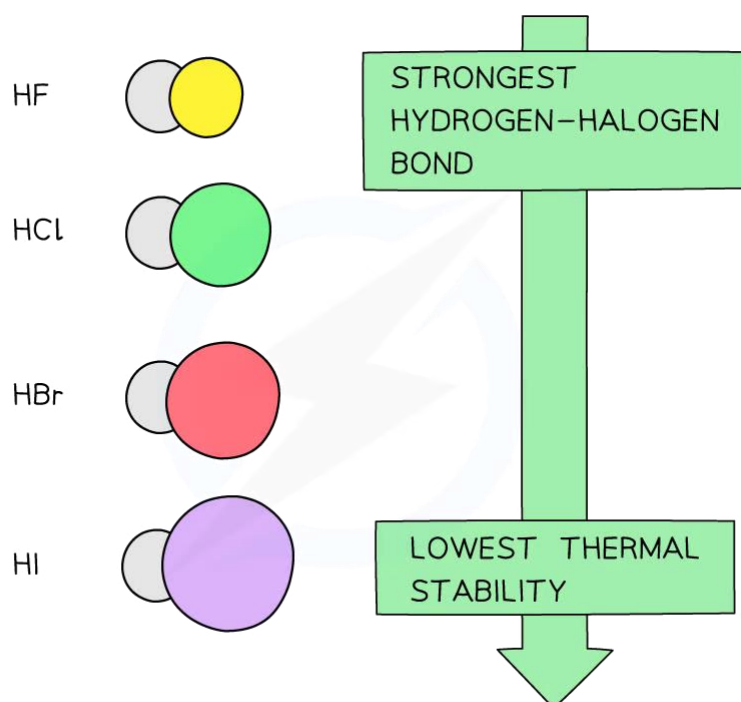
| Equation for Reaction | Description of Reaction |
|---|--|
| $\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \longrightarrow 2\text{HF}(\text{g})$ | Reacts explosively even in cool, dark conditions |
| $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \longrightarrow 2\text{HCl}(\text{g})$ | Reacts explosively in sunlight |
| $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \longrightarrow 2\text{HBr}(\text{g})$ | Reacts slowly on heating |
| $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ | Forms an equilibrium mixture on heating |

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Hydrogen Halides

Thermal stability refers to how well a substance can resist breaking down when heated

- A substance that is thermally stable will break down only at high temperatures
- The hydrogen halides formed from the reaction of halogen and hydrogen gas decrease in **thermal stability** going down the group
- The decrease in thermal stability can be explained by looking at the bond energies of the hydrogen-halogen bond
 - Going down the group, the atomic radius of the halogens increases
 - The overlap of its outer shell with a hydrogen atom therefore gives a longer bond length
 - The longer the bond, the weaker it is, and the less energy required to break it
- As the bonds get weaker, the hydrogen halogens become less stable to heat going down the group



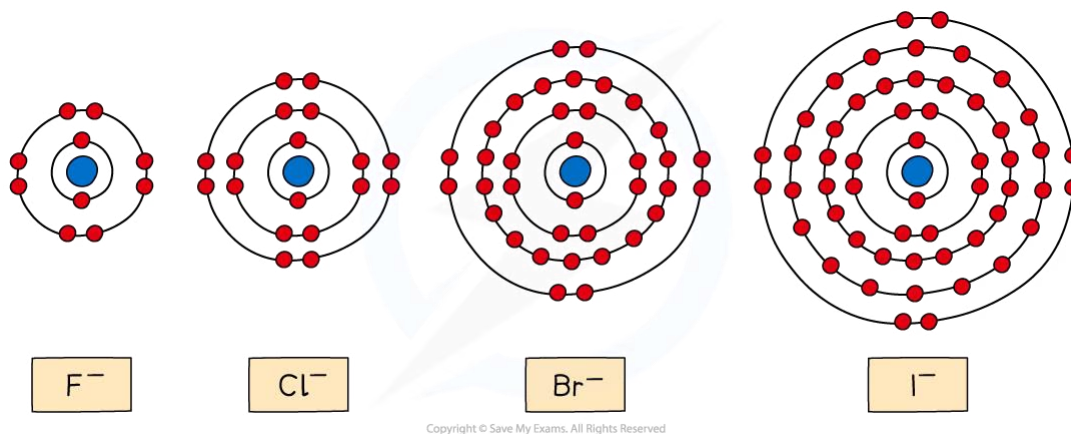
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The thermal stability of the hydrogen halide decreases going down the group as their bonds become weaker due to the increased atomic radius of the halogens

Halide Ions Reducing Agents

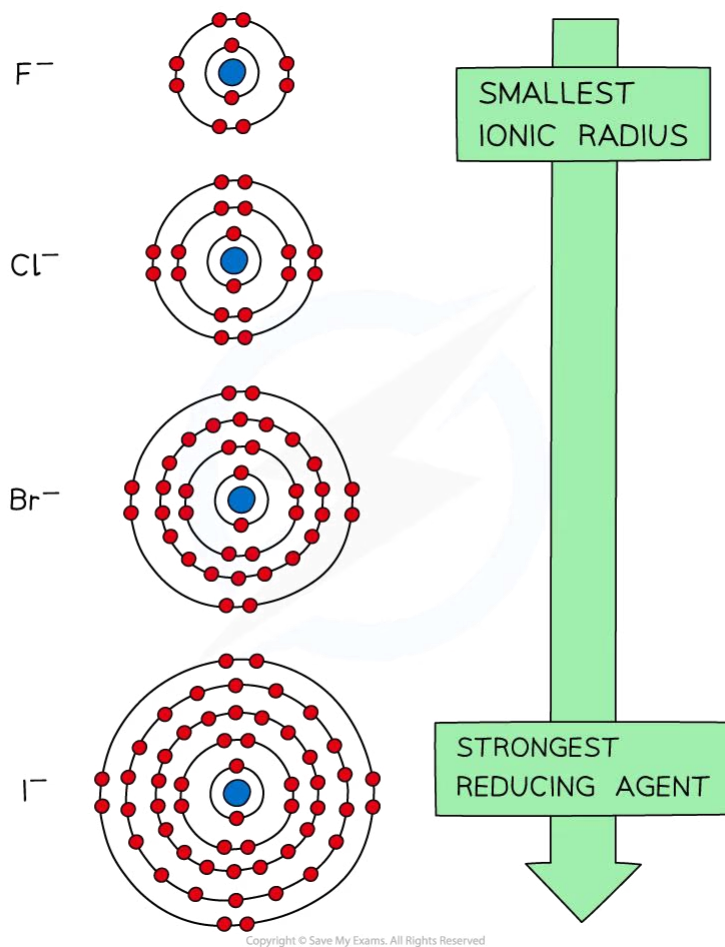
Halide ions can also act as **reducing agents** and donate electrons to another atom

- The halide ions themselves get **oxidised** and lose electrons
- The **reducing power** of the halide ions **increases** going down the group
- This trend can be explained by looking at the ionic radii of the halide ions



The diagram shows that going down the group the ionic radii of the halide ions increases

- Going down the group, the halide ions become **larger**
- The outermost electrons get further away from the nucleus
- The outermost electrons also experience more **shielding** by inner electrons. As a result of this, the
- outermost electrons are held **less tightly** to the positively charged nucleus
- Therefore, the halide ions lose electrons more easily going down the group and their **reducing power** increases



The reducing power of the halide ions increases going down the group



Exam Tip

Halogen is derived from a Greek word meaning salt-maker, a good reminder of their reaction with metals to form salts.

2.3.3 Testing for Halide Ions

Testing for Halides

Silver ions & ammonia

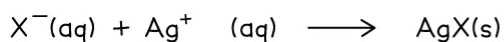
- Halide ions can be identified in an **unknown solution** by dissolving the solution in **nitric acid** and then adding **silver nitrate solution dropwise**
- The nitric acid is to prevent any false positive results from carbonate ions precipitating out with silver ions
- The halide ions will react with the silver nitrate solution as follows:



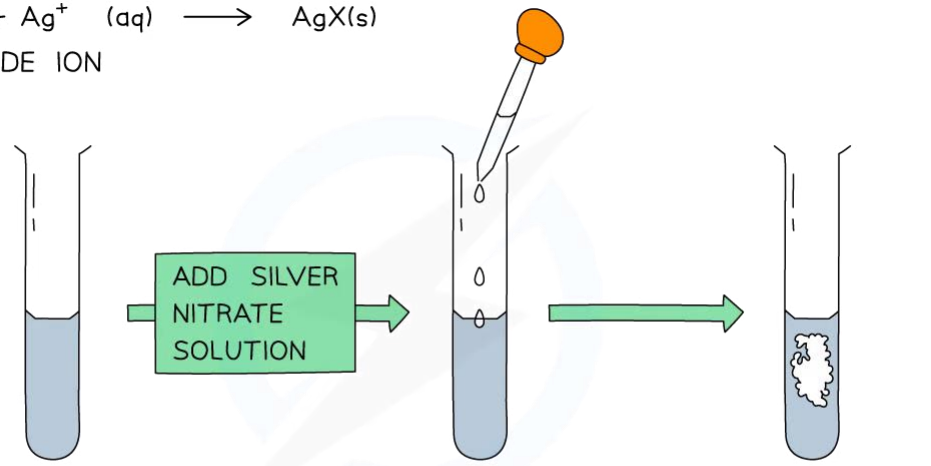
(ionic equation)

Where X^- is the halide ion

- The state symbols are key in this equation
- If the unknown solution contains halide ions, a **precipitate** of the silver halide will be formed (AgX)



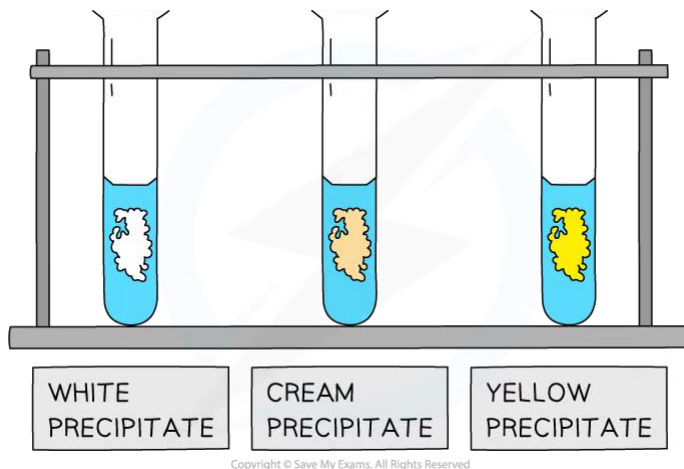
X = HALIDE ION



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A silver halide precipitate is formed upon addition of silver nitrate solution to halide ion solution

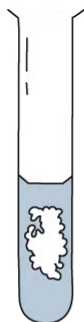
-
-
- Silver iodide (AgI) is a yellow precipitate
Silver chloride (AgCl) is a white precipitate
Silver bromide (AgBr) is a cream precipitate



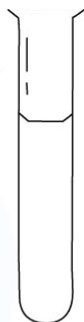
The silver halide precipitates are dense and characteristically coloured

Adding ammonia

- Because the white, cream and yellow precipitates could look very similar in colour, ammonia is often used as a follow up test to determine which halide ion is present **Dilute** followed by
- **concentrated ammonia** is added to the silver halide solution to identify the halide ion
- If the precipitate dissolves in **dilute** ammonia the unknown halide is **chloride** If the precipitate
- does not dissolve in dilute, but does dissolve in **concentrated** ammonia the unknown halide is
- **bromide**
- If the precipitate does not dissolve in **dilute** or **concentrated** ammonia, then the unknown halide is iodide

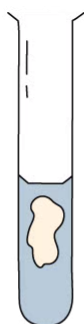


AgCl
PRECIPITATE



AgCl FORMS
A SOLUBLE
COMPLEX WITH
DILUTE AMMONIA

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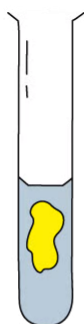


AgBr
PRECIPITATE

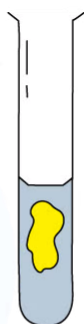


AgBr FORMS
A SOLUBLE
COMPLEX WITH
CONCENTRATED
AMMONIA

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AgI
PRECIPITATE



AgI DOESN'T
FORM A SOLUBLE
COMPLEX WITH
AMMONIA

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Silver chloride and silver bromide precipitates dissolve on addition of ammonia solution whereas silver iodide is insoluble in ammonia

Reaction of Halide Ions with Silver Nitrate & Ammonia Solutions

| Halide Ion | Colour of Silver Halide Solution | Effect of Adding Dilute Ammonia Solution to Precipitate | Effect of Adding Concentrated Ammonia Solution to Precipitate |
|----------------------|----------------------------------|---|---|
| Cl ⁻ (aq) | White | Dissolves | Dissolves |
| Br ⁻ (aq) | Cream | Remains insoluble | Dissolves |
| I ⁻ (aq) | Pale yellow | Remains insoluble | Remains insoluble |

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Concentrated sulfuric acid

- Chloride, bromide and iodide ions react with concentrated sulfuric acid to produce **toxic gases**. These reactions should therefore be carried out in a fume cupboard
- The general reaction of the halide ions with concentrated sulfuric acid is:

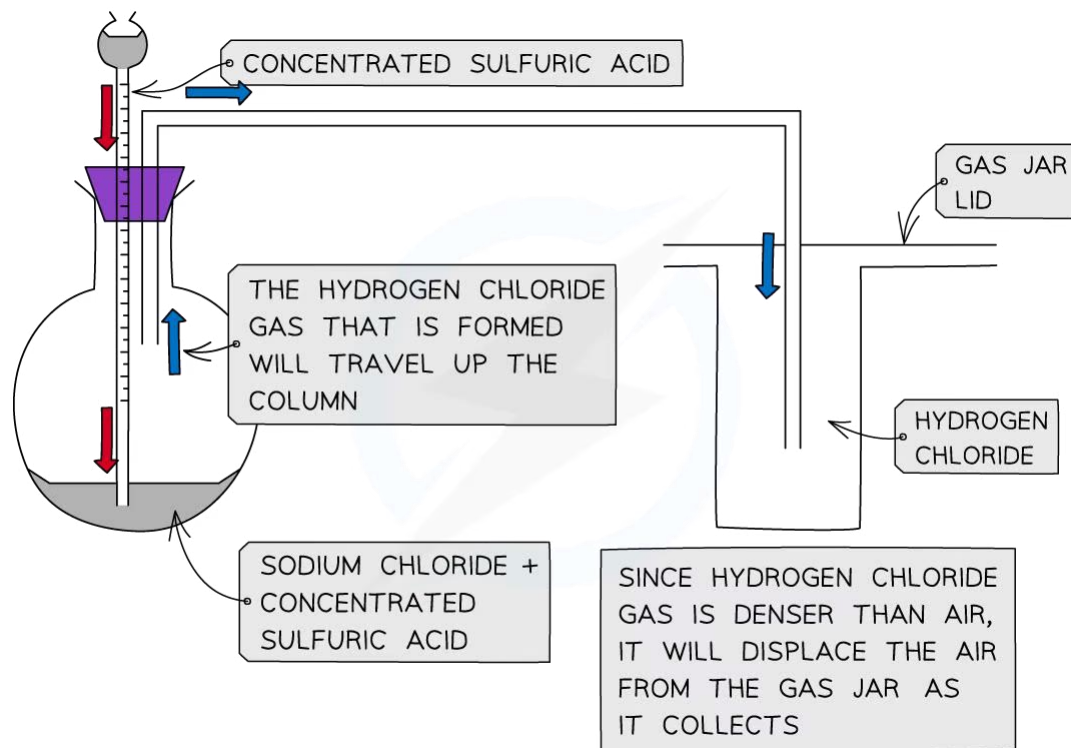


(general equation) where X is

the halide ion

Reaction of chloride ions with concentrated sulfuric Acid

- Concentrated sulfuric acid is dropwise added to sodium chloride crystals to produce **hydrogen chloride gas**



Apparatus set up for the preparation of hydrogen chloride gas from sodium chloride with concentrated sulfuric acid

- The reaction that takes place is:



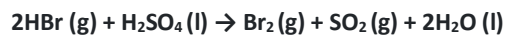
- The HCl gas produced is seen as **white fumes**

Reaction of bromide ions with concentrated sulfuric acid

- The reaction of sodium bromide and concentrated sulfuric acid is:



- The concentrated sulfuric acid **oxidises** HBr which decomposes into **bromine** and **hydrogen gas** and sulfuric acid itself is **reduced** to **sulfur dioxide gas**:



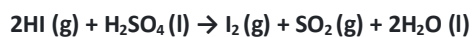
- The bromine is seen as a **reddish-brown gas**

Reaction of iodide ions with concentrated sulfuric acid

- The reaction of sodium iodide and concentrated sulfuric acid is:



- **Hydrogen iodide** decomposes readily
- Sulfuric acid oxidises the hydrogen iodide to form several products:
- The concentrated sulfuric acid **oxidises** HI and is itself **reduced** to **sulfur dioxide gas**:



- Iodine is seen as a violet/purple vapour
- The concentrated sulfuric acid **oxidises** HI and is itself **reduced** to **sulfur**:



- Sulfur is seen as a **yellow solid**
- The concentrated sulfuric acid **oxidises** HI and is itself **reduced** to **hydrogen sulfide**:



- Hydrogen sulfide has a **strong smell of bad eggs**

Summary of the Halide Ion Reactions with Concentrated Sulfuric Acid

| Halide Ion | Reaction with Concentrated Sulfuric Acid | Observations |
|---------------------------|--|--|
| $\text{Cl}^- (\text{aq})$ | $\text{H}_2\text{SO}_4(\text{l}) + \text{NaCl}(\text{s}) \longrightarrow \text{HCl}(\text{g}) + \text{NaHSO}_4(\text{s})$ | White fumes of HCl gas |
| $\text{Br}^- (\text{aq})$ | $\text{H}_2\text{SO}_4(\text{l}) + \text{NaBr}(\text{s}) \longrightarrow \text{HBr}(\text{g}) + \text{NaHSO}_4(\text{s})$ $\text{H}_2\text{SO}_4(\text{l}) + 2\text{HBr}(\text{s}) \longrightarrow \text{Br}_2(\text{g}) + \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ | Reddish brown gas of Br_2 |
| $\text{I}^- (\text{aq})$ | $\text{H}_2\text{SO}_4(\text{l}) + \text{NaI}(\text{s}) \longrightarrow \text{HI}(\text{g}) + \text{NaHSO}_4(\text{s})$ $2\text{HI}(\text{g}) + \text{H}_2\text{SO}_4(\text{l}) \longrightarrow \text{I}_2(\text{g}) + \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ $6\text{HI}(\text{g}) + \text{H}_2\text{SO}_4(\text{l}) \longrightarrow 3\text{I}_2(\text{g}) + \text{S}(\text{s}) + 4\text{H}_2\text{O}(\text{l})$ $8\text{HI}(\text{g}) + \text{H}_2\text{SO}_4(\text{l}) \longrightarrow 4\text{I}_2(\text{g}) + \text{H}_2\text{S}(\text{s}) + 4\text{H}_2\text{O}(\text{l})$ | Violet/purple vapour of I_2 Yellow solid of S Strong, bad smell of H_2S |

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Exam Tip

It gets easier to oxidise the hydrogen halides going down Group 7: the halides become stronger **reducing agents**.

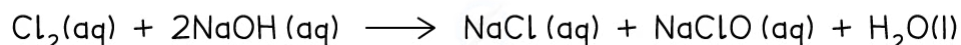
2.3.4 Uses & Reactions of Chlorine

Reactions of Halide Ions

- A **disproportionation reaction** is a reaction in which the same species is both oxidised and reduced
- The reaction of **chlorine** with **dilute alkali** is an example of a disproportionation reaction
- In these reactions, the chlorine gets oxidised and reduced at the same time
- Different reactions take place at different **temperatures** of the dilute alkali

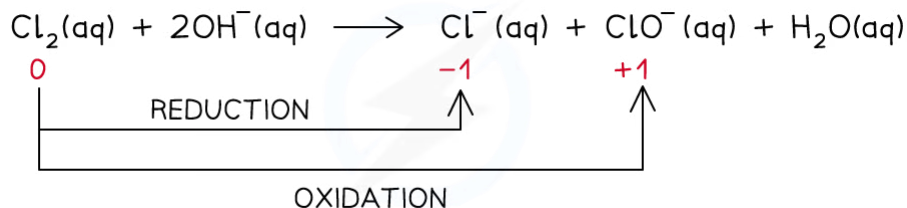
Chlorine in cold alkali (15 °C)

- The reaction that takes place is:



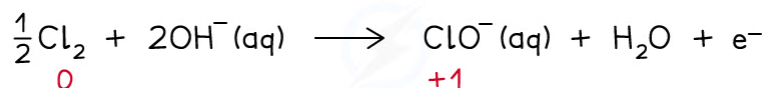
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- The ionic equation is:



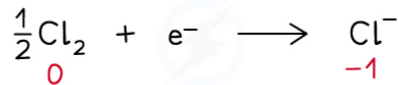
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- The ionic equation shows that the chlorine gets both oxidised and reduced
- Chlorine gets oxidised as there is an increase in ox. no. from 0 to +1 in $\text{ClO}^-(\text{aq})$
 - The half-equation for the oxidation reaction is:



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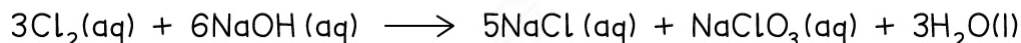
- Chlorine gets reduced as there is a decrease in ox. no. from 0 to -1 in $\text{Cl}^-(\text{aq})$
 - The half-equation for the reduction reaction is:



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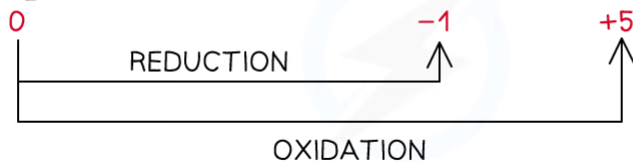
Chlorine in hot alkali (70 °C)

- The reaction that takes place is:



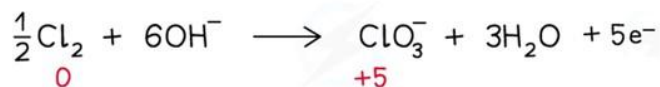
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- The ionic equation is:



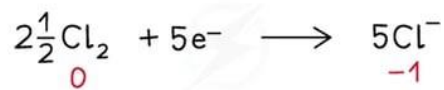
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- The ionic equation shows that the chlorine gets both oxidised and reduced
- Chlorine gets oxidised as there is an increase in ox. no. from 0 to +5 in $\text{ClO}_3^-(\text{aq})$
 - The half-equation for the oxidation reaction is:



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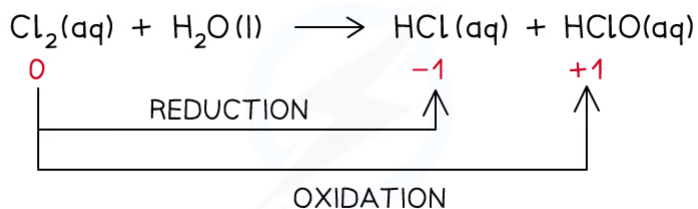
- Chlorine gets reduced as there is a decrease in ox. no. from 0 to -1 in $\text{Cl}^-(\text{aq})$
 - The half-equation for the reduction reaction is:



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Chlorine

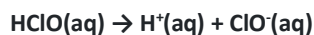
- Chlorine can be used to clean water and make it drinkable
- The reaction of chlorine in water is a **disproportionation reaction** in which the chlorine gets both **oxidised** and **reduced**



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The disproportionation reaction of chlorine with water in which chlorine gets reduced to HCl and oxidised to HClO

- Chloric(I) acid (HClO) **sterilises** water by killing **bacteria**
- Chloric acid can further dissociate in water to form $\text{ClO}^-(\text{aq})$:



- $\text{ClO}^-(\text{aq})$ also acts as a **sterilising agent** cleaning the water

? Exam Tip

Adding chlorine to a water supply is an effective way to make the water safe to drink as it forms the sterilising agent HClO which in turn dissociates in water into another sterilising agent, $\text{ClO}^-(\text{aq})$