2.1 Bonding, Structure & Properties

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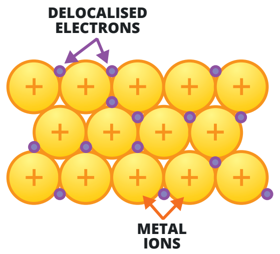
Metallic Bonding:

- **Metallic bonding** - when **metal atoms** bond together.
- Metals have **giant structures** of **regularly arranged atomsr**

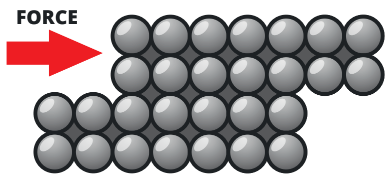


The electrons from the outer shells of the atoms are **delocalised** *– meaning they are* **free to move** *through the whole structure.*

- By **sharing delocalised electrons - strong metallic bonds** are formed.
- The strength of a metallic bond is due to the force of attraction between the metal ions (+) and the delocalised electrons (-).

Metallic Properties

- **Conduct electricity** - the <u>delocalised electrons carry</u> electrical <u>charge</u> through the structure.
- **Conduct heat** – the delocalised electrons and closely packed ions <u>transfer energy through the structure</u> by conduction.
- **Malleable** and **ductile**



The <u>layers of metal ions</u> are able to <u>slide over each other</u> when hammered or stretched

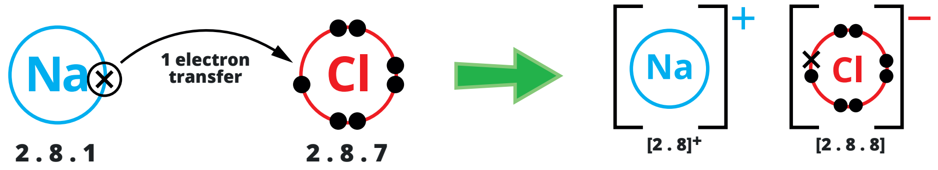
- **High melting** and **boiling points** – <u>large amounts of energy</u> are needed to <u>break the strong metallic bonds</u> in melting / boiling.

Higher Tier – the melting and boiling points increase as you move across any period of the Periodic table, because there are more delocalised electrons increasing the attraction between the ions and the free electrons (stronger bonds).

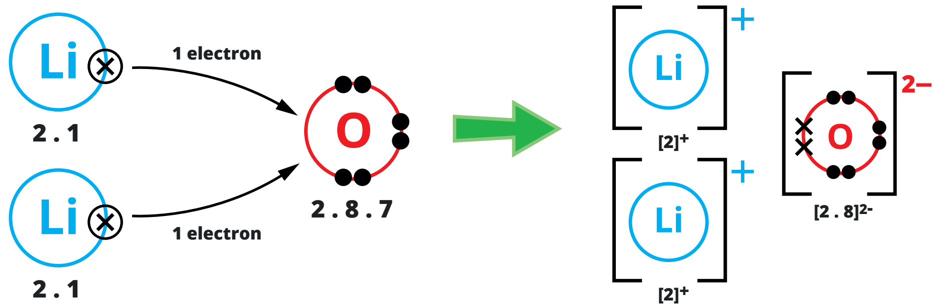
Ionic Bonding:

- **Ion** – a charged particle. Has different numbers of protons and electrons.
  - » **Positive ion** – has more protons (+) than electrons (-)
  - » **Negative ion** – has more electrons (-) than protons (+)
- **Ionic bonding** – the bonding between **metal and non-metal atoms**.
- Ionic bonds form when **electrons transfer** from a metal to a non-metal atom so that **both atoms** achieve **full outer shells**.
- **Dot and Cross** diagram – representation to show ionic bonding.

sodium chloride



lithium oxide

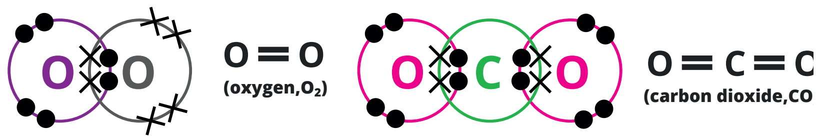


Covalent Bonding:

- **Covalent bonding** – the bonding between non-metal atoms.
- Covalent bonds form when the atoms **share electrons** so that **both atoms** achieve **full outer shells**.
- **Dot and Cross** diagram – modelling used to show covalent bonding:

Single bonds – 1 pair of electrons shared

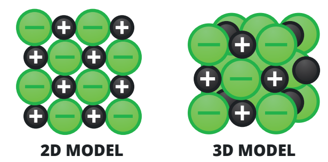
Double bonds – 2 pairs of electrons shared (HT only)



Simple and Giant Structures:

Giant Ionic Structures

- Ionic compounds have giant ionic structures.
- **Ionic lattice** – the **regular arrangement** of the ions in ionic structures.



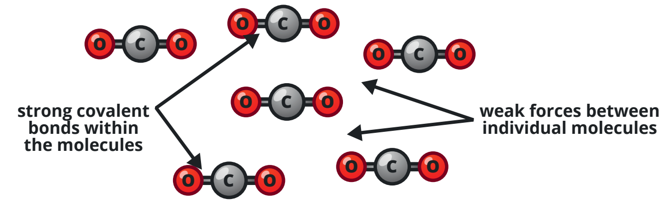
The oppositely charged ***ions attract*** *each other in a* ***regular pattern****.*

Properties:

- **High melting and boiling points** – due to the **strength** of the **electrostatic forces between** the ions.
- Conduct electricity when dissolved or molten – only then are the ions free to move to carry the charge.

Simple Molecular Structures

- Simple molecules consist of a few atoms held together by covalent bonds.
- Hydrogen, water and carbon dioxide are examples of simple molecular structures.

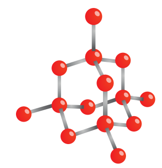


Properties

- **Low melting and boiling points** – due to the **weak intermolecular forces between** the molecules.
- **Do not conduct electricity** – no free electrons to carry the electric current.

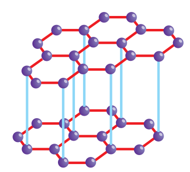
Giant Covalent Structures:

- **Giant covalent structures** consist of **lots of atoms** held together by covalent bonds
- They are arranged into **giant lattices**, which are **extremely strong** because of the large number of bonds in the structure.



Diamond - each carbon **bonded to 4 others**.

- <u>Does not conduct</u> electricity.
- Used in <u>drill bits</u>, <u>glass cutting</u>, gemstones.



Graphite - each carbon **bonded to 3 others**.

- <u>Conducts</u> electricity – delocalised electrons between layers carries charge.
- Used in <u>pencils</u> and <u>lubricants</u> – <u>layers can slide over</u> each other.