

Edexcel Chemistry A-Level

Topic 2: Bonding and Structure Detailed Notes





Topic 2A: Bonding

Ionic Bonding

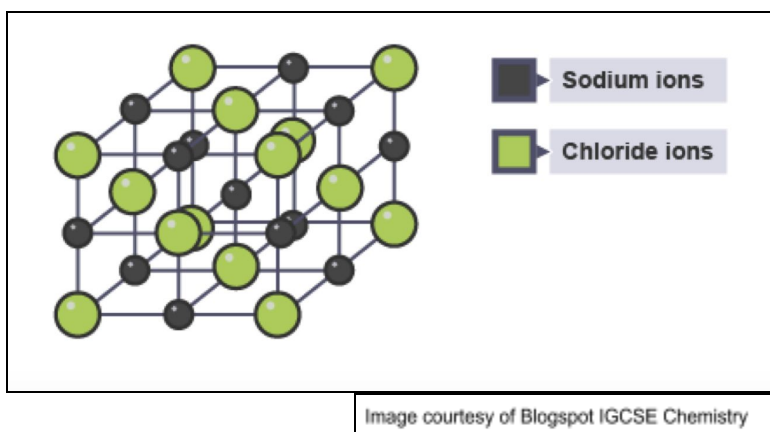
Ionic bonding occurs between a **metal and a non-metal**. Electrons are **transferred** from the metal to the non-metal to achieve full outer shells.

When the electrons are transferred, it creates **charged particles** called **ions**. Oppositely charged ions **attract through electrostatic forces** to form a **giant ionic lattice**.

Example:

Sodium Chloride is an ionic compound formed from Na^+ and Cl^- ions. Sodium loses an electron and Chlorine gains an electron to produce ions with a full outer electron shell.

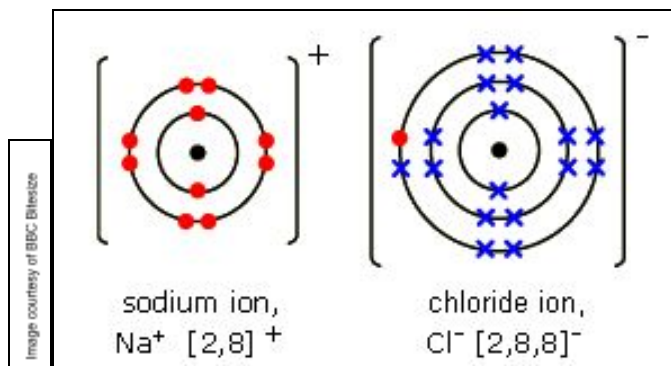
These then attract into an ionic lattice:



The **charge** of an ion is related to the strength of the ionic bond that forms. Ions with a **greater charge** will have a **greater attraction** to the other ions resulting in stronger forces of attraction and therefore **stronger ionic bonding**.

Larger ions that have a **greater ionic radius** will have a **weaker attraction** to the oppositely charged ion because the attractive forces have to act over a **greater distance**.

Cations (+ve) and **Anions** (-ve) can be represented using **dot and cross diagrams** and so can ionic bonding. The electrons being transferred from the cation can be seen on the outer shell of the anion.



The red dot clearly shows the transferred electron from sodium to chloride to produce two ions with full outer electron shells.



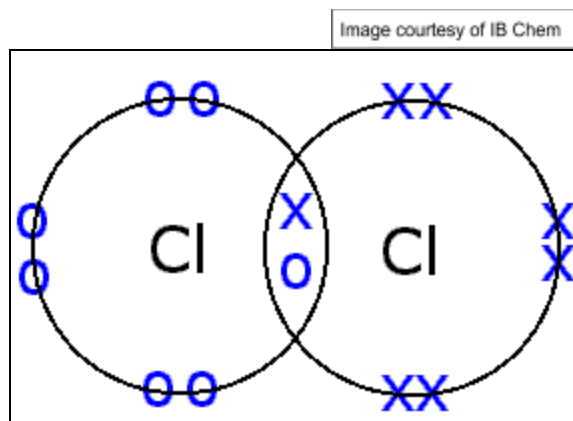


Covalent Bonding

Covalent bonds form between **two non-metals**. Electrons are **shared** between the two outer shells in order to achieve a **full outer shell**. **Multiple electron pairs** can be shared to produce **multiple covalent bonds**.

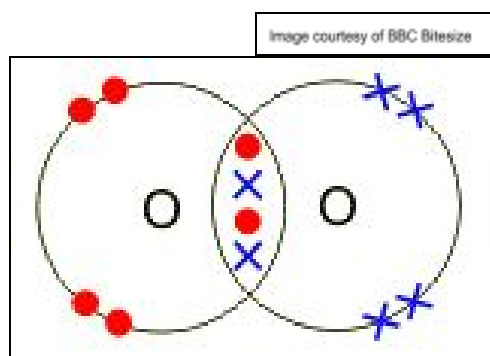
The shared electron pairs can be represented using **dot and cross diagrams** and a covalent bond shown with a straight line.

Example:



Double and triple bonds can also be shown on dot and cross diagrams with the multiple electron pairs being displayed in the shared segment between the two atoms.

Example:



The **length** of a covalent bond is strongly linked to its **strength**. **Shorter bonds** tend to be **stronger** as the atoms are **held closer together** so the forces of attraction are greater, requiring more energy to be overcome. Double and triple bonds are shorter than single covalent bonds explaining why they are so much stronger.



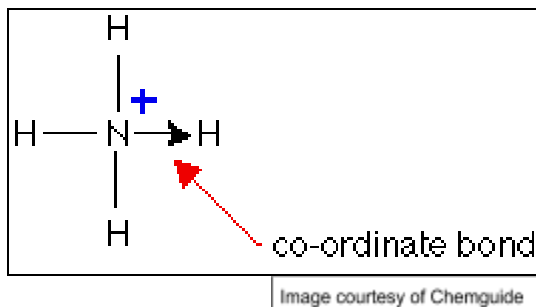


Dative Bonding

Dative or coordinate bonds form when both of the **electrons in the shared pair** are supplied from a **single atom**. It is indicated using an **arrow** from the lone electron pair.

Example:

Ammonia (NH_3) has a lone electron pair that can form a dative bond with a H^+ ion to produce an ammonium ion (NH_4^+).

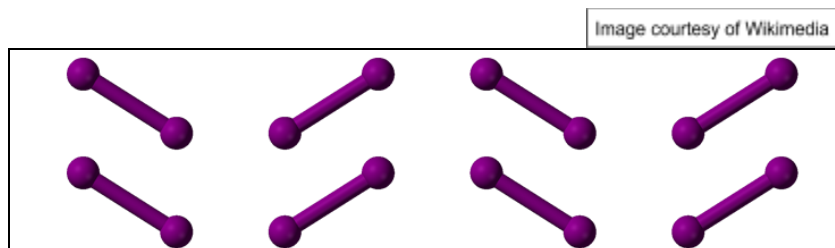


Once a dative bond has formed, it is treated as a **standard covalent bond** as it **reacts in exactly the same way** and has the same properties regarding length and strength.

Simple Covalent

Substances with a simple molecular structure consist of **covalently bonded molecules** held together with weak **van der waals** forces. These are a type of intermolecular force that act between the molecules holding them in a structure.

Example:



Shapes of Simple Molecules

The shape of a simple molecule or ion is determined by the number of **electron pairs around the central atom** and the repulsion between them. Each electron pair **naturally repels** each other so that the **largest bond angle possible** exists between the covalent bonds.





Lone Pair Repulsion

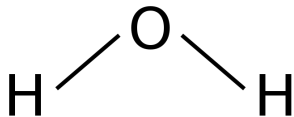
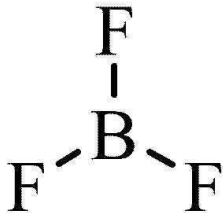
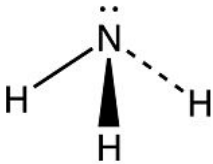
Any lone pairs present around the central atom provide **additional repulsive forces**, which changes the bond angle. For every lone pair present, the bond angle between covalent bonds is **reduced by 2.5°**.

Molecule Shapes

The shape of a molecule can be determined by considering the **type and quantity of electron pairs**:

1. Find the number of electron pairs.
2. Determine how many of the pairs are bonding pairs and how many are lone pairs.
3. Bonding pairs indicate the basic shape and lone pairs indicate any additional repulsion.

This table shows some common molecule shapes:

Name	Bonding e ⁻ Pairs	Lone e ⁻ Pairs	Bond Angle (°)	Example
Linear	2	0	180	Cl - Be - Cl
V - Shaped	2	2	104.5	
Trigonal Planar	3	0	120	
Triangular Pyramid	3	1	107	





Tetrahedral	4	0	109.5	
Trigonal Bipyramid	5	0	180 and 120	
Octahedral	6	0	90	

Images courtesy of World of Chemicals, Socratic, Quora, and Alchetron

Bond Polarity

The negative charge around a covalent bond is **not evenly spread** around the orbitals of the bonded atoms.

Electronegativity

Every atom has electronegativity, which is defined as:

The power of an atom to attract negative charge towards itself within a covalent bond.

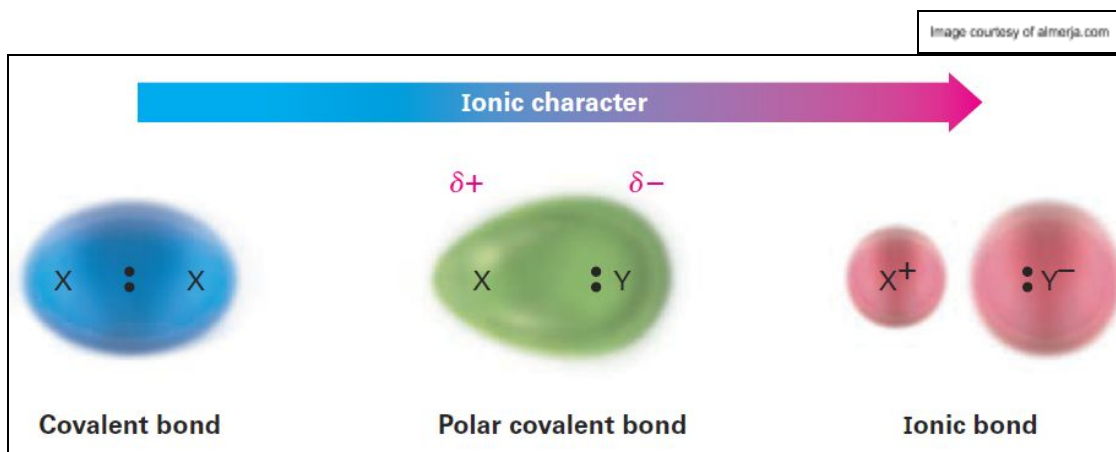




This 'power' is different for every atom depending on its **size and nuclear charge**.
Electronegativity **increases along a period** as atomic radius decreases and **decreases down a group** as shielding increases.

Ionic and covalent bonding are the extremes in a continuous scale of bonding as shown below.

Example:

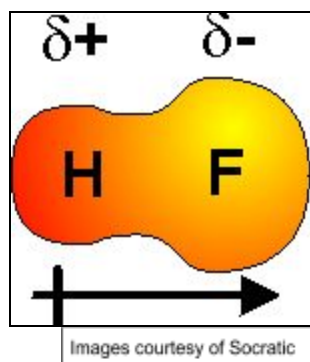


This bond polarity can be permanent or induced depending on the molecule and how it interacts with things around it.

Permanent Dipole

If the two atoms that are bonded have different electronegativities, a **polar bond** forms. The more electronegative atom **draws more of the negative charge towards itself** and away from the other atom, producing a δ^- region and a δ^+ region. This is a **permanent dipole**.

Example:



Hydrogen Fluoride is a polar molecule as Fluorine is a lot more electronegative than hydrogen so electrons are drawn to the right.

Polar molecules with a permanent dipole can align to form a **lattice of molecules** similar to an ionic lattice.

Induced Dipole

Before





An induced dipole can form when the electron orbitals around a molecule are **influenced by another charged particle**.

Example:

Intermolecular Forces

There are **three main types of intermolecular force**. Each one differs in strength and in what they act between.

Van der Waals Forces

This is the **weakest** type of intermolecular force. It acts as an **induced dipole** between molecules.

The strength of van der waals forces varies depending on the Mr of the molecule and its shape. The **larger the Mr** of the molecule, the **stronger the intermolecular forces**. Straight chain molecules experience stronger van der waals forces than branched chain molecules as they can **pack closer together**. This **reduces the distance** over which the force acts making it stronger.

Van der waals forces act between organic **alkane chains** and are affected by the **chain length** and any **branching**. As the chain length of the alkane increases, so does the **Mr** of the molecule. This results in **stronger** intermolecular forces between the chains and the compound has a **higher boiling point** as a result.

Branching of alkane chains weakens van der waal forces between the chains as they are less able to **pack tightly** together. Therefore the distance over which the intermolecular forces act is increased and the **attractive forces weakened**. This means branched chain alkanes have **lower boiling points** than straight chain alkanes.

Permanent Dipole

This type of intermolecular force acts between molecules with a **polar bond**. The $\delta+$ and $\delta-$ regions attract each other and hold the molecules together in a **lattice-like structure**.

Example:



Hydrogen

This is the

intermolecular force. Hydrogen bonds only act between hydrogen and the three most

Bonding

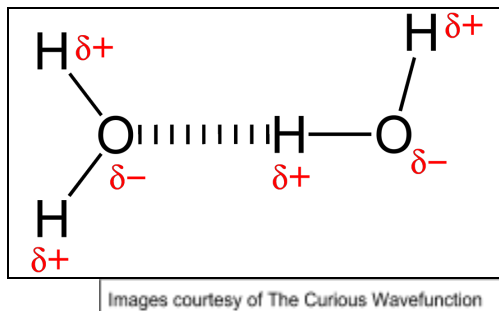
strongest type of





electronegative atoms: **Nitrogen, Oxygen and Fluorine**. The **lone pair** on these atoms forms a bond with a hydrogen atom from another molecule, shown with a **dotted line**.

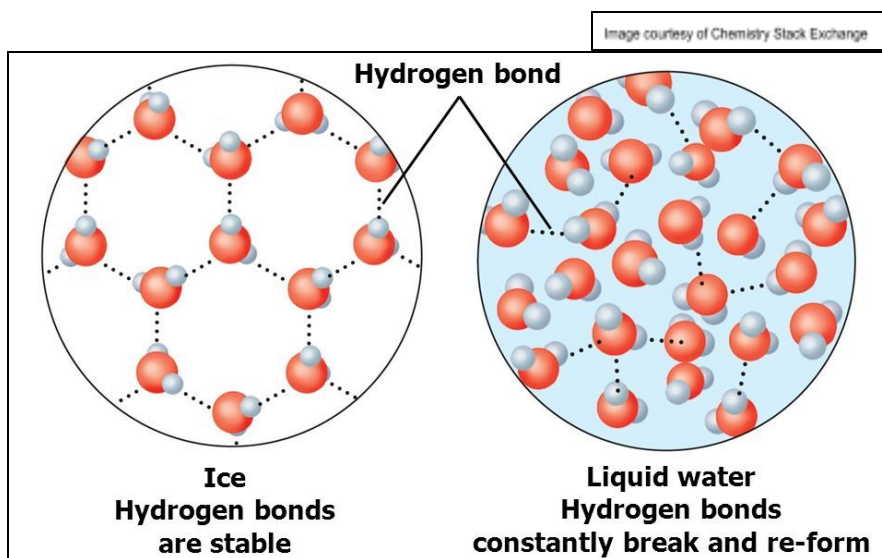
Example:



Molecules held together with hydrogen bonds have **much higher melting and boiling points** compared to similar sized molecules without hydrogen bonding. This shows how the type of intermolecular force heavily influences the **physical properties** of a substance.

Water has a simple molecular structure but has an **unusually high boiling point** for the size of molecule due to the presence of hydrogen bonds that require more energy to be overcome. These bonds also result in ice having a much **lower density** than liquid water as they hold the molecules in a **rigid structure** with lots of air gaps.

Example:

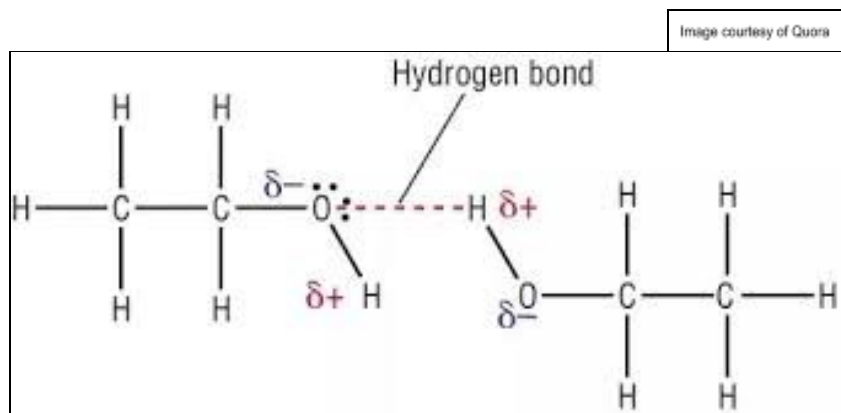


Hydrogen bonding is also responsible for the fact that **alcohols** have **much higher boiling points** than alkanes with a similar Mr value. This is because the lone electron pair on the oxygen atom is able to form **hydrogen bonds** with a hydrogen on another alcohol molecule.





Example:

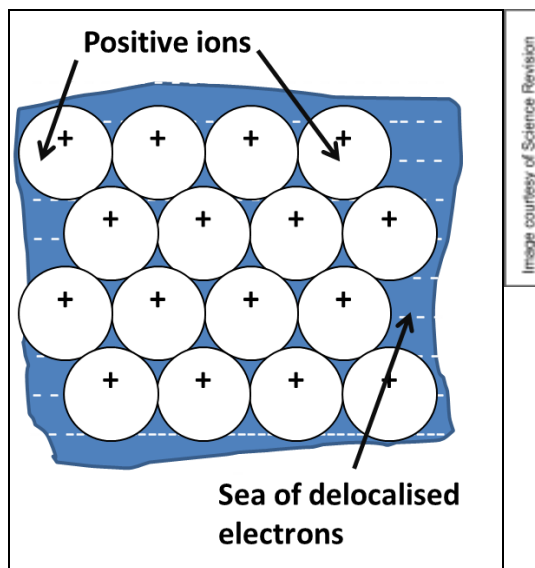


This same property makes alcohols and water **good solvents** for compounds that are able to form hydrogen bonds in solution. However it makes them poor for the dissolving of some **polar molecules** such as halogenoalkanes that cannot form this type of intermolecular force.

Metallic Bonding

Metallic bonding consists of a **lattice of positively charged ions** surrounded by a **'sea' of delocalised electrons**. This produces a very strong **electrostatic force of attraction** between these oppositely charged particles.

Example:



The **greater the charge** on the positive ion, the **stronger the attractive force** as more electrons are released into the 'sea'.

Ions that are **larger in size**, such as Barium, produce a **weaker attraction** due to their **greater atomic radius**.





Topic 2B: Structure

Bonding and Physical Properties

Physical properties of a substance include the boiling point, melting point, solubility and conductivity. They are different depending on the **type of bonding** and the **crystal structure** of the compound.

Crystal Structures

There are four main types of crystal structure; **ionic, metallic, simple molecular, macromolecular**, each with different physical properties.

Ionic (eg. Sodium Chloride)

Substances with an ionic crystal structure have a **high melting and boiling point**. This is because the electrostatic forces holding the ionic lattice together are strong and require a lot of energy to overcome.

When **molten or in solution**, ionic substances can **conduct electricity**. In this state, the ions separate and are no longer held in a lattice. Therefore they are free to move and **carry a flow of charge**, an electrical current.

Ionic substances are often **brittle** materials. When the layers of alternating charges are distorted, like charges repel, breaking apart the lattice into fragments.

Example:

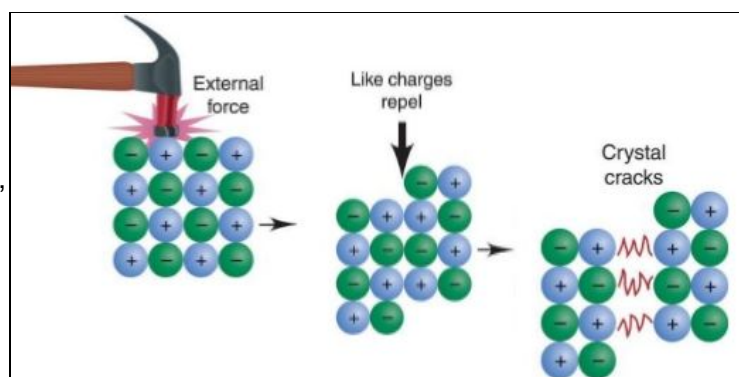


Image courtesy of Saints Chemistry

Metallic (eg. Aluminium)

Substances with metallic structure are often **good conductors**. The 'sea' of delocalised electrons is able to move and **carry a flow of charge**.

Metals are also **malleable** as the layers of positive ions are able to slide over one another. The delocalised electrons prevent fragmentation as they can move around the lattice.

Example:

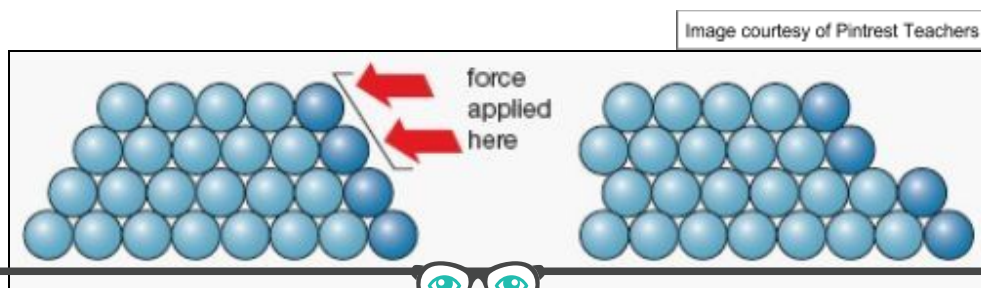


Image courtesy of Pintrest Teachers



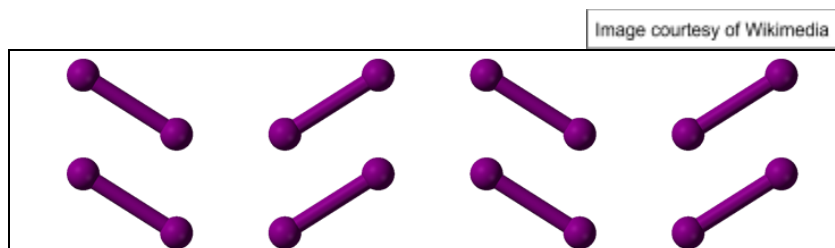


The electrostatic forces of attraction between the positive ions and delocalised electrons are very strong and therefore require a lot of energy to overcome. This means metallic substances have **high melting points** and are nearly always **solid at room temperature**. Mercury is the only liquid metal at room temperature.

Simple covalent (eg. Water)

Substances with a simple molecular structure consist of **covalently bonded molecules** held together with weak **van der waals** forces. These are a type of intermolecular force that act between the molecules holding them in a structure.

Example:



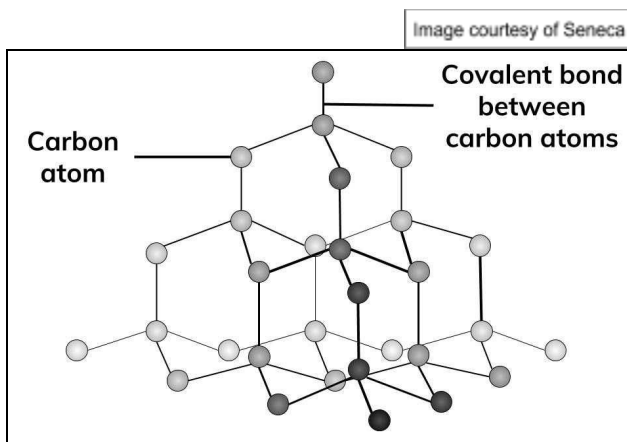
These van der waals forces are **very weak** and not much energy is required to overcome them meaning simple molecular substances have **low melting and boiling points**. Simple molecular substances are **very poor conductors** as their structure contains no charged particles.

Macromolecular (eg. Diamond)

Substances that have a macromolecular structure are **covalently bonded** into a **giant lattice** structure. Each atom has **multiple covalent bonds** which are very strong, giving the substance a **very high melting point**.

The strength of the covalent lattice makes macromolecular substances **rigid**. Diamond is a macromolecular structure made up of carbon atoms each bonded to four further carbon atoms. This makes diamond one of the hardest, strongest materials known.

Example:



Graphite is another structure made up of

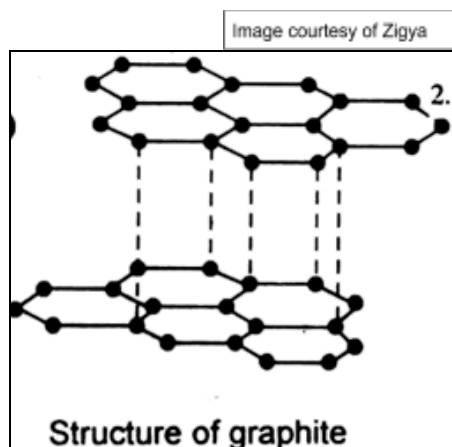
macromolecular carbon atoms.





However, in graphite, each carbon atom is bonded to three others in **flat sheets**. The electrons not used in bonding are released as **free electrons** which move between layers, meaning it can **conduct electricity**.

Example:



Graphene is a relatively recent material consisting of single, **2D sheets of graphite** that are just **one atom thick**. These sheets are formed of **hexagonal carbon rings** that create a very strong, rigid material that is extremely **lightweight**. Delocalised electrons move through each layer allowing it to **conduct** electricity.

Example:

