# 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 form full outer shells.

Transfer of electrons 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 form an ionic lattice with strong electrostatic attraction between oppositely charged ions:**

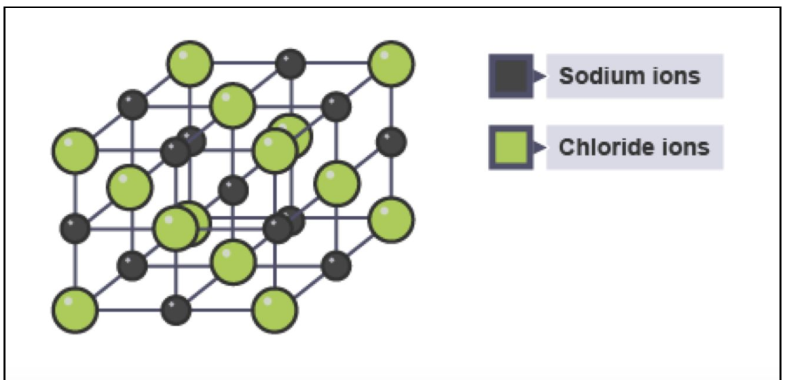


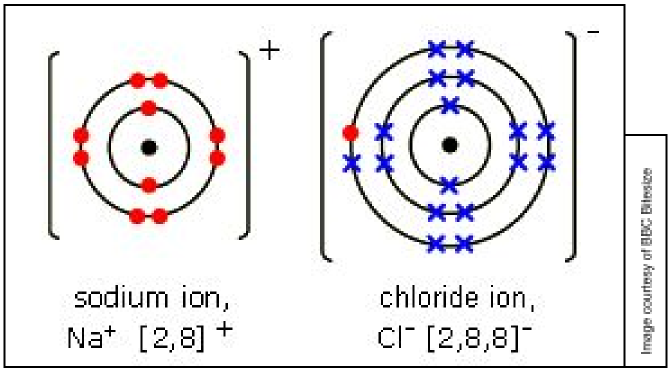

Image courtesy of Blogspot IGCSE Chemistry

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.

*Example: Sodium chloride dot and cross diagram*




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

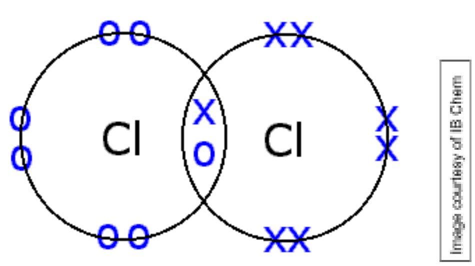
## Covalent Bonding

Covalent bonds form between **two non-metals**. There is a strong electrostatic attraction between the two nuclei and the shared electrons between them. Electrons are **shared** between the two outer shells in order to form 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**. The overlap includes a covalent bond. The **number of electrons** within the overlap tells you the nature of the covalent bond:

- 2 electrons (1 from each atom): **single bond**, displayed formula represented as —
- 4 electrons (2 from each atom): **double bond**, displayed formula represented as =
- 6 electrons (3 from each atom): **triple bond**, displayed formula represented as ≡

*Example: Dot and cross diagram of chlorine, $Cl_2$*



**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: Dot and cross diagram of oxygen,* $O_2$

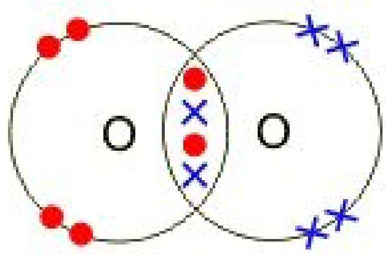



Image courtesy of BBC Bitesize

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 ($NH_4^+$)**

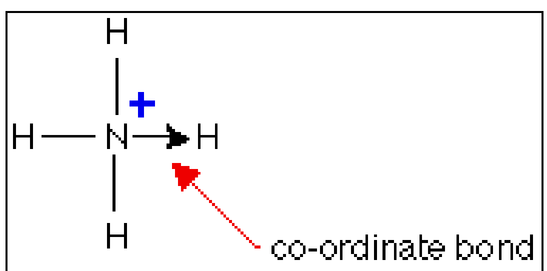
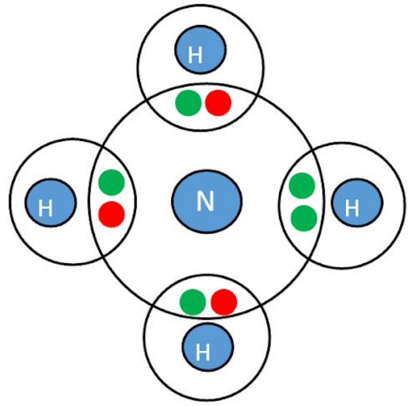


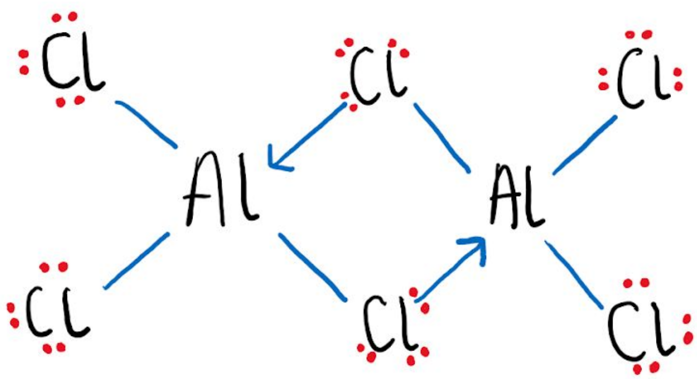

Image courtesy of Chemguide

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

Since both electrons come from the **same atom** in a dative covalent bond, in dot and cross diagrams both electrons in that bond will have the **same shape**. In other words, they will both be dots or both be crosses.

*Example dative bonded structures:* $NH_4^+$ and $Al_2Cl_6$



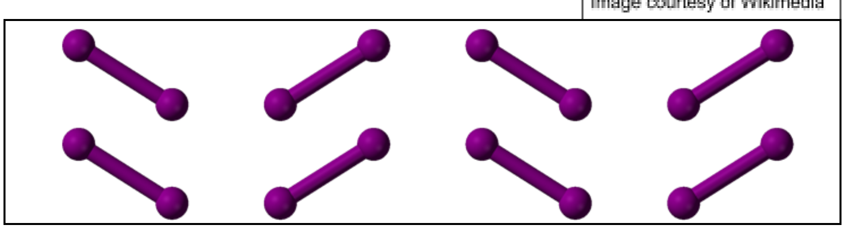




## Simple Covalent

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:*

Image courtesy of Wikimedia



## 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

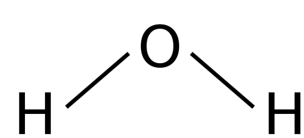
The shape of a molecule can be determined by considering the **type and quantity of electron pairs**:

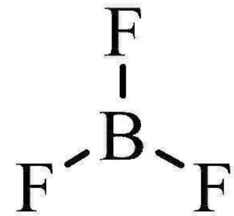
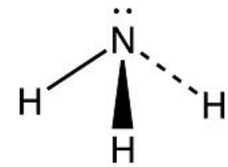
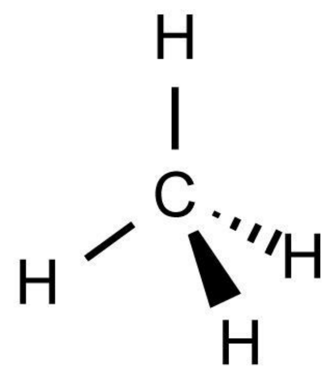
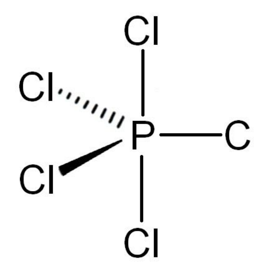
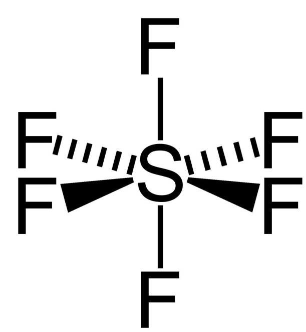
1. Find the number of electron pairs.
2. Determine how many of the pairs are bonding pairs and how many are lone pairs.
3. Bonding pairs indicate the basic shape and lone pairs indicate any additional repulsion.

Molecules may also be described in terms of their bond lengths and bond angles.

- **Bond length** - the average distance between two nuclei in a covalent bond.
- **Bond angle** - the angle between two covalent bonds from the same atom.

This table shows some common molecule shapes:

| Name | Bonding $e^-$ Pairs | Lone $e^-$ Pairs | Bond Angle (°) | Example |
|---|---|---|---|---|
| Linear | 2 | 0 | 180 | Cl - Be - Cl<br>O=C=O |
| V - Shaped | 2 | 2 | 104.5 |  |

| | | | | |
|---|---|---|---|---|
| Trigonal Planar | 3 | 0 | 120 |  |
| Triangular Pyramid | 3 | 1 | 107 |  |
| Tetrahedral | 4 | 0 | 109.5 |  |
| Trigonal Bipyramidal | 5 | 0 | 90 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

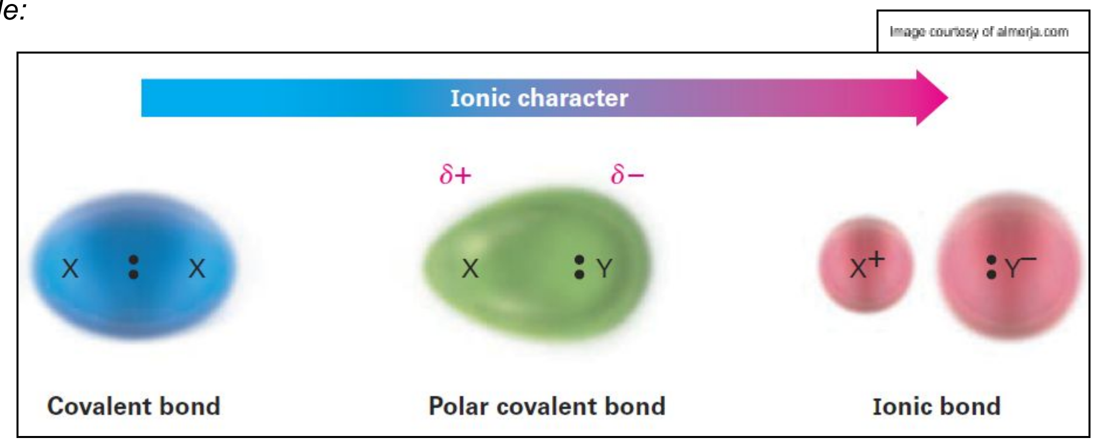
Every atom has electronegativity, which is defined as:

**The power of an atom to attract the electron pair in a covalent bond towards itself.**

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

Ionic and covalent bonding are the **extremes** in a **continuous scale of bonding** as shown below. If the electronegativity between two atoms is great enough, an ionic bond will form between them.

*Example:*




Bond polarity can be permanent or induced depending on the molecule and its interactions.

### Polar Bonds and Polar Molecules

A polar bond results from a **large difference in electronegativity** between two atoms forming a covalent bond. Tables of electronegativity can be used to workout if a bond will be polar or not. If the difference in electronegativity between two bonding atoms is between 0.4 and 1.7, the bond will be a polar **covalent bond**. If the difference in electronegativity is greater than that, the bond will be **ionic**.

Since electronegativity is a **periodic trend**, elements that are close together on the periodic table will not form polar bonds.

**Polar molecules**
**Polar molecules** arise when there is an **overall** difference in polarity **across the molecule**, due to the arrangement of polar bonds and the geometry of the molecule. Polar molecules must have polar bonds, however a molecule with polar bonds may not necessarily be a polar molecule.

*Example:*


**$CO_2$** - the C=O bonds in $CO_2$ are polar, however the molecule is linear so the dipoles created by each polar bond cancel out.

**$H_2O$** - the O-H bonds in water are polar and the geometry of the molecule is bent, so overall there is a polarity and the molecule is polar.

## Intermolecular Forces

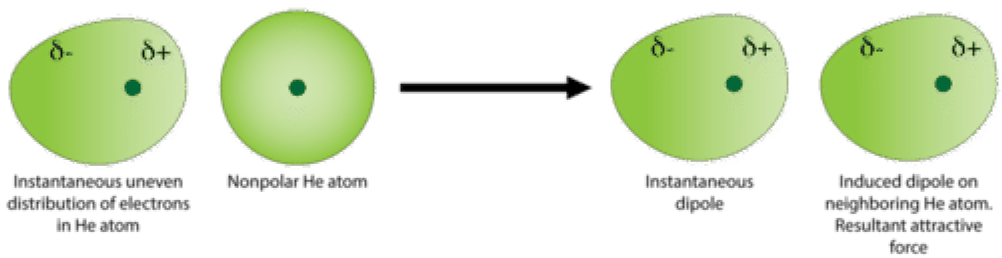
There are **three main types of intermolecular force**. Each one differs in strength and in what they act between.

## Van der Waals Forces

Van der Waals forces are the **weakest** type of intermolecular force. They act as an **induced dipole** between molecules. They are also called **London forces** or instantaneous dipole-induced forces.

The strength of van der waals forces varies depending on the Mr of the molecule and its shape. The **greater 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 much closer together**. This **reduces the distance** over which the force acts, making the intermolecular force stronger.

*Example:*




[CC BY-NC-SA 3.0 US](), [Allison Soult, Chemistry LibreTexts]()

### Boiling Point Trends of Alkanes

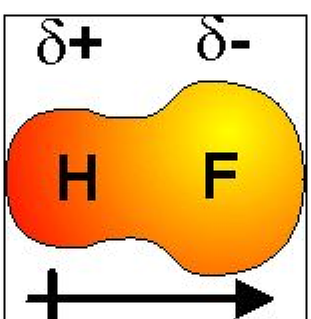
Van der waals forces act between organic **alkane chains** and are affected by the **chain length** and any **branching**. As the chain length of the alkane increases, so does the **Mr** of the molecule. This results in **stronger** intermolecular forces between the chains, and so the compound has a **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. This means the distance over which the intermolecular forces act is increased, weakening the **attractive forces.** Therefore, branched-chain alkanes have **lower boiling points** than straight-chain alkanes.

### Permanent Dipole

If the two atoms that are bonded have sufficiently 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 **∂-** region and a **∂+** region. This produces 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 towards the fluorine atom.**

The **∂+** and **∂-** regions of neighbouring polar molecules attract each other and hold the molecules together in a **lattice-like structure**.

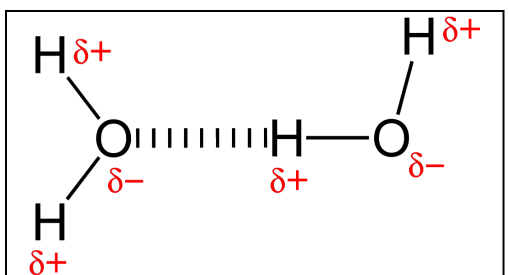
*Example:*

| ∂+   ∂- | ∂+   ∂- | ∂+   ∂- |
|---|---|---|
| H - Cl | H - Cl | H - Cl |

## Hydrogen Bonding

Hydrogen bonding is the **strongest** type of intermolecular force. Hydrogen bonds only act between hydrogen and the three most electronegative atoms: **nitrogen, oxygen and fluorine**. The **lone pair** on these atoms form a bond with a δ+ hydrogen atom from another molecule, shown with a **dotted line**. **$H_2O$, $NH_3$** and **HF** all have hydrogen bonds between molecules.

*Example:*




Images courtesy of The Curious Wavefunction

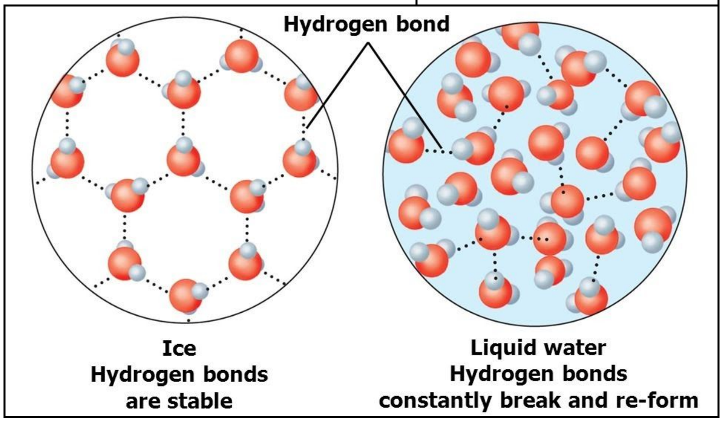
### Properties

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 the molecule. This is due to the presence of hydrogen bonds that require a lot of 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:*

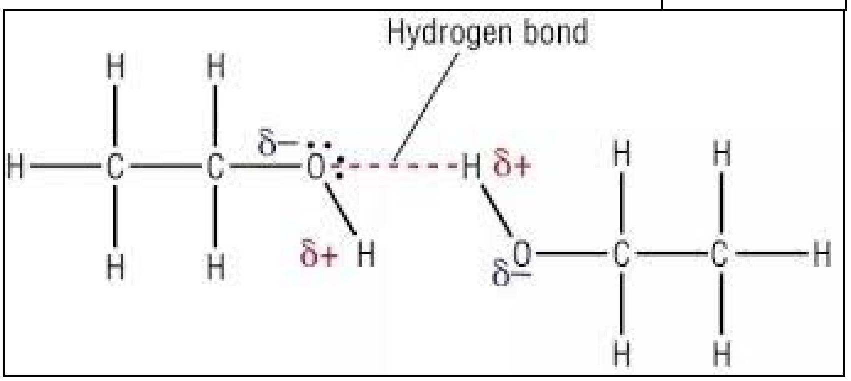
Image courtesy of Chemistry Stack Exchange




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:*

Image courtesy of Quora



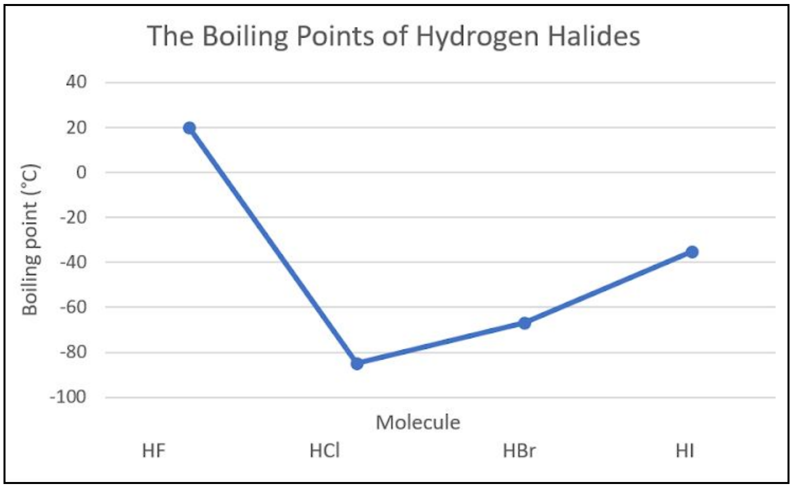

## Solvents

Water is a popular choice of **solvent**. Its hydrogen bonding capabilities allow it to dissolve some **ionic compounds** by solvating the individual ions, and to dissolve some alcohols by forming **hydrogen bonds** with their **hydroxyl** group.

However, both water and alcohols are poor solvents for the dissolving of some **polar molecules** such as halogenoalkanes that cannot form hydrogen bonds.

**Non-aqueous solvents** are often used for compounds which have the same type of intermolecular force.

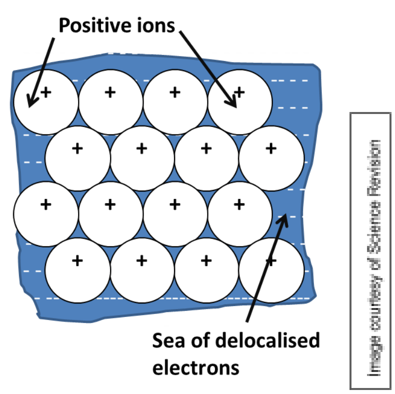
### Boiling Point Trends of Hydrogen Halides




**Hydrogen fluoride** is the only hydrogen halide that forms **hydrogen bonds** between molecules. This gives it the highest boiling point because hydrogen bonds are much **stronger** than van der waals and permanent dipole forces.

The boiling point **increases** as you move down the group past hydrogen fluoride because as the halide increases in size, their **number of electrons** also increases. This means more **van der waals** forces form between molecules, so more energy is required to separate them.

## Metallic Bonding




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

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

The **physical properties** of a substance include its boiling point, melting point, solubility and conductivity. They are different depending on the **type of bonding** present, the types of particle present 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 (e.g. Sodium Chloride)

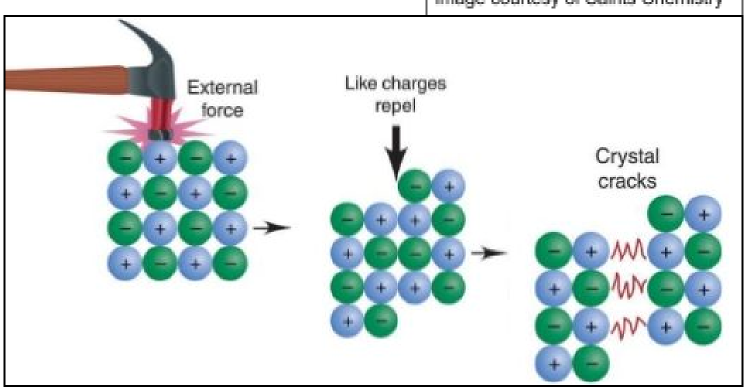
Substances with an ionic crystal structure have a **high melting and boiling point**. This is because the electrostatic forces holding the ionic lattice together are strong and require a lot of energy to overcome.

When **molten or dissolved 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** so can conduct 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



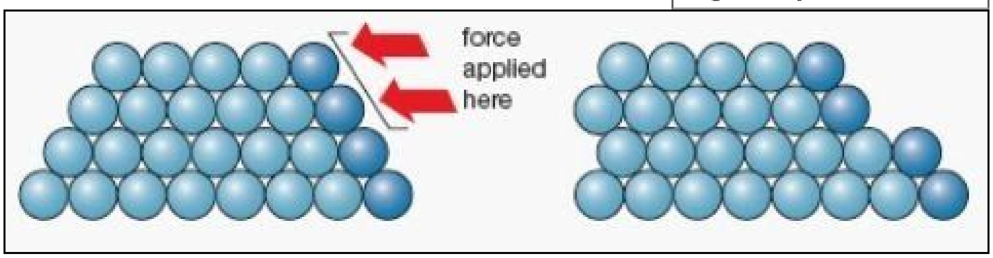
### Metallic (e.g. Aluminium)

Substances with metallic structures are often **good conductors**. The delocalised electrons are 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.

*Example:*

Image courtesy of Pintrest Teachers



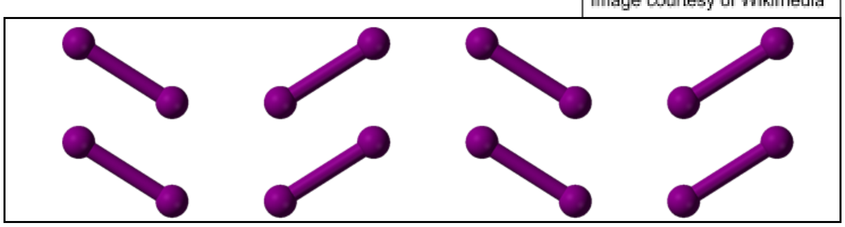

The electrostatic forces of attraction between the positive ions and delocalised electrons are very strong and therefore require a lot of energy to overcome. This means metallic substances have **high melting points** and are nearly always **solid at room temperature**. Mercury is the only metal that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

### Simple molecular (e.g. Water)

Substances with a simple covalent molecular structure consist of **covalently bonded molecules** held together with weak **van der waals** forces. This is the structure formed by water and iodine, $I_2$.

*Example:*

Image courtesy of Wikimedia



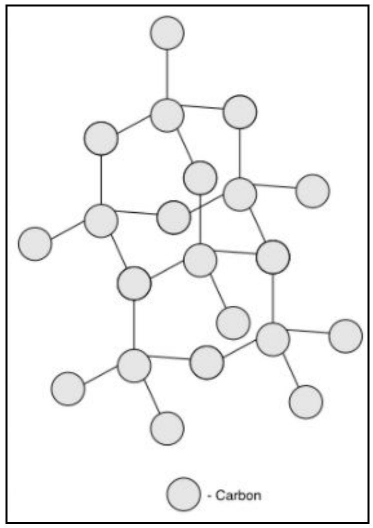
These van der waals forces are **very weak** and not much energy is required to overcome them. This means simple molecular substances have **low melting and boiling points**. Simple molecular substances are **very poor conductors** as their structure contains no charged particles.

**Giant Covalent Structures**

Macromolecular covalent substances 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**.

**Diamond** is a **macromolecular structure** made up of carbon atoms each bonded to four other carbon atoms. This forms a **rigid tetrahedral structure**, making diamond one of the **hardest, strongest** materials known - which is why it is often used on the tips of drills.

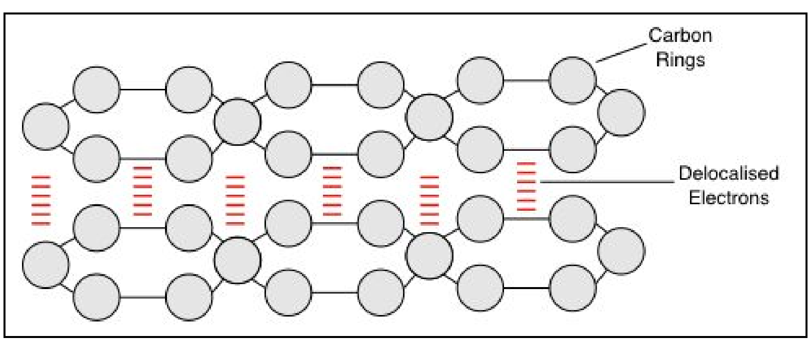
*Example: Structure of diamond*




**Graphite** is another macromolecular structure made up of carbon atoms. However, in graphite, each carbon atom is bonded to three others in **flat hexagonal sheets**. This means there is one **delocalised electron** per carbon atom. These electrons can move freely, allowing graphite to **conduct electricity**. Graphite can therefore be used in an electrode.

The **intermolecular forces** between layers of graphite are **weak**, allowing the layers to easily slide over each other, meaning graphite can be used as a **lubricant**.

*Example: Structure of graphite*



**Graphene** consists 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, making graphene a useful material in electronics.

*Example: Structure of graphene*