Edexcel IAL Chemistry A-Level

Topic 3: Bonding and Structure

Detailed Notes

Topic 3A: Ionic Bonding

Evidence for the existence of ions

Physical evidence

Ionic substances are **brittle** and they are unable to conduct electricity in solid form. However, when ionic substances are **molten** or **aqueous** the ions are free to move and so can **conduct electricity** in these states. These properties, in comparison to metals which are malleable and conduct electricity when solid, can be explained by the model of **oppositely charged ions** in a **giant ionic lattice.**

Electron density maps

Electron density maps show the region around a nucleus in which electrons are distributed. A **high density** corresponds to a **high probability** of an electron being there. Different types of chemical bonds have different electron density maps. In covalent bonds, for instance, there is a high electron density between the bonding nuclei, whereas in ionic bonds there is a low electron density between ions. This is explained by the fact that ionic bonds are formed by the **physical transfer of electrons** whereas covalent bonds are formed by the **sharing of electrons.**

Electron density maps therefore allow chemists to work out the **type of bonding** present, and also the **distance** between ions in an ionic lattice.

Ion migration

In electrolysis, ions are free to move and ions move to the **oppositely charged electrode -** where they gain or lose electrons to form atoms. This ion migration can be explained by the **electrostatic attraction** that results between oppositely charged species.

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, according to the octet rule. The **octet** rule states that atoms will often react to form a full outer shell containing 8 electrons.
When the electrons are transferred, it creates **charged particles** called **ions**. Oppositely charged ions **attract through strong electrostatic forces** to form a **giant ionic lattice**.

Positive ions are formed when an atom **loses** at least one electron.
Negative ions are formed when an atom **gains** at least one electron.

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. The ions are then attracted to one another, forming an ionic lattice:

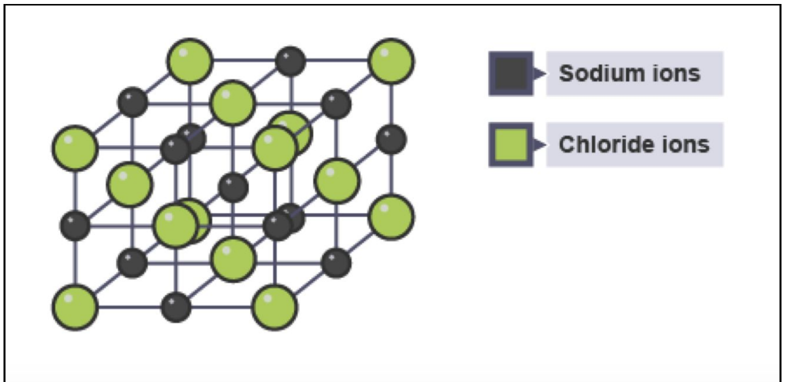


Image courtesy of Blogspot IGCSE Chemistry

Ionic bond strength

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**.

Dots and Cross Diagrams

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 displayed on the outer shell of the anion.

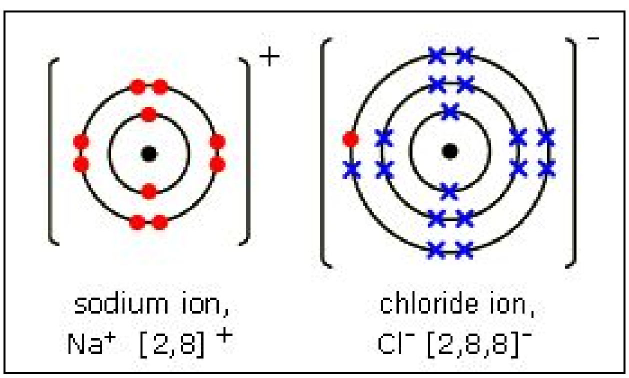


Image courtesy of BBC Bitesize

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

Ionic radii

Down a group
As you move **down** a group, the number of **electron shells increases**, and therefore the **ionic radius increases.**

Across a period
Groups 1 and 2: The atoms in these groups **lose electrons** to form **positive ions**. As you go across the period, the magnitude of the positive charge increases for the same amount of electron shielding. This means there is a greater electrostatic attraction which pulls the outer electrons more tightly in towards the nucleus, **decreasing the ionic radius**.

The ions formed are said to be **isoelectronic** as they have the **same electronic configuration** (but lose different numbers of electrons to form this configuration).

Example: Ne, Na^+ and Mg^{2+} all have the following electron configuration: $1s^22s^22p^6$.

Groups 6, 7 and 8: The atoms in these groups **gain electrons** to form **negative ions.** As you go across the period the additional proton makes less of a difference, so the only change is a **slight decrease** in ionic radius.

Example: N^{3-}, O^{2-} and F^- all have the following electron configuration: $1s^22s^22p^4$.

Polarisation

In ionic bonds, polarisation refers to the **distortion** of the **electron cloud**. In the **perfect ionic model**, ions are spherical with no distortion. In reality, the positive ion will attract some of the negative ion's electron cloud towards itself.

The **polarising power** of a cation and the **polarisability** of an anion depends on the radius and charge of the ions.

- A highly charged **cation** will have a strong electrostatic attraction with an anion which will distort the anion's electron cloud. If a cation has a smaller radius, there will generally be a greater concentration of charge in a small area, causing a distortion of the electron cloud. Therefore, the cation's **polarising power increases** with an **increase in charge** and **decrease in radius**.
- The more negatively charged an **anion** is, the greater the electron cloud. The larger the electron cloud, the more easily it is distorted. Also, a larger radius means the outer electrons are held more loosely, allowing them to be more easily distorted.

Therefore, the anion's **polarisability increases** as the **radius increases** and charge becomes **more negative**.

Topic 3B: Covalent Bonding

Evidence for the existence of covalent substances

Physical evidence

Giant covalent structures have **high melting** and **boiling points** due to **strong covalent bonds.** They have no free ions or electrons (except graphite and graphene) so generally cannot conduct electricity.
Simple covalent molecules have **low melting** and **boiling points** because, although the covalent bonds are strong, the **intermolecular forces** between molecules are **weak.**

The properties of these covalent structures can all be explained by the model of covalent bonding.

Electron density maps

As described in section 3A, electron density maps show the regions around nuclei in which electrons are likely to be found. They can be used to identify the type of bonding present. Covalent bonds involve the sharing of electrons between nuclei, so the electron density map shows a **high density of electrons between bonding atoms.**

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

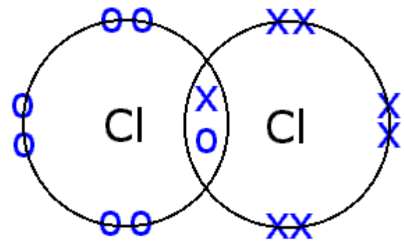


Image courtesy of IB Chem

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

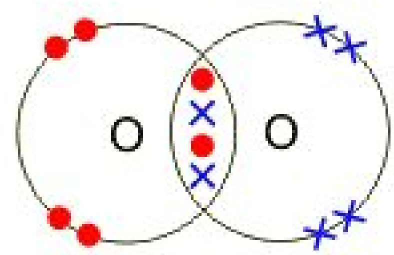


Image courtesy of BBC Bitesize

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.

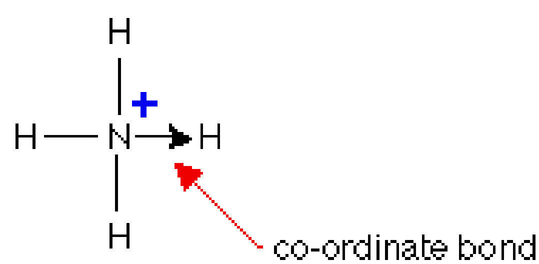
Dative Bonding

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^+).

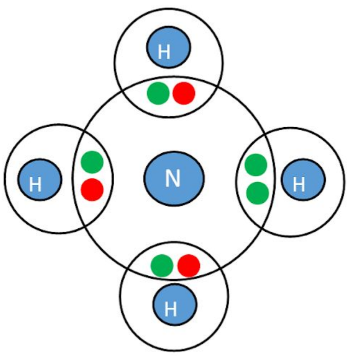
Once a dative bond has formed, it is treated as a **standard covalent bond** because it **reacts in exactly the same way** and has the same properties regarding length and strength.

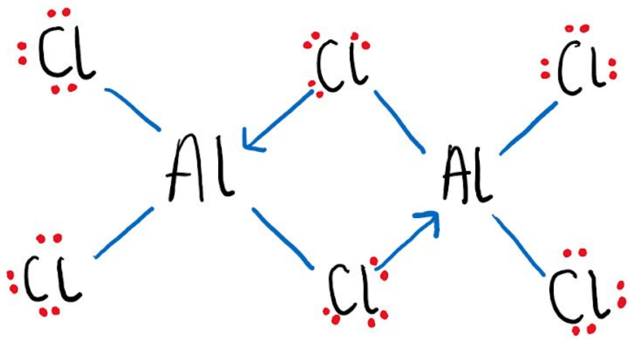


Since both electrons come from the **same atom** in a dative covalent bond, in dot and cross diagrams both electrons in that bond will have the **same shape**.
In other words, they will both be dots or both be crosses.

Image courtesy of Chemguide

Example structures: NH_4^+ and Al_2Cl_6



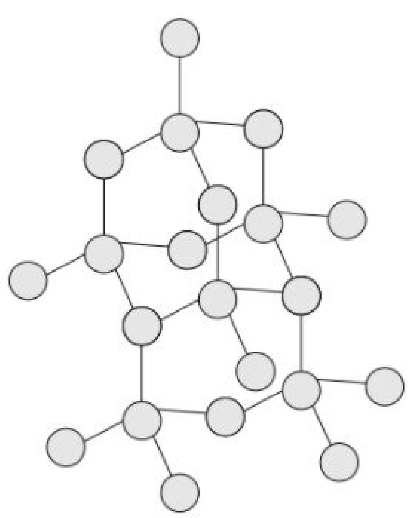


Giant Covalent Structures

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**.

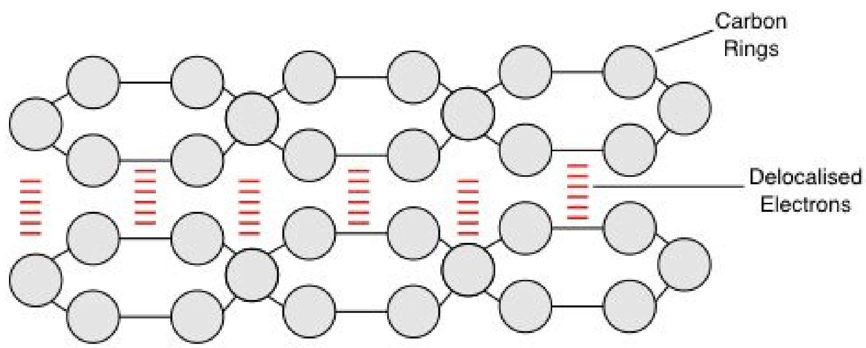
Diamond is a **macromolecular structure** made up of carbon atoms each bonded to four further carbon atoms. This forms a **rigid tetrahedral structure**, making diamond one of the **hardest, strongest** materials known - which is why it is often used on the tips of drills.

Example: Structure of diamond



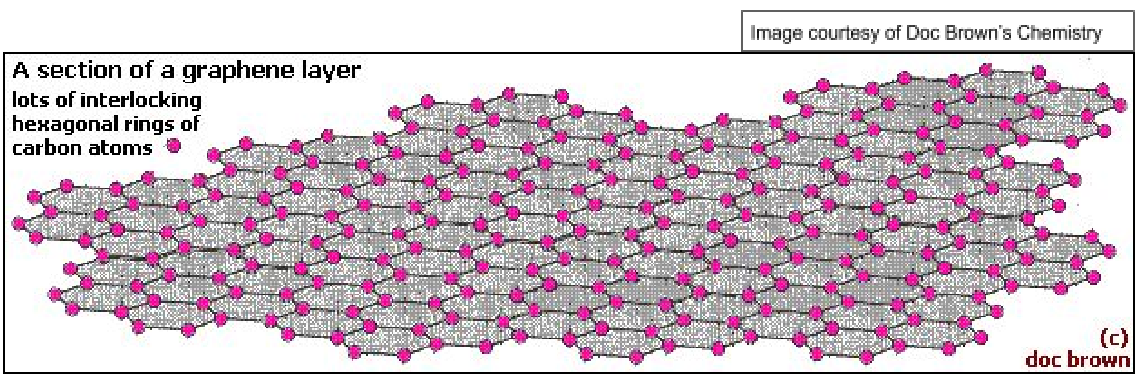
Graphite is another macromolecular structure made up of carbon atoms. However, in graphite, each carbon atom is bonded to three others in **flat hexagonal sheets**. This means there is one **delocalised electron** per carbon atom. These electrons can move freely, allowing graphite to **conduct electricity**.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



Bond Polarity

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

Electronegativity

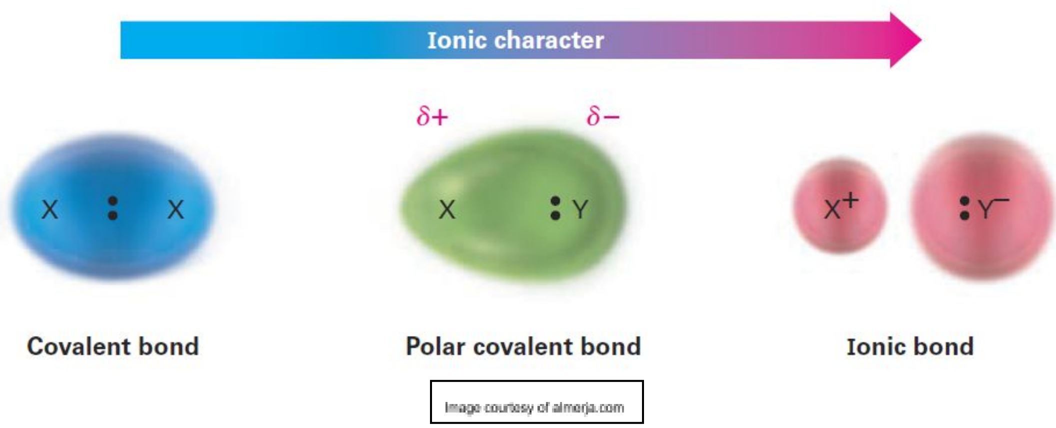
Every atom has **electronegativity**, which is defined as:

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 you move across a period, atoms have a greater nuclear charge and a smaller covalent radius. This allows the nucleus to attract the bonding electrons more strongly. Electronegativity **decreases down a group**. Going down a group, atoms increase in size due to the extra electron shells, increasing shielding towards the bonding electrons.

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:

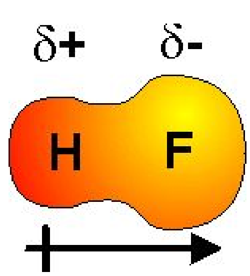


This bond polarity can be **permanent** or **induced** depending on the molecule and how it interacts with things around it.

Permanent Dipole

If the two atoms that are bonded have 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:



Hydrogen fluoride is a polar molecule as fluorine is a lot more electronegative than hydrogen. This causes electrons to be drawn towards the fluorine atom.

Images courtesy of Socratic

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

Polarisability of ions

Where electronegativity difference is large enough the electrons will not be shared between atoms but will instead spend most of their time localised around one of the atoms only - forming **ions**. A cation is able to **distort** the shape of the **electron cloud** on a nearby anion. The extent by which the electron cloud is distorted is known as **polarisability**.

Smaller cations with a **greater charge** are the most polarising. **Anions** with a **larger ionic radius** can have their electron clouds distorted more easily.

Polar Bonds and Polar Molecules

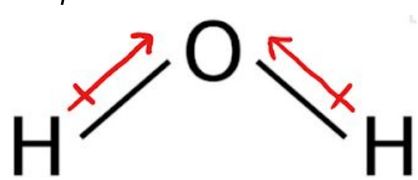
A polar bond results from a **large difference in electronegativity** between two atoms forming a covalent bond. Tables of electronegativity can be used to workout if a bond will be polar or not. If the difference in electronegativity between two bonding atoms is between 0.4 and 1.7, the bond will be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

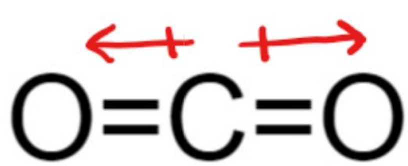
Since electronegativity is a **periodic trend**, elements that are close together on the periodic table will not form polar bonds.

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.

Topic 3C: Shapes of 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**:

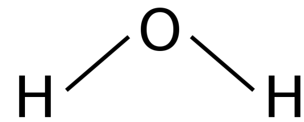
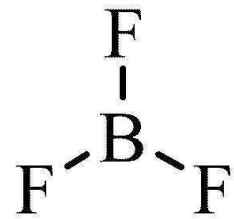
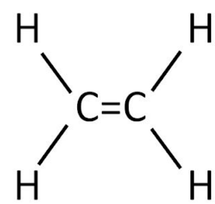
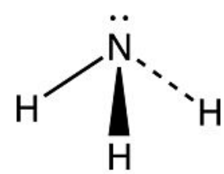
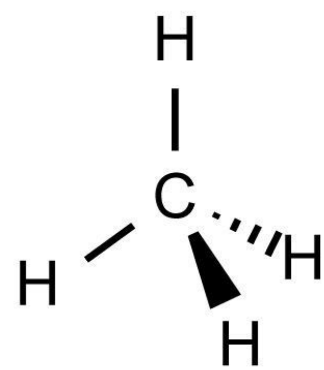
1. Find the number of electron pairs.
2. Determine how many of the pairs are bonding pairs and how many are lone pairs.
3. Bonding pairs indicate the basic shape and lone pairs indicate any additional repulsion.

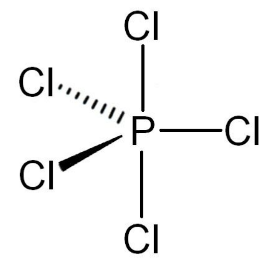
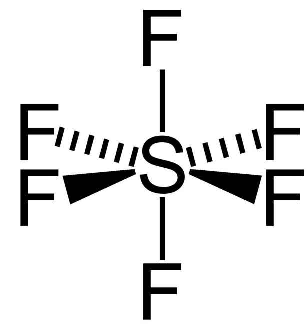
Molecules may also be described in terms of their bond lengths and bond angles.

- **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 O=C=O

V - Shaped	2	2	104.5	
Trigonal Planar	3	0	120	 
Triangular Pyramid	3	1	107	
Tetrahedral	4	0	109.5	CH_4, NH_4^+ 

Trigonal Bipyramidal	5	0	180 and 120	
Octahedral	6	0	90	

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

Topic 3D: Metallic Bonding

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

Example:

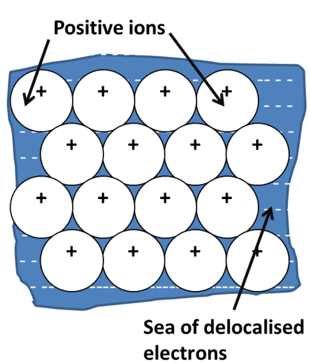


Image courtesy of Science Revision

The **greater the charge** on the positive ion, the **stronger the attractive force** since 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**.

Metallic Structure (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** because the uniform layers of positive ions are able to slide over one another. The delocalised electrons prevent fragmentation as they can move around the lattice.

Example:

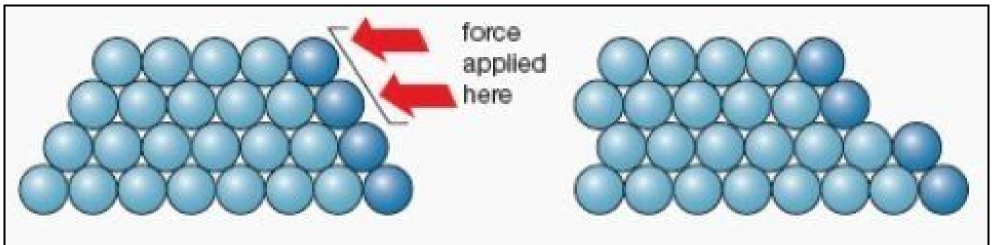


Image courtesy of Pintrest Teachers

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 metal which is liquid at room temperature.