AQA Chemistry A-level

3.1.3: Bonding and Structures

Detailed Notes

3.1.3.1 - 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:

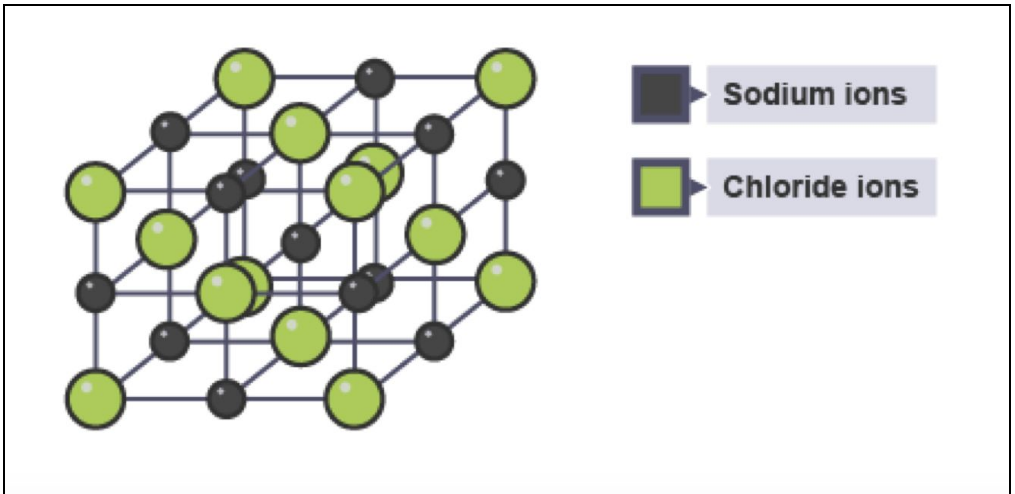


Image courtesy of Blogspot IGCSE Chemistry

Common compound ions include:

Sulfate - SO_4^{2-}
Hydroxide - OH^-
Nitrate - NO_3^-
Carbonate - CO_3^{2-}
Ammonium - NH_4^+

3.1.3.2 - Covalent and Dative Bonding

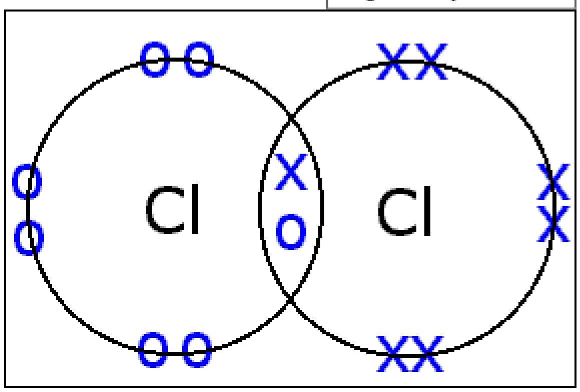
Covalent Bonding

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:

Image courtesy of IB Chem



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^+).

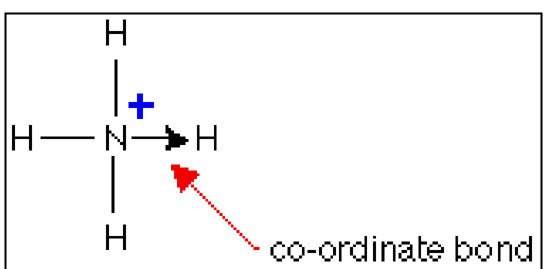


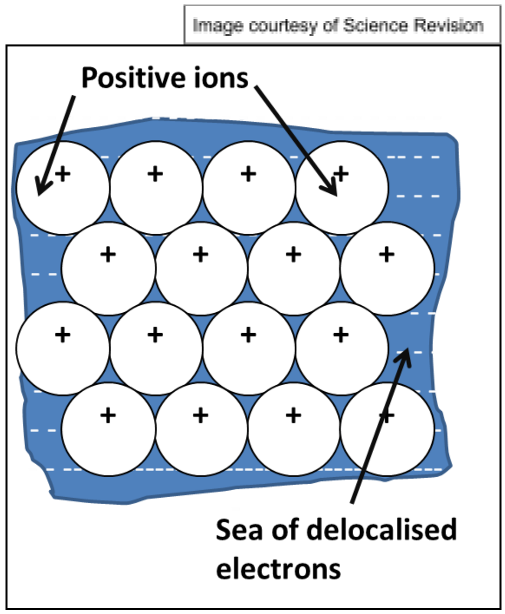
Image courtesy of Chemguide

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

3.1.3.3 - 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**.

3.1.3.4 - 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 and 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:

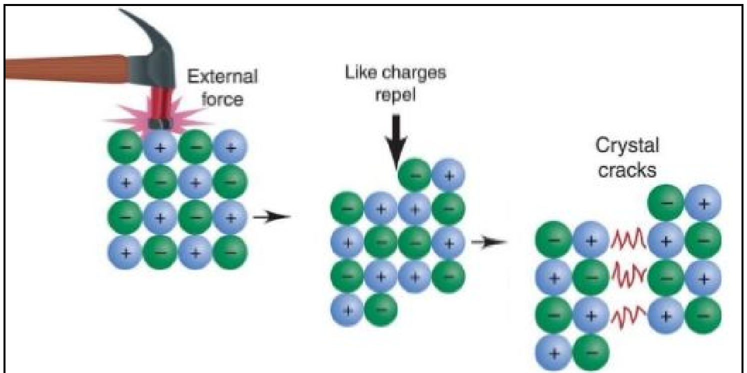


Image courtesy of Saints Chemistry

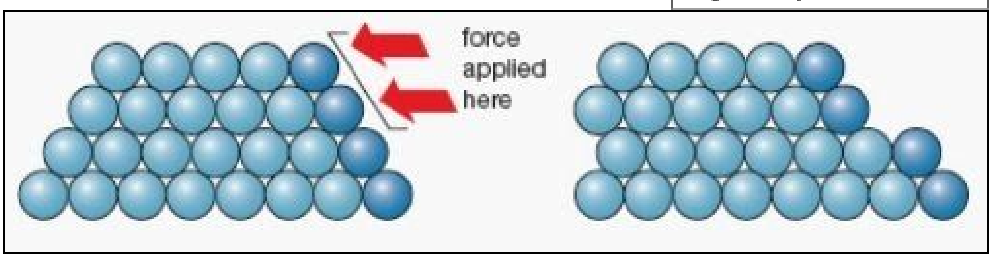
Metallic (eg. Aluminium)

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



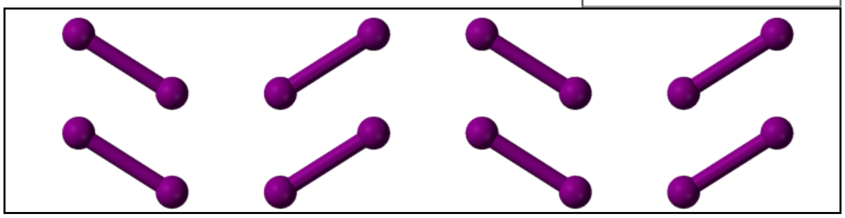
The electrostatic forces of attraction between the positive ions and delocalised electrons are very strong and therefore require a lot of energy to overcome. This means metallic substances have **high melting points** and are nearly always **solid at room temperature**. Mercury is the only liquid metal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

Simple Molecular (eg. Iodine)

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:

Image courtesy of Wikimedia



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**. Water has a simple molecular structure but has an unusually high boiling point for the size of molecule due to the presence of hydrogen bonding.

Simple molecular substances are **very poor conductors** as their structure contains no charged particles.

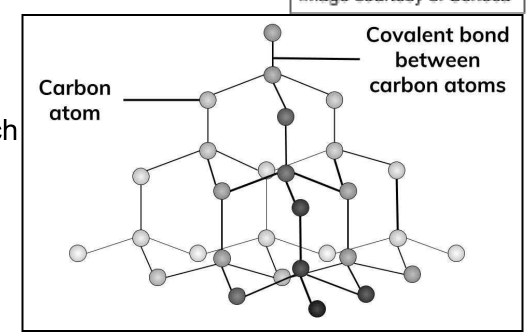
Macromolecular (eg. diamond)

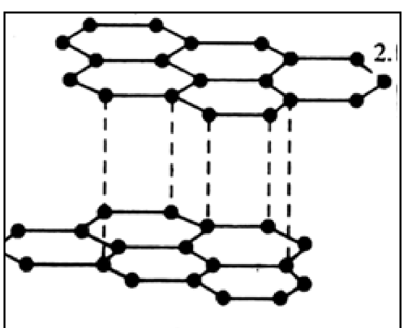
Substances that have a macromolecular structure are **covalently bonded** into a **giant lattice** structure. Each atom has **multiple covalent bonds** which are very strong, giving the substance a **very high melting point**.

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:

Image courtesy of Seneca





Structure of graphite

Image courtesy of Zigya

Graphite is another macromolecular structure made up of 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:

3.1.3.5 - Shapes of Molecules

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

Lone Pair Repulsion

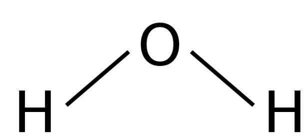
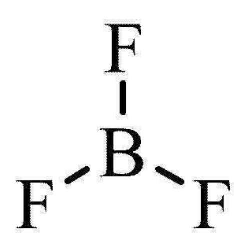
Any lone pairs present around the central atom provide **additional repulsive forces**, which changes the bond angle. For every lone pair present, the bond angle between covalent bonds is **reduced by 2.5°**.

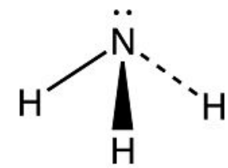
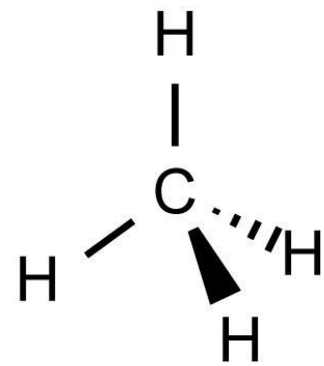
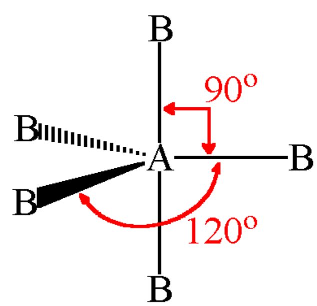
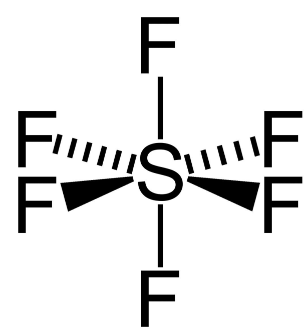
Molecule Shapes

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	90 and 120	
Octahedral	6	0	90	

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

3.1.3.6 - 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:

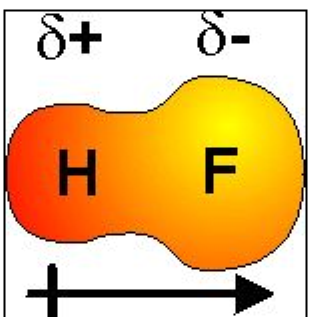
The power of an atom to attract negative charge towards itself within a covalent bond.

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.

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 other atom, producing a **∂-** region and a **∂+** region. This is a **permanent dipole**.

Example:



Images courtesy of Socratic

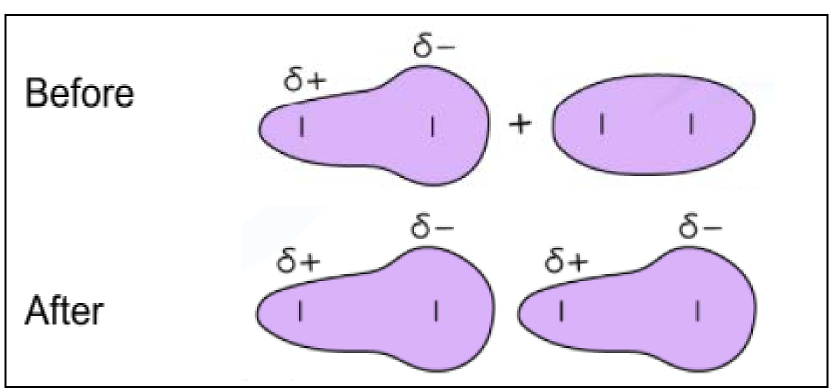
Hydrogen fluoride is a polar molecule as fluorine is a lot more electronegative than hydrogen so electrons are drawn to the right.

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

Induced Dipole

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

Example:



3.1.3.7 - 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 line up and **pack closer together**. This **reduces the distance** over which the force acts, therefore they are stronger.

Permanent Dipole

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

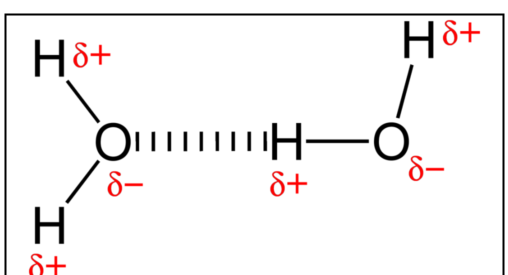
Example:

∂+		∂-	∂+		∂-	∂+		∂-
H	-	F	H	-	F	H	-	F

Hydrogen Bonding

This is the **strongest** type of intermolecular force. Hydrogen bonds only form between hydrogen and the three most 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:



Images courtesy of The Curious Wavefunction

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 forces acting between the molecules heavily influences the **physical properties** of a substance.