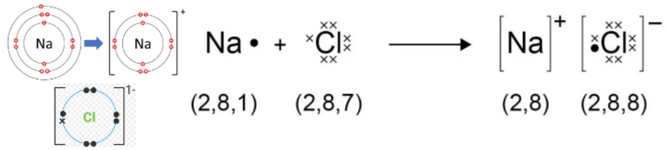
# Bonding, structure and the properties of matter

In chemical equations, the three states of matter are shown as (s), (l) and (g), with (aq) for aqueous solutions.

## Ionic bonding (metal to non-metal): the particles are oppositely charged ions.

There are **strong electrostatic attraction between oppositely charged ions**.

Electrons in the outer shell of the metal atom are transferred. **Metal atoms lose electrons to become positively charged ions. Non-metal atoms gain electrons to become negatively charged ions**. The ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 have the electronic structure of a noble gas (Group 0). The electron transfer during the formation of an ionic compound can be represented by a dot and cross diagram, eg for sodium chloride




Dot and cross – doesn't show how the ions are arranged in terms of space.

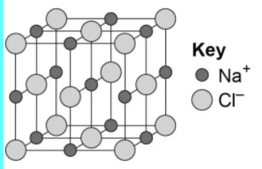
**The sodium atom loses 1 electron to become a 1+ ion to achieve a full outer shell.**
**The chlorine atom gains 1 electron to become a 1- chloride ion to achieve a full outer shell.**

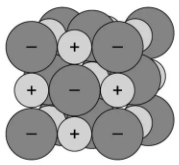
Ionic compounds have regular structures (giant ionic lattices) in which there are **strong electrostatic forces of attraction in all directions between oppositely charged ions.**

These compounds have **high melting points and high boiling points** because of the large amounts of energy needed to break the many strong bonds.

When **melted or dissolved in water, ionic compounds conduct electricity** because **the ions are free to move and so charge can flow**.

Ball and stick- doesn't show that forces of attraction between ions actually act in all directions.



If a diagram has alternating charged particles it is ionic. The ratios of ions are its empirical formula. Eg NaCl (1 to 1)

2 dimensional models – only show one layer, not other layers.

3- dimensional models – not to scale, gives no information about forces of attraction between ions or movement of electrons to form ions.

## Covalent bonding (non-metal to non-metal): atoms share pairs of electrons, they form covalent bonds.

These bonds **between atoms** are **strong**.

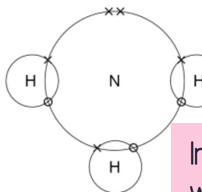
Covalently bonded substances may consist of **small molecules**, such as $H_2O$, $CO_2$

Some covalently bonded substances have **very large molecules**, such as polymers.

Some covalently bonded substances have **giant covalent structures**, such as diamond and silicon dioxide.

The covalent bonds in molecules and giant structures can be represented in the following forms:

**For ammonia ($NH_3$)**

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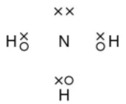


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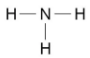
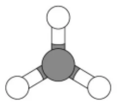
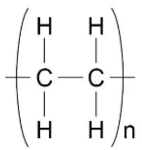


Substances that consist of small molecules are usually gases or liquids that have relatively low melting points and boiling points.

**Intermolecular** forces are **weak compared** with **covalent bonds**

These substances have only weak forces between the molecules (intermolecular forces). It is these intermolecular forces that are overcome, not the covalent bonds, when the substance melts or boils.



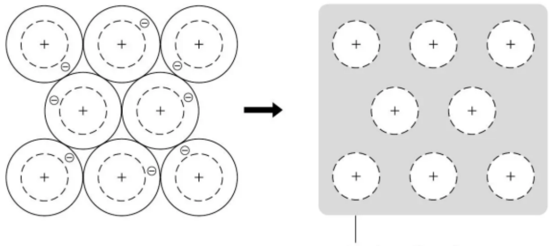

poly(ethene)

Polymers have very large molecules. The atoms in the polymer molecules are linked to other atoms by strong covalent bonds. The intermolecular forces between polymer molecules are relatively strong and so these substances are solids at room temperature.

## Metallic bonding (metal to metal): Metals consist of giant structures of atoms arranged in a regular pattern.

(most metals have high melting and boiling points).

The electrons in the outer shell of metal atoms are **delocalised** and so are free to move through the whole structure. **The sharing of delocalised electrons gives rise to strong metallic bonds**. The bonding in metals may be represented in the following form:




Metals are good conductors of electricity because the **delocalised electrons in the metal carry electrical charge** through the metal. Metals are good conductors of thermal energy because **energy is transferred by the delocalised electrons**

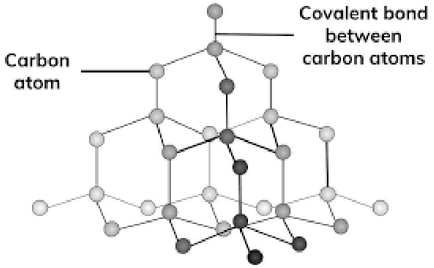
In pure metals, atoms are arranged in layers, which allows metals to be bent and shaped as the layers slide easily.


Alloys are harder than pure metals due to distortion of the layers of atoms in the structure of a pure metal by the other metal atoms.

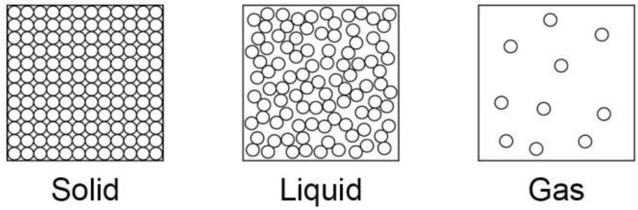
# Giant covalent carbon structures

Substances that consist of giant covalent structures are solids with **very high melting points**. All of the atoms in these structures are linked to other atoms by **strong covalent bonds. These bonds must be overcome** to melt or boil these substances. **Diamond** and **graphite** (forms of carbon) and **silicon dioxide** (silica) are examples of giant covalent structures.

DIAMOND: In diamond, each carbon atom forms **four covalent bonds** with other carbon atoms in a giant covalent structure, so diamond is **very hard**, has a **very high melting point** and **does not conduct** electricity.

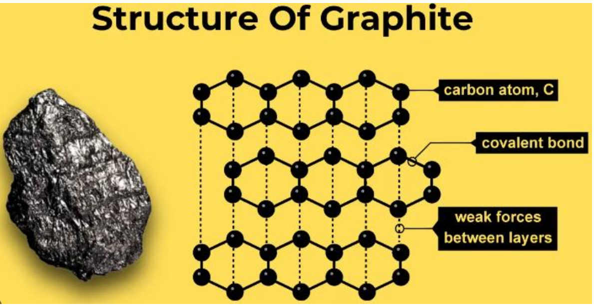



The intermolecular forces increase with the size of the molecules, so larger molecules have higher melting and boiling points. These substances do not conduct electricity because the molecules do not have an overall electric charge. BUT, graphite does!




In this model there are no forces, all particles are represented as spheres and the spheres are solid.

The amount of energy needed to change state from solid to liquid and from liquid to gas depends on the strength of the forces between the particles of the substance. The nature of the particles involved depends on the type of bonding and the structure of the substance. The stronger the forces between the particles the higher the melting point and boiling point of the substance.

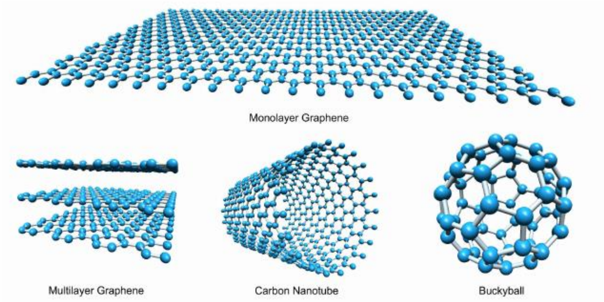



In graphite, one electron from each carbon atom is delocalised. This means it can conduct electricity.

In GRAPHITE, each carbon atom forms **three covalent bonds with three other carbon atoms**, forming **layers of hexagonal rings** which have no covalent bonds between the layers.

Key Terms to Remember:

- **Ionic bond**: Between metal and non-metal, electrostatic force of attraction between ions that are charged (positive and negative) due to gaining or losing electrons.
- **Covalent bond**: Between non-metal atoms. Electrons are shared. Very strong!
- **Intermolecular**: Between molecules (not the bonds IN the molecule).
- .**Metallic**: Between metals – sharing delocalised electrons.
- **Delocalised electrons**: Electrons that flow and can carry charge/current.
- **Charge**: Electrons have a negative charge, metals lose electrons to carry a positive charge when an ion, non-metals gain electrons and have a negative charge when an ion.
- **Giant covalent**: Very large covalent structure like diamond, graphite, silicon dioxide
- **Electrostatic force of attraction**: The attractive force acting in ionic bonding.




Graphene is a single layer of **graphite** and has properties that make it useful in **electronics and composites**.

Fullerenes are molecules of carbon atoms with **hollow shapes**. The structure of fullerenes is based on hexagonal rings of carbon atoms but they may also contain rings with five or seven carbon atoms. The first fullerene to be discovered was Buckminsterfullerene ($C_{60}$) which has a spherical shape.

Carbon **nanotubes are cylindrical fullerenes** with very high length to diameter ratios. Their properties make them useful for **nanotechnology, electronics and materials**