Edexcel International AS Chemistry

Your notes

Ionic & Metallic Bonding & Structure

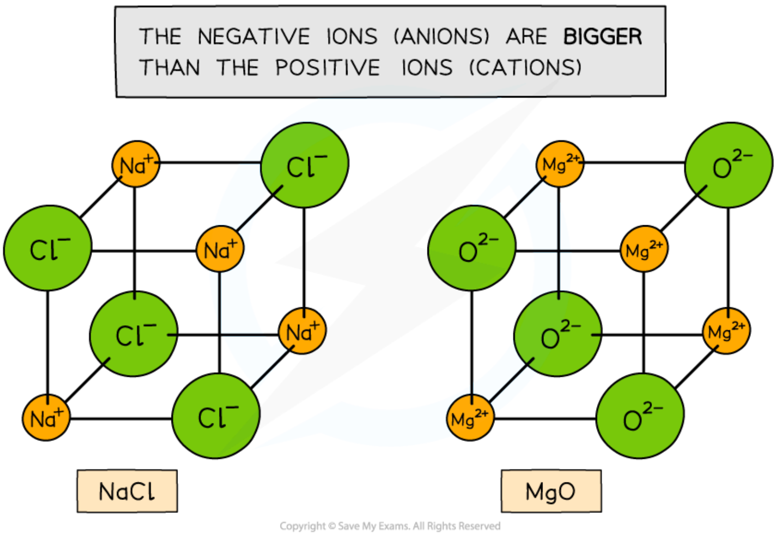
Contents

Evidence of Ions

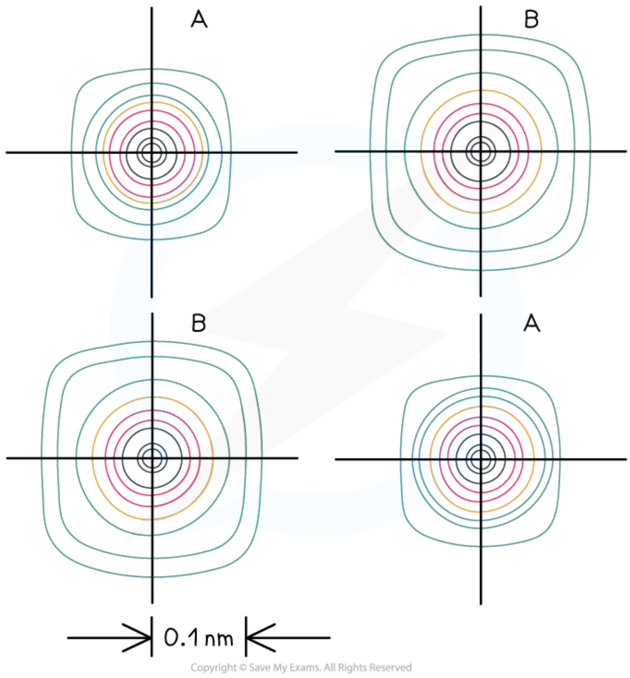
- Most ionic, metallic and covalent solids are **crystalline lattices**
- The ions, atoms or molecules are arranged in a **regular and repeating arrangement**

Giant ionic lattices

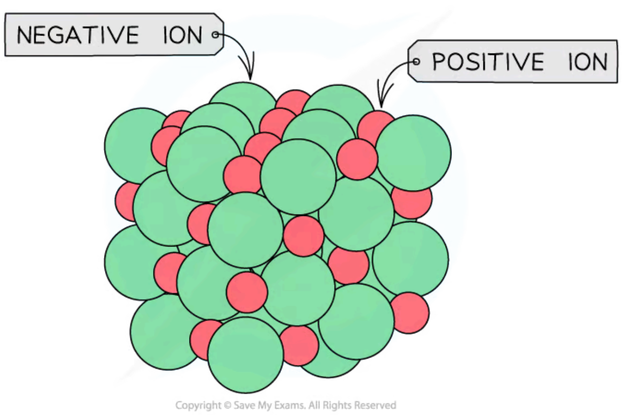
- An **ionic bond** is an electrostatic force of attraction between a positively charged metal (**cation**) ion and a negatively charged non-metal (**anion**) ion
  - The metal becomes positively charged as it transfers electrons to the non-metal which then becomes negatively charged
  - When an ionic compound is formed, the attraction between the ions happens in all directions
- Ionic compounds are arranged in giant **ionic lattices** (also called **giant ionic structures**)
- The type of lattice formed depends on the sizes of the **positive** and **negative** ions which are arranged in an **alternating** fashion
  - The ionic lattice of MgO and NaCl are **cubic**



Ionic lattices of the ionic compounds NaCl and MgO



Electron density maps shown above show the likelihood of finding electrons in a region. This shows the electron density map for NaCl. The contours are lines of equal electron density. A = Na⁺ ions (smaller), B = Cl⁻ ions (larger). Between the ions the electron density falls to zero.



General ionic lattice which shows the actual packing of the ions

Melting and boiling point

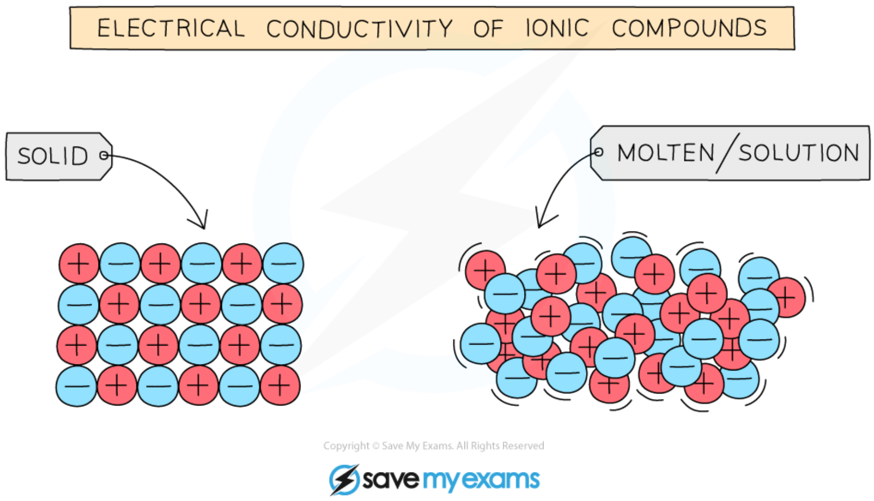
- Most ionic compounds are solids at room temperature
- This is because there isn't enough energy to overcome the strong electrostatic forces of attraction between the oppositely charged ions that make up the lattice
- Therefore, high temperatures are required to make an ionic compound melt or boil

Your notes

- Melting (and boiling) points are also higher for lattices that contain ions with a greater ionic charge
  - For example, the melting point of sodium oxide, Na_2O, is 1405 K while the melting point of calcium oxide, CaO, is 2845 K
  - This is due to a stronger attraction between the ions - the size of the ions is not a factor here as the Na^+ and Ca^{2+} ions are a similar size

Electrical Conductivity

- For electrical current to flow there must be present freely moving charged particles such as electrons or ions
- Ionic compounds can conduct electricity in the **molten** state or in **solution** as they have ions that can move and carry charge
- They cannot conduct electricity in the solid state as the ions are in fixed positions within the lattice and are unable to move



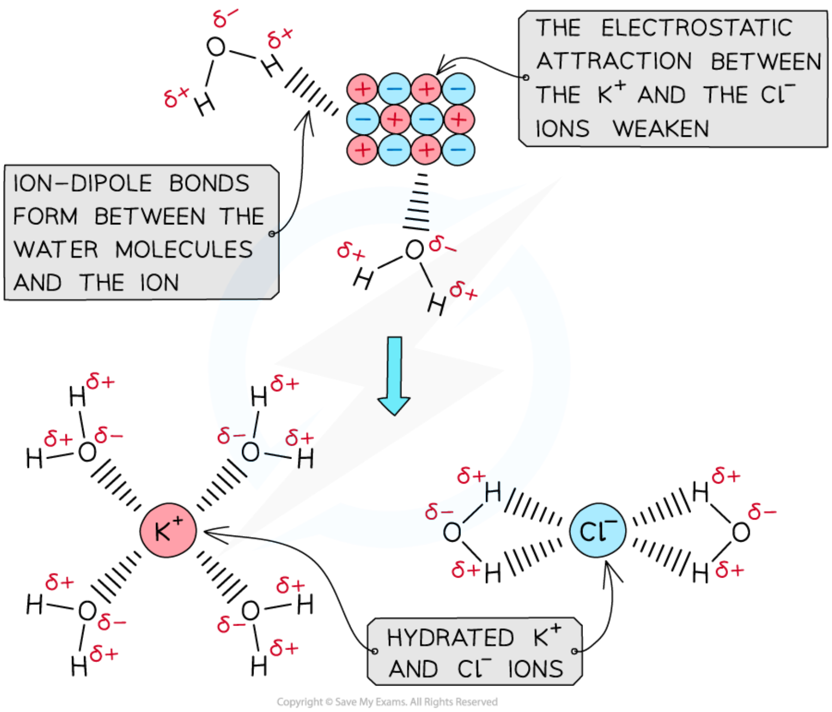
Molten or aqueous particles move and conduct electricity but cannot in the solid state

Solubility

- Many ionic compounds will dissolve in **polar** solvents, e.g. water
- Solubility is dependent on two main factors:
  1. Breaking down the ionic lattice
  2. The polar molecules attracting and surrounding the ions
- Polar molecules, such as water, can break down or disrupt the ionic lattice and surround each ion in solution
- The δ+ end of the polar molecule can surround the negative anion
- The δ- end of the polar molecule can surround the positive cation

Your notes

- The solubility of an ionic compound depends on the relative strength of the electrostatic forces of attraction within the ionic lattice and the attractions between the ions and the polar molecule
- In general, the greater the ionic charge the less soluble an ionic compound is
  - For example, 356.9 g of sodium chloride, NaCl, will dissolve in one dm^3 of water while only 74.4 g of calcium chloride will dissolve in one dm^3 of water
  - This is a general rule though and there are many exceptions

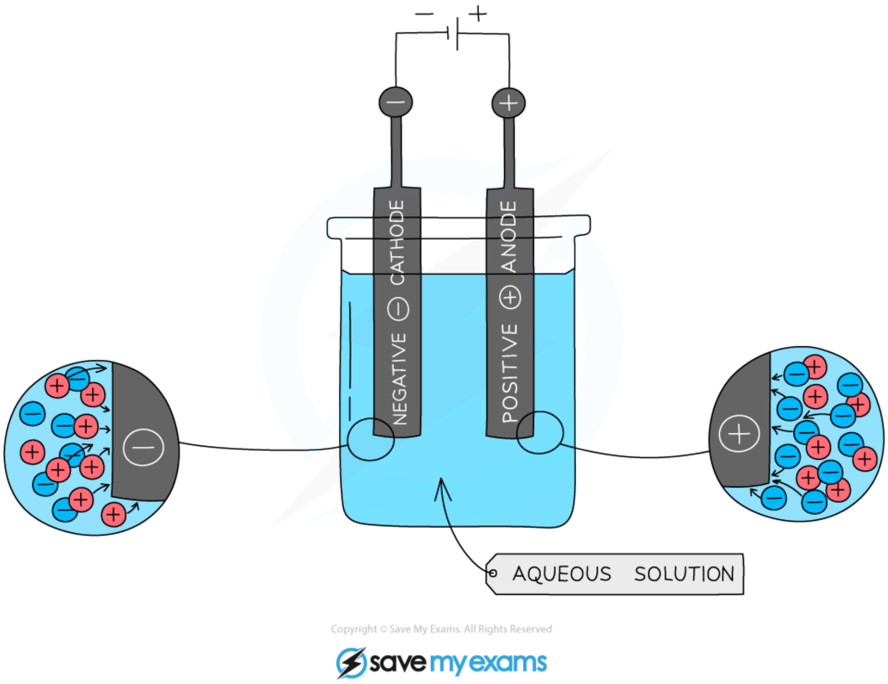


The polar water molecules will form ion-dipole bonds with the ions in solution causing the ions to become hydrated

Evidence

- The behaviour of ionic substances during electrolysis is a clear piece of evidence for the existence of ions
  - Positive ions in solution are attracted to the negative electrode
  - Negative ions in solution are attracted to the positive electrode

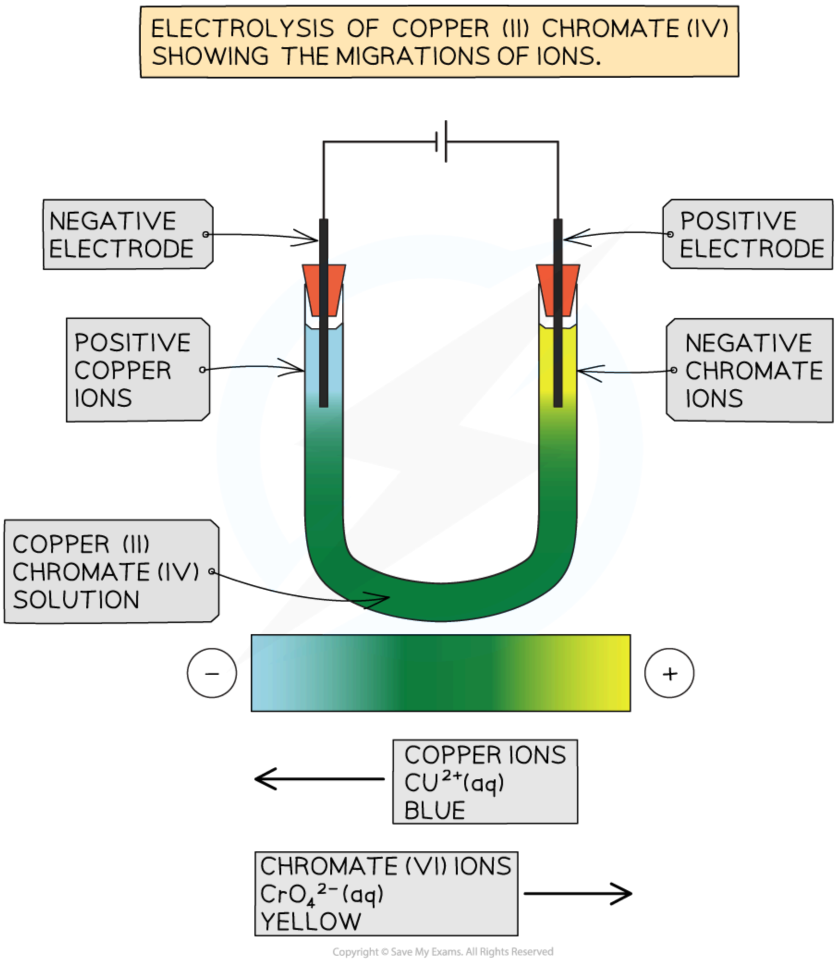
Your notes



Your notes

Set up of electrolysis

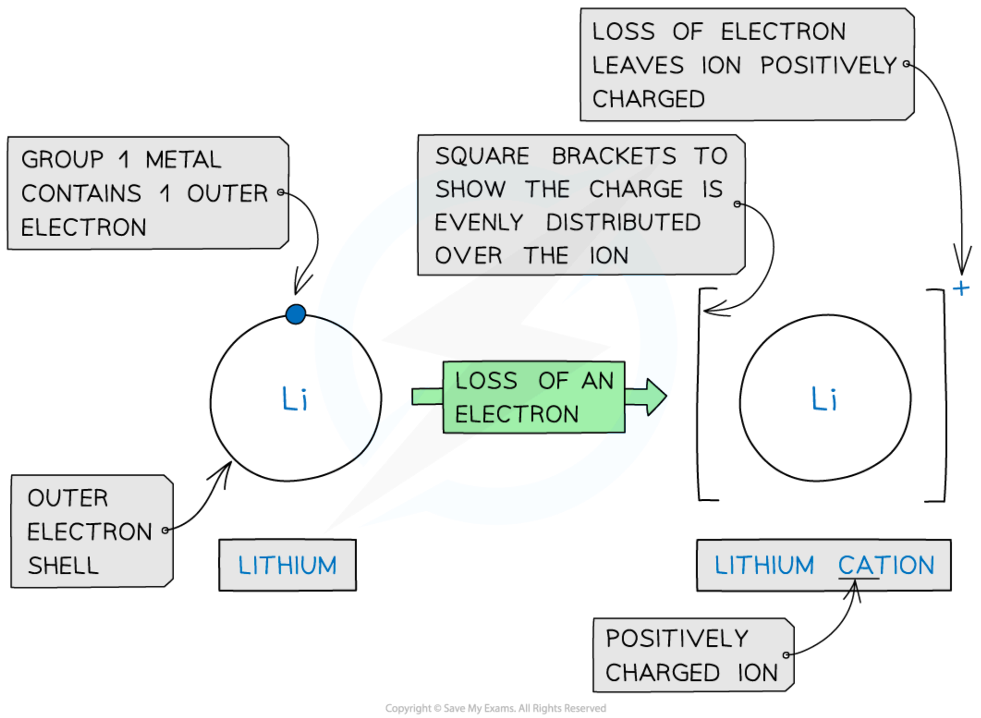
- An example which is simple to see the separation is using copper(II) chromate(VI), $CuCrO_4$
- The solution contains
  - **Cu^{2+} ions (blue)**
  - **CrO_4^{2-} ions (yellow)**



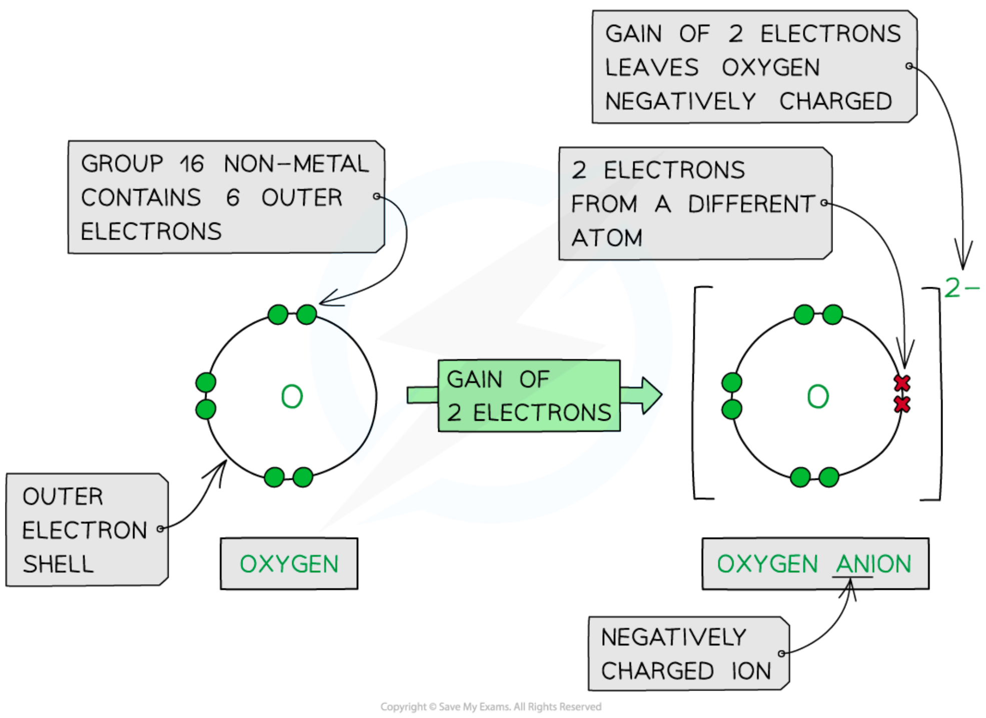
- Overall the solution is an olive green colour, but when the solution undergoes electrolysis a blue colour appears around the negative electrode, and a yellow colour appears around the positive electrode
- This is because the Cu^{2+} ions are attracted to the negative electrode and their blue colour is observed and the CrO_4^{2-} ions are attracted to the positive electrode and their yellow colour is observed

Ionic Dot-and-Cross Diagrams

- As a general rule, **metals** are on the **left** of the periodic table and **nonmetals** are on the **right-hand** side
- **Ionic bonding** involves the **transfer** of electrons from a **metallic** element to a **non-metallic** element
- Transferring electrons usually leaves the metal and the non-metal with a **full outer shell**
- Metals **lose** electrons from their valence shell forming **positively charged cations**
- Non-metal atoms **gain** electrons forming **negatively charged anions**
- Once the atoms become ions, their electronic configurations are the same as a noble gas
    - A potassium ion (K^+) has the same electronic configuration as argon: $[2,8,8]^+$
    - A chloride ion (Cl^-) also has the same electronic configuration as argon: $[2,8,8]^-$



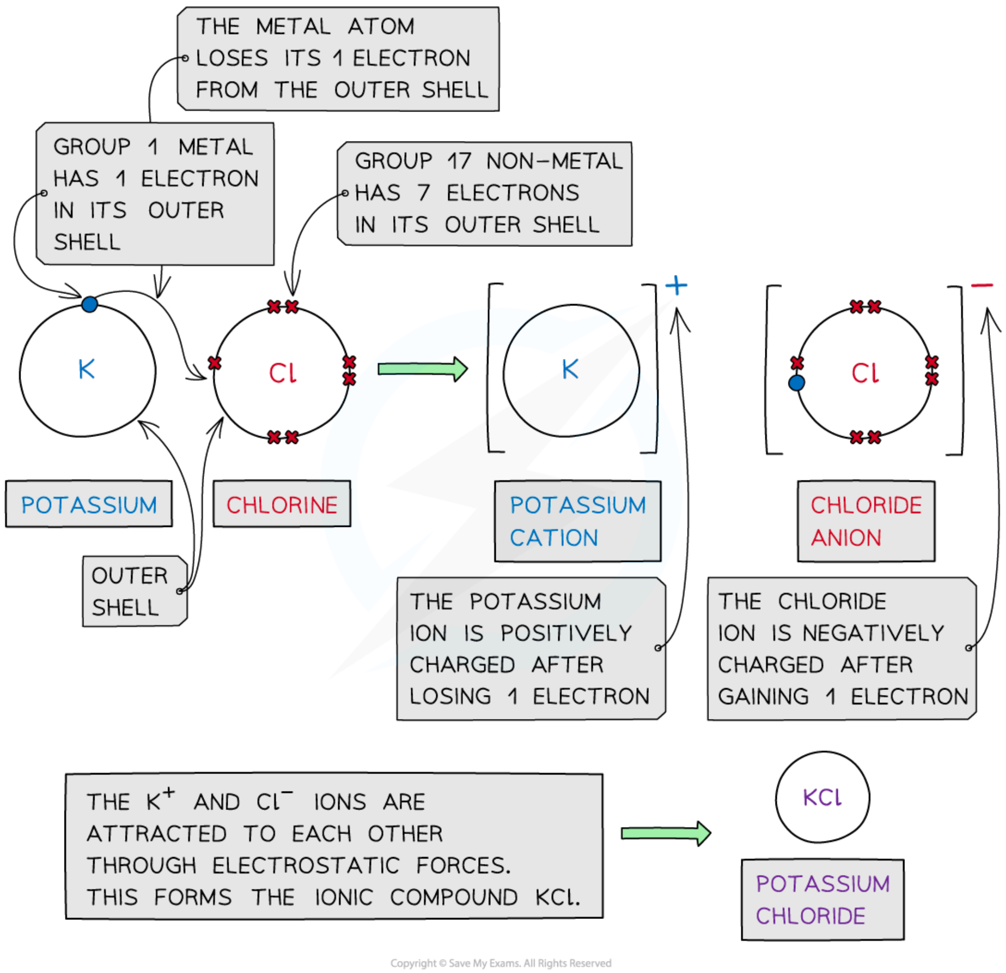
Forming cations by the removal of electrons from metals



Forming anions by the addition of electrons to nonmetals

- **Cations** and **anions** are oppositely charged and therefore attracted to each other
- **Electrostatic attractions** are formed between the oppositely charged ions to form **ionic compounds**
- The **ionic bond** is the **electrostatic attraction** formed **between the oppositely charged ions**, which occurs in all directions (this called **non-directional bonding**)
- This form of attraction is **very strong** and requires a **lot of energy** to overcome
  - This causes high melting points in ionic compounds

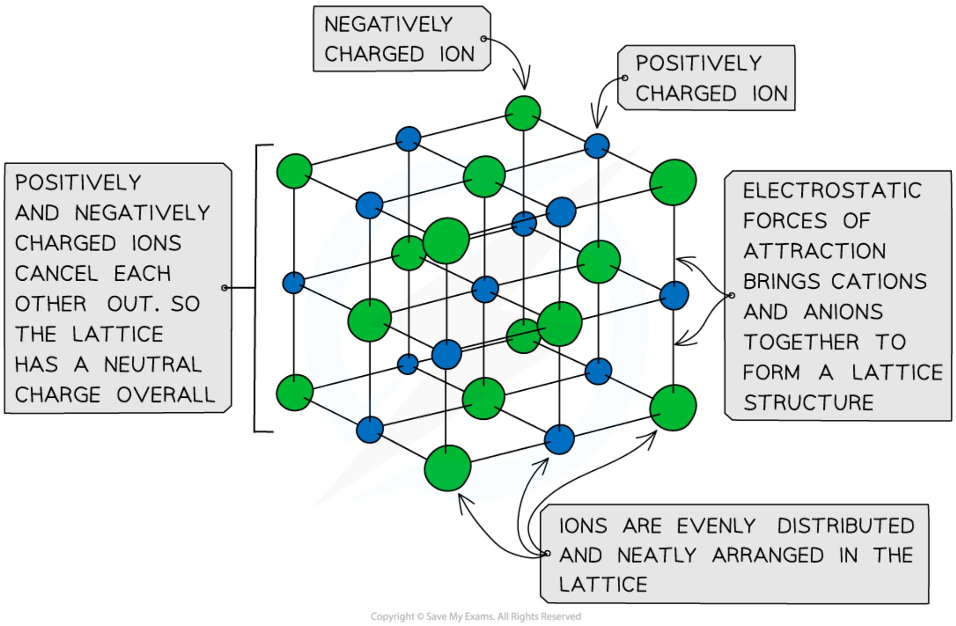
Your notes



Cations and anions bond together using strong electrostatic forces, which require a lot of energy to overcome

- The ions form a **lattice structure** which is an evenly distributed **crystalline** structure
- Ions in a lattice are arranged in a **regular repeating pattern** so that positive charges cancel out negative charges
- The attraction between the cations and anions is occurring in all directions
  - Each ion is attracted to all of the oppositely charged ions around it
- Therefore the final lattice is overall electrically **neutral**

Your notes



Ionic solids are arranged in lattice structures

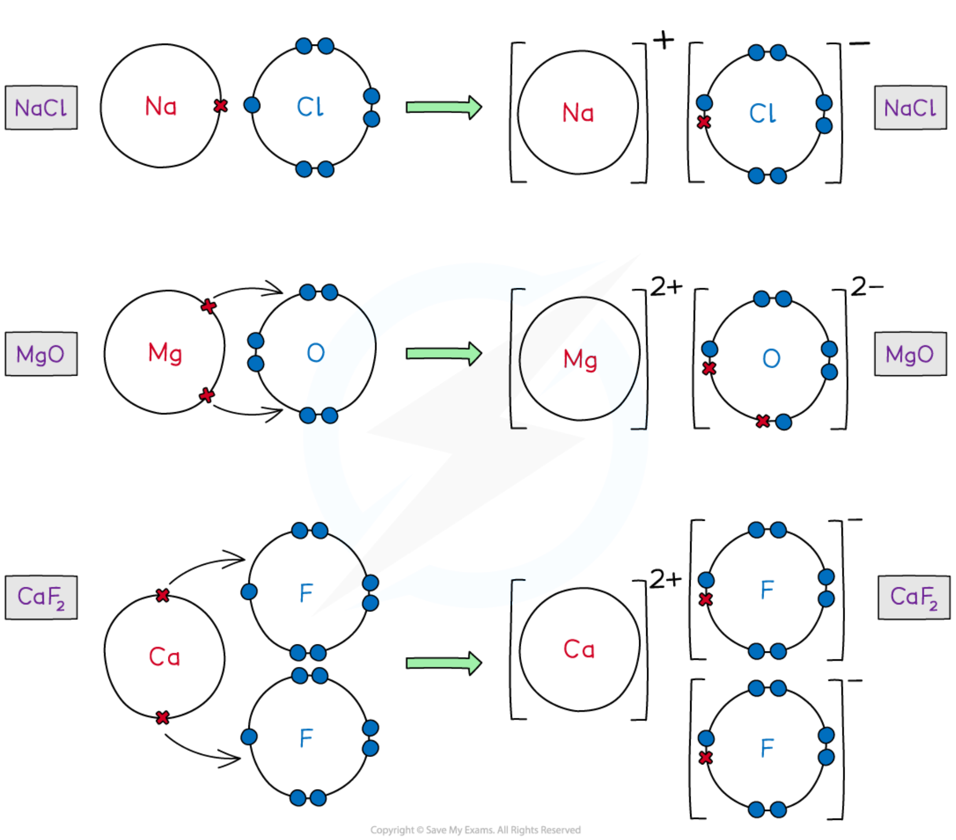
Dot and cross diagrams

- These are diagrams that show the arrangement of the outer-shell electrons in an **ionic** or **covalent** compound or element
- The electrons are shown as dots and crosses
- In a dot and cross diagram:
  - Only the outer electrons are shown
  - The charge of the ion is spread evenly which is shown by using brackets
  - The charge on each ion is written at the top right-hand corner

Ionic compounds

- Ionic bonds are formed when **metal atoms** transfer electrons to a **non-metal** to form a positively charged and negatively charged ion
- The atoms achieve a **noble gas** configuration

Your notes



Dot-and-cross diagrams of ionic compounds in which one of the atoms transfers their valence electrons to the other

Calcium fluoride

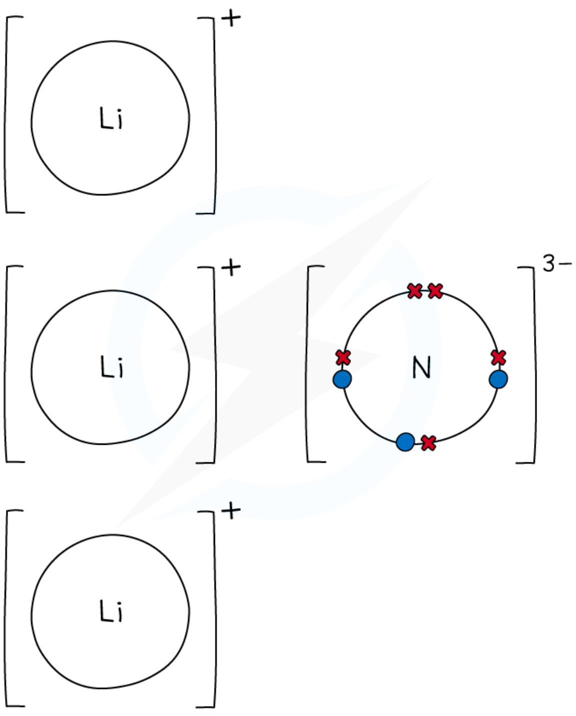
- Calcium is a Group 2 **metal**
  - It **loses** its 2 outer electrons to form a calcium ion with a +2 charge (Ca^{2+})
- Fluorine is a Group 7 **non-metal**
  - It **gains** 1 electron to form a fluoride ion with a -1 charge (F^-)
- As before, the positive and negative ions are attracted to each other via an ionic bond
- However, to cancel out the 2+ charge of the calcium ion, 2 fluorine atoms are needed
  - Each fluorine atom can only accept 1 electron from the calcium atom
  - 2 fluoride ions will be formed
- Calcium fluoride is made when 1 calcium ion and 2 fluoride ions form ionic bonds, CaF_2
- The final ionic solid of CaF_2 is **neutral** in charge

Worked Example

Draw a dot cross diagram for lithium nitride

Answer

- Lithium is a Group 1 **metal**
    - It **loses** its outer electron to form a lithium ion with a +1 charge (Li^+)
    - Nitrogen is a Group 5 **non-metal**
    - It **gains** 3 electrons to form a nitride ion with a -3 charge (N^{3-})
    - To cancel out the -3 charge of the nitride ion, 3 lithium atoms are needed and 3 lithium ions will be formed
    - Lithium nitride is made when 1 nitride ion and 3 lithium ions form ionic bonds
    - The final ionic solid of Li_3N is **neutral** in charge



Dot and cross diagram to show the ionic bonding in lithium nitride

Worked Example

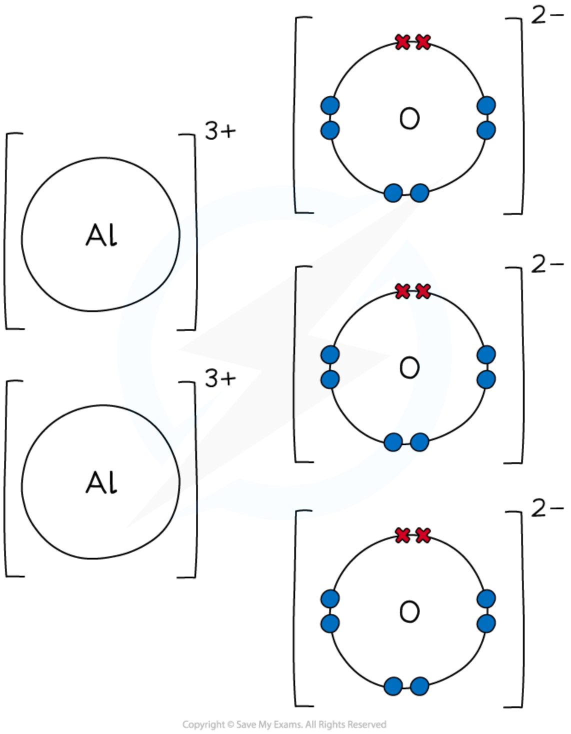
Draw a dot cross diagram for aluminium oxide

Answer

- Aluminium is a Group 3 **metal**
    - It **loses** its outer electrons to form an aluminium ion with a +3 charge (Al^{3+})
    - Oxygen is a Group 6 **non-metal**
    - It **gains** 2 electrons to form an oxide ion with a -2 charge (O^{2-})

Your notes

- To cancel out the negative and positive charges, 2 aluminium and 3 oxygen atoms are needed
- Aluminium oxide is made when 2 aluminium ions and 3 oxygen ions form ionic bonds
- The final ionic solid of Al_2O_3 is **neutral** in charge

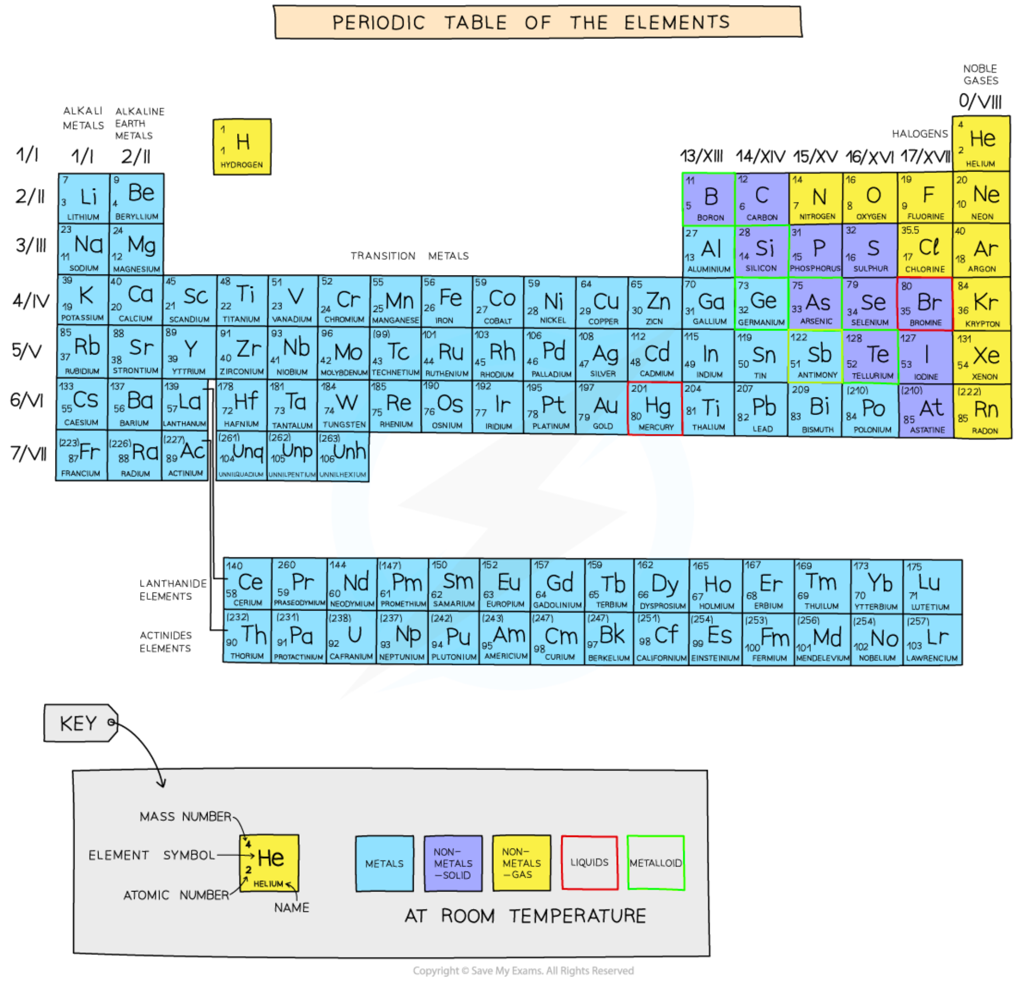


Dot and cross diagram to show the ionic bonding in aluminium oxide

Your notes

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**



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

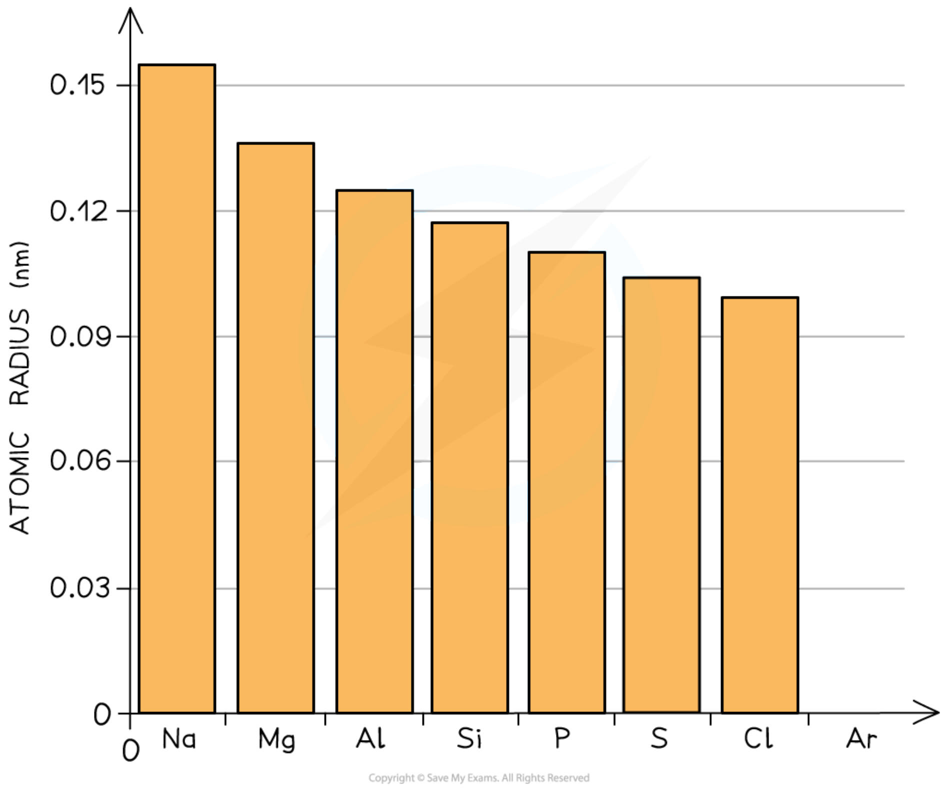
- In **metals** this is also called the **metallic radius** and in **non-metals**, the **covalent radius**

Your notes

Atomic radii of Period 3 elements

- You can see a clear trend across the period:

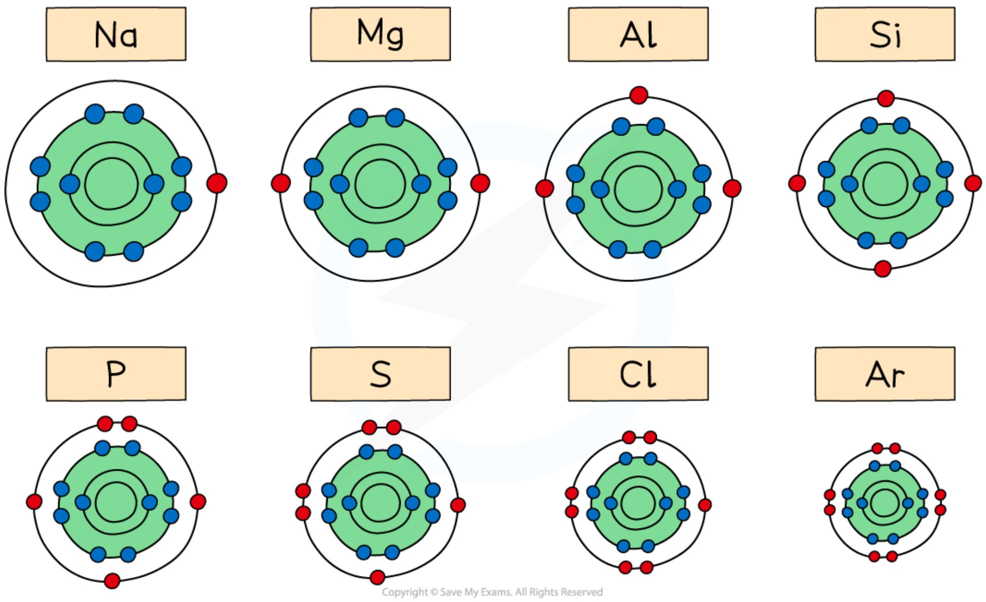
Period 3 Element	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic Radius (nm)	0.157	0.136	0.125	0.117	0.110	0.104	0.099	No data



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



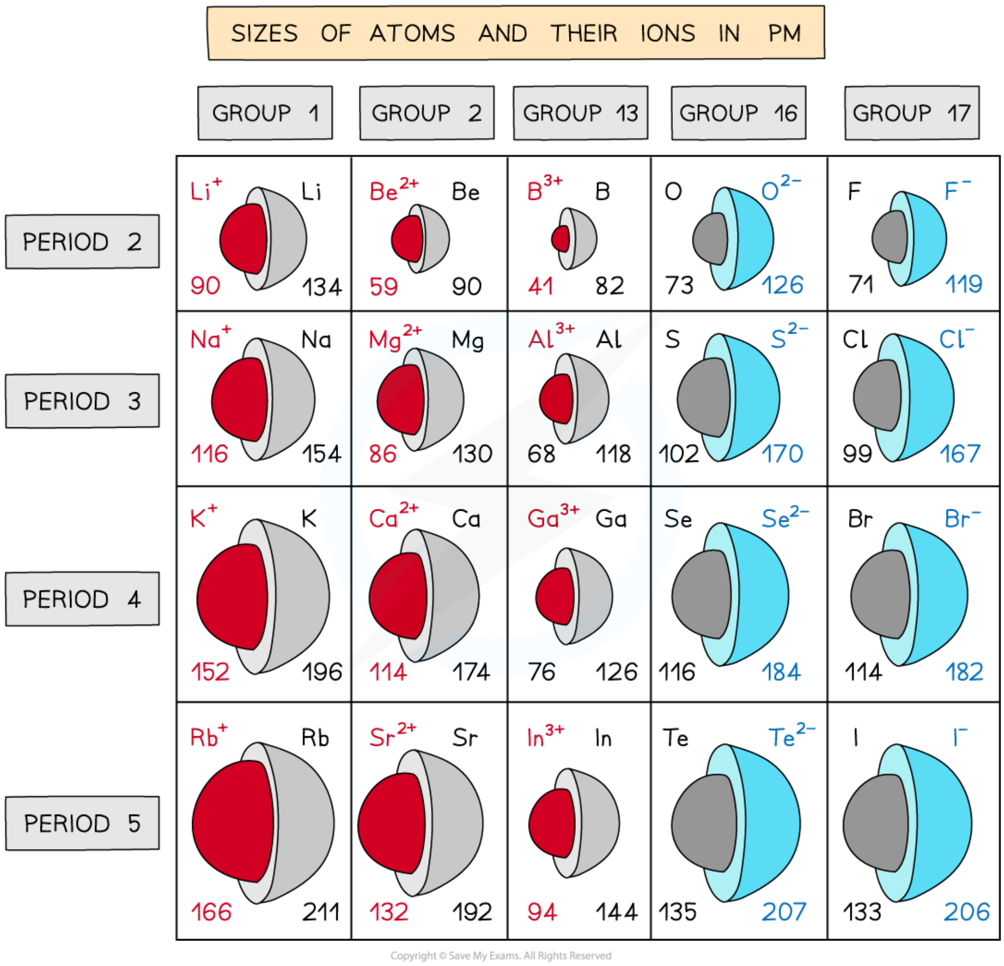
The diagram shows that across Period 3, the elements gain extra electrons in the same principal quantum shell

Ionic Radius & Charge

Ionic radius

- The **ionic radius** of an element is a measure of the size of an ion
- Ionic radii show predictable patterns
  - Ionic radii **increase** with increasing **negative charge**
  - Ionic radii **decrease** with increasing **positive charge**
- For negative ions
  - They are formed by atoms **gaining** electrons
  - The outermost electrons are further away from the positively charged nucleus and are therefore held only weakly to the nucleus which **increases the ionic radius**
  - The greater the negative charge, the larger the ionic radius
- For positive ions
  - Positively charged ions are formed by atoms **losing** electrons
  - The nuclear charge remains the same but there are now fewer electrons that undergo a greater electrostatic force of attraction to the nucleus which **decreases the ionic radius**
  - The greater the positive charger, the smaller the ionic radius

Your notes



Trends in the ionic radii across a period and down a group

- The strength of ionic bonding, therefore, depends on both the charges and the sizes of the two ions
- **Higher charges** and **smaller sizes** produce stronger bonds
- This is because:
  - As the ionic charge increases, or the ionic radius decreases, the charge density increases

Trends in Ionic Radius

Isoelectronic Ions

- These are ions that have the **same electronic configuration**
- For example the following ions all have the electronic configuration of **$1s^2$ $2s^2$ $2p^6$**
- However, they all have different numbers of protons, therefore will have a different ionic radius
  - N^{3-} (7 protons)

Your notes

- O^{2-} (8 protons)
- F^- (9 protons)
- Na^+ (11 protons)
- Mg^{2+} (12 protons)
- Al^{3+} (13 protons)

- As the number of protons in the nucleus of the ion increases, the electrons get pulled in more closely to the nucleus
- The radii of the isoelectronic ions therefore fall across this series of ions
  - N^{3-} has an ionic radius of 0.171 nm and Al^{3+} has an ionic radius of 0.054 nm

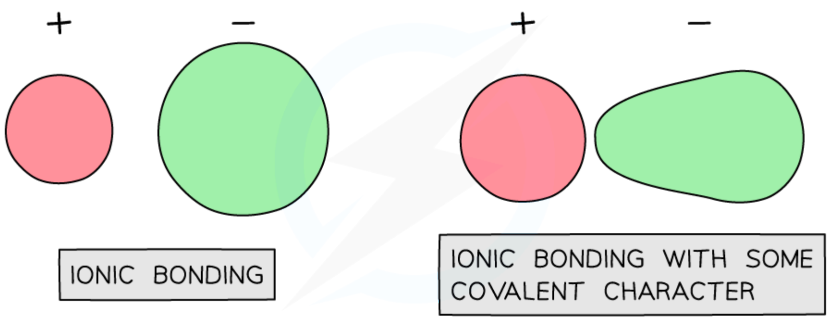
Your notes

Polarisation

Your notes

The Polarisation of anions & Cations

- Theoretical lattice energies assume a perfect ionic model where the ions are 100% spherical and the attractions are purely electrostatic
  - So overall, there has a been a complete and irreversible exchange of outer shell electrons between the two elements in the compound so that one ion is negatively charged and the other ion positively charged and the charges are whole values
- Theoretical lattice energy can differ from measured lattice energy
- This is because covalent character is introduced when there is **polarisation** of the anions
- When this occurs, the cation attracts electrons from the anion therefore distorting electron density of the anion



Polarisation of anion leading to covalent character

Cations

- A cation with a large charge density will have a larger **polarising power**
- The charge density can be calculated as the charge of the cation divided by the surface area, for example
  - K^+ will have a lower polarising power than Li^+ as the ionic radius is greater
  - Na^+ will have a lower polarising power than Mg^{2+} as the charge is smaller
- So a cation that is small and highly charged will have the greatest polarising power

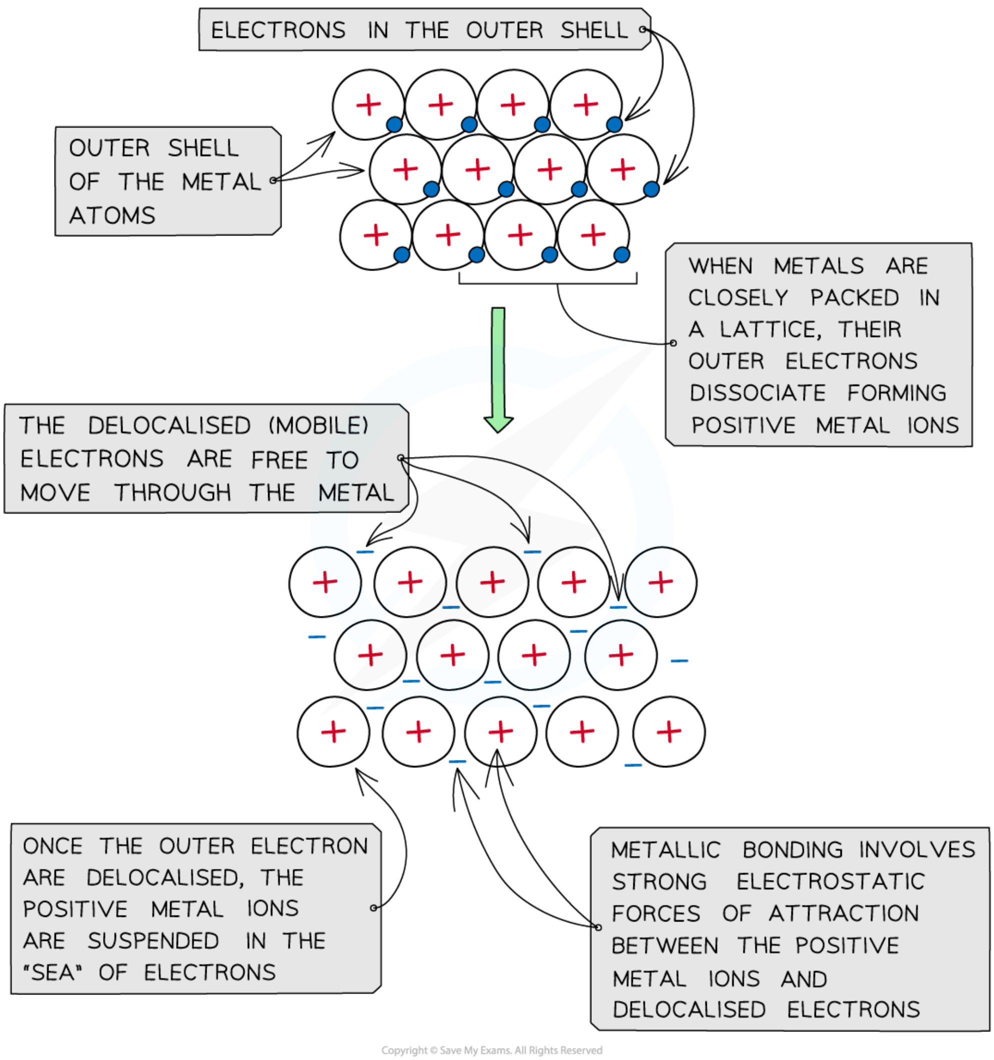
Anions

- Anions are polarised
- The polarisability of the anion depends of its ionic radius
- The larger the ionic radius the more easily it will be distorted, for example
  - Br- will be more polarisable than Cl^- as the ionic radius is larger

Your notes

Metallic Bonding & Lattices

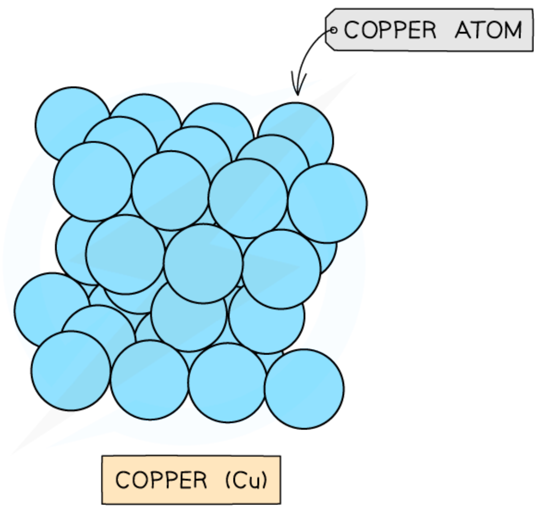
- **Metal** atoms are tightly packed together in **lattice** structures
- When the metal atoms are in lattice structures, the electrons in their outer shells are free to move throughout the structure
- The free-moving **electrons** are called '**delocalised electrons'** and they are not bound to their atom
- When the electrons are delocalised, the metal atoms become **positively** charged ions
- The positive ions **repel** each other
- The strong **forces** between the positive metal centres and the 'sea' of delocalised electrons holds the lattice together and keeps it stable



The positive metal ions are suspended in a 'sea' of delocalised electrons

Simple Properties of Metals

- Metals form **giant metallic lattices** in which the metal ions are surrounded by a 'sea' of **delocalised** electrons
- The metal ions are often packed in **hexagonal layers** or in a **cubic arrangement**
- This layered structure with the delocalised electrons gives a metal its key properties



Layers of copper ions (the delocalised electrons are not shown in the diagram)

- If other atoms are added to the metal structure, such as carbon atoms, this creates an **alloy**
- Alloys are much stronger than pure metals, because the other atoms stop the layers of metal ions sliding over each other easily
- The **strength** of the metallic attraction can be increased by:
  - **Increasing** the number of **delocalised electrons** per metal atom
  - **Increasing** the **positive charges** on the metal centres in the lattice
  - **Decreasing** the **size** of the metal ions
- Due to the delocalised 'sea' of electrons, metallic structures have some **characteristic properties** shown below:

Metallic Bonding Properties Table

Property	Reason

Your notes

Your notes

High melting and boiling point	There are strong electrostatic forces of attraction between the delocalised electrons and positive metal ions which results in strong metallic bonds within giant metallic structures. These require large amounts of energy to overcome the forces and break the bonds
Good conductors of electricity	The delocalised electrons are able to move and carry charge. For example, in a circuit, electrons entering one end of the metal cause a delocalised electron to be displaced from the other need. Hence, electrons can flow and conduct electricity
Good conductors of heat	The delocalised electrons are free to move within the metal structure. The conversion of the kinetic energy of the electron to heat energy when it collides with the metal atoms/nuclei results in the transfer of energy. This is alongside the increased vibrations and movement of the particles (positive ions and delocalised electrons)
Malleable and ductile	Layers of positive ions can easily slide over one another to take up different positions. This does not disrupt the metallic bonding as the delocalised electrons do not belong to any particular metal atom and so they can move with the layers of positive ions, maintaining the electrostatic forces. Therefore, the metallic bonds are not broken and, as a result, metallic bonds are strong but flexible/

Examiner Tips and Tricks

You should be able to draw the structure of a metal with positive ions in layers and the delocalised electrons surrounding the ions

If drawing the structure of a metal in the exam, make sure to include labels for metal ions and delocalised electrons