

## Bonding

**Definition: Ionic bonding** is the **electrostatic force** of attraction between **oppositely charged ions** formed by electron transfer.

Metal atoms lose electrons to form +ve ions.  
Non-metal atoms gain electrons to form -ve ions.

Mg goes from  $1s^2 2s^2 2p^6 3s^2$   
to  $Mg^{2+} 1s^2 2s^2 2p^6$

O goes from  $1s^2 2s^2 2p^4$  to  
 $O^{2-} 1s^2 2s^2 2p^6$

Ionic bonding is stronger and the melting points higher when **the ions are smaller** and/ or have **higher charges**. E.g. MgO has a higher melting point than NaCl as the ions involved ( $Mg^{2+}$  &  $O^{2-}$  are smaller and have higher charges than those in NaCl,  $Na^+$  &  $Cl^-$ )

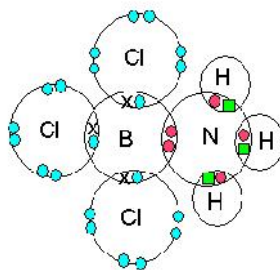
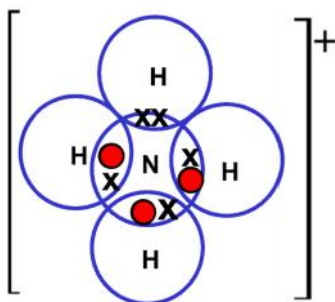
Definition: covalent bond

A **covalent bond** is a **shared pair** of electrons

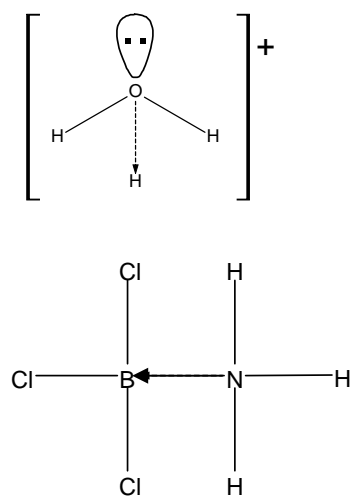
### Dative Covalent bonding

A **Dative covalent bond** forms when **the shared pair of electrons** in the covalent bond come from **only one of the bonding atoms**. A **dative covalent bond** is also called co-ordinate bonding.

Common examples you should be able to draw that contain dative covalent bond (e.g.  $NH_4^+$ ,  $H_3O^+$ ,  $NH_3BF_3$ )



The dative covalent bond acts like an ordinary covalent bond when thinking about shape so in  $NH_4^+$  the shape is tetrahedral



The direction of the arrow goes from the atom that is providing the lone pair to the atom that is deficient

### Metallic bonding

**Definition: Metallic bonding** is the **electrostatic force** of attraction between the **positive metal ions** and the **delocalised electrons**

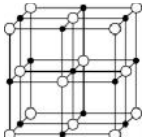
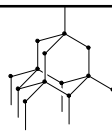
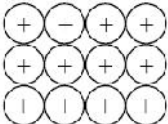
The three main factors that affect the strength of **metallic bonding** are:

1. Number of protons/ Strength of nuclear attraction.  
*The more protons the stronger the bond*
2. Number of delocalised electrons per atom (the outer shell electrons are delocalised)  
*The more delocalised electrons the stronger the bond*
3. Size of ion.  
*The smaller the ion, the stronger the bond.*

#### Example

**Mg has stronger metallic bonding than Na and hence a higher melting point. The Metallic bonding gets stronger because in Mg there are more electrons in the outer shell that are released to the sea of electrons. The Mg ion is also smaller and has one more proton. There is therefore a stronger electrostatic attraction between the **positive metal ions** and the **delocalised electrons** and **higher** energy is needed to break bonds.**

## Bonding and Structure

Bonding	Structure	Examples
Ionic : <b>electrostatic force</b> of attraction between <b>oppositely charged ions</b>	 <p style="text-align: center;"><b>Giant Ionic Lattice</b></p>	<b>Sodium chloride</b> Magnesium oxide
<b>Covalent</b> : <b>shared pair</b> of electrons	<b>Simple molecular:</b> With intermolecular forces (van der Waals, permanent dipoles, hydrogen bonds) between molecules	<b>Iodine</b> <b>Ice</b> Carbon dioxide Water Methane
<b>Covalent</b> : <b>shared pair</b> of electrons	<b>Macromolecular:</b> giant molecular structures. 	<b>Diamond</b> <b>Graphite</b> Silicon dioxide Silicon
<b>Metallic:</b> <b>electrostatic force</b> of attraction between the <b>metal positive ions</b> and the <b>delocalised electrons</b>	 <p style="text-align: center;"><b>Giant metallic lattice</b></p>	<b>Magnesium, Sodium</b> (all metals)

Only use the words **molecules** and **intermolecular forces** when talking about simple molecular substances

Property	Ionic	Molecular (simple)	Macromolecular	Metallic
<b>boiling and melting points</b>	high- because of giant lattice of ions with strong electrostatic forces between oppositely charged ions.	low- because of weak intermolecular forces between molecules (specify type e.g van der waals/hydrogen bond)	high- because of many strong covalent bonds in macromolecular structure. Take a lot of energy to break the many strong bonds	high- strong electrostatic forces between positive ions and sea of delocalised electrons
<b>Solubility in water</b>	Generally good	generally poor	insoluble	insoluble
<b>conductivity when solid</b>	poor: ions can't move/ fixed in lattice	poor: no ions to conduct and electrons are localised (fixed in place)	diamond and sand: poor, because electrons can't move (localised) graphite: good as free delocalised electrons between layers	good: delocalised electrons can move through structure
<b>conductivity when molten</b>	good: ions can move	poor: no ions	poor	(good)
<b>general description</b>	crystalline solids	mostly gases and liquids	solids	shiny metal Malleable as the positive ions in the lattice are all identical. So the planes of ions can slide easily over one another -attractive forces in the lattice are the same whichever ions are adjacent

## Shape of molecules

Name	No bonding pairs	No lone pairs	Diagram	Bond angle	Examples
linear	2	0		180	CO <sub>2</sub> , CS <sub>2</sub> , HCN, BeF <sub>2</sub>
Trigonal planar	3	0		120	BF <sub>3</sub> , AlCl <sub>3</sub> , SO <sub>3</sub> , NO <sub>3</sub> <sup>-</sup> , CO <sub>3</sub> <sup>2-</sup>
Tetrahedral	4	0		109.5	SiCl <sub>4</sub> , SO <sub>4</sub> <sup>2-</sup> , ClO <sub>4</sub> <sup>-</sup> , NH <sub>4</sub> <sup>+</sup>
Trigonal pyramidal	3	1		107	NCl <sub>3</sub> , PF <sub>3</sub> , ClO <sub>3</sub> , H <sub>3</sub> O <sup>+</sup>
Bent	2	2		104.5	OCl <sub>2</sub> , H <sub>2</sub> S, OF <sub>2</sub> , SCl <sub>2</sub>
Trigonal Bipyramidal	5	0		120 and 90	PCl <sub>5</sub>
Octahedral	6	0		90	SF <sub>6</sub>

### How to explain shape

1. State number of bonding pairs and lone pairs of electrons.
2. State that electron pairs repel and try to get as far apart as possible (or to a position of minimum repulsion.)
3. If there are no lone pairs state that the electron pairs repel equally
4. If there are lone pairs of electrons, then state that lone pairs repel more than bonding pairs.
5. State actual shape and bond angle.

Remember lone pairs repel more than bonding pairs and so reduce bond angles (by about 2.5° per lone pair in above examples)

Occasionally more complex shapes are seen that are variations of octahedral and trigonal bipyramidal where some of the bonds are replaced with lone pairs. You do not need to learn the names of these but ought to be able to work out these shapes using the method below



Square planar  
Bond angle  $90^\circ$

e.g.  $\text{XeF}_4$

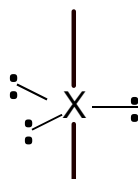


Xe has 8 electrons in its outer shell. 4 F's add 4 more electrons. This makes a total of 12 electrons made up of 4 bond pairs and 2 lone pairs. This means it is a variation of the 6 bond pair shape (octahedral)



Bond angle  $\sim 89^\circ$   
(Reduced by lone pair)

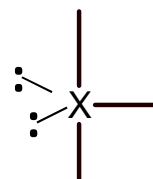
e.g.  $\text{BrF}_5$



Bond angle  $180^\circ$

e.g.  $\text{I}_3^-$

Cl has 7 electrons in its outer shell. 3 F's add 3 more electrons. This makes a total of 10 electrons made up of 3 bond pairs and 2 lone pairs. This means it is a variation of the 5 bond pair shape (trigonal bipyramidal)



Bond angle  $\sim 89^\circ$   
(Reduced by lone pairs)

e.g.  $\text{ClF}_3$



I has 7 electrons in its outer shell. 4 F's add 4 more electrons. Remove one electron as positively charged. This makes a total of 10 electrons made up of 4 bond pairs and 1 lone pair. This means it is a variation of the 5 bond pair shape (trigonal bipyramidal)



Bond angles  $\sim 119^\circ + 89^\circ$   
(Reduced by lone pair)

e.g.  $\text{SF}_4$  &  $\text{IF}_4^+$



## Electronegativity and intermediate bonding

### Definition

**Electronegativity** is the relative tendency of an **atom in a covalent bond** in a molecule to **attract electrons** in a covalent bond to itself.

F, O, N and Cl are the most electronegative atoms

Electronegativity is measured on the **Pauling scale** (ranges from 0 to 4)

The **most** electronegative element is **fluorine** and it is given a value of 4.0

### Factors affecting electronegativity

Electronegativity increases across a period as the **number of protons increases** and the atomic radius decreases because the **electrons in the same shell** are pulled in more.

It decreases down a group because the **distance** between the nucleus and the outer electrons **increases and the shielding** of inner shell electrons increases

### Intermediate bonding

Ionic and covalent bonding are the extremes of a continuum of bonding type. Differences in electronegativity between elements can determine where a compound lies on this scale

A compound containing elements of similar electronegativity and hence a **small electronegativity difference** will be purely **covalent**

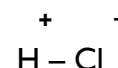
A compound containing elements of very different electronegativity and hence a very **large electronegativity difference** (> 1.7) will be **ionic**

### Formation of a permanent dipole – (polar covalent) bond

A **polar covalent bond** forms when the elements in the bond have **different electronegativities**. (Of around 0.3 to 1.7)

When a bond is a **polar covalent bond** it has an **unequal distribution of electrons** in the bond and produces a **charge separation, (dipole)** + - ends.

The element with the larger electronegativity in a polar compound will be the - end



## Polar and Non Polar molecules

### Symmetric molecules

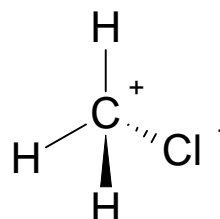
A symmetric molecule (**all bonds identical and no lone pairs**) will **not** be polar **even if individual bonds within the molecular ARE polar**.

The individual dipoles on the bonds '**cancel out**' due to the symmetrical shape of the molecule. There is no **NET** dipole moment: the **molecule is NON POLAR**



CO<sub>2</sub> is a symmetrical molecule and is a non-polar molecule

e.g. CCl<sub>4</sub> will be non-polar whereas CH<sub>3</sub>Cl will be polar



## Intermolecular Forces

### Van der Waals' Forces

Van der Waals forces occur between **all molecular substances** and noble gases. **They do not occur in ionic substances.**

These are also called **transient, induced dipole-dipole interactions**. They occur between all simple covalent molecules and the separate atoms in noble gases. In any molecule the electrons are moving constantly and randomly. As this happens the electron density can fluctuate and parts of the molecule become more or less negative i.e. small temporary or transient dipoles form. These instantaneous dipoles can cause dipoles to form in neighbouring molecules. These are called induced dipoles. The induced dipole is always the opposite sign to the original one.

#### Main factor affecting size of Van der Waals

The **more electrons** there are in the molecule the higher the chance that temporary dipoles will form. This makes the **Van der Waals stronger between the molecules** and so boiling points will be greater.

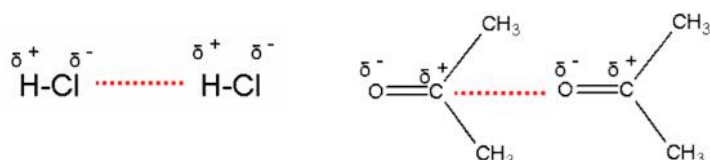
The increasing boiling points of the halogens down the group 7 series can be explained by the increasing number of electrons in the bigger molecules causing an increase in the size of the Van der Waals between the molecules. This is why  $I_2$  is a solid whereas  $Cl_2$  is a gas.

The increasing boiling points of the alkane homologous series can be explained by the increasing number of electrons in the bigger molecules causing an increase in the size of the Van der Waals between molecules.

The shape of the molecule can also have an effect on the size of the Van der Waals forces. Long chain alkanes have a larger surface area of contact between molecules for Van der Waals to form than compared to spherical shaped branched alkanes and so have stronger Van der Waals.

### Permanent dipole-dipole forces

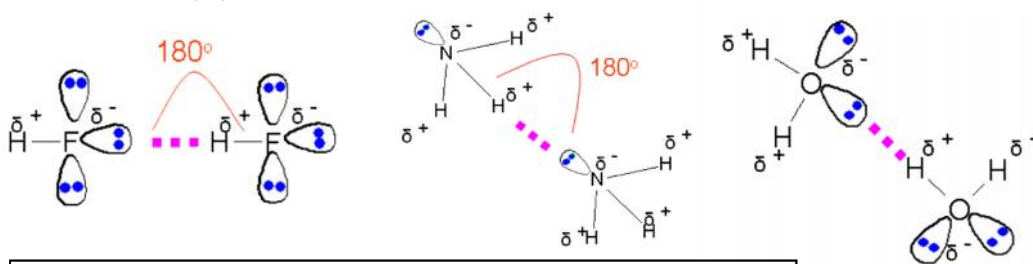
- Permanent dipole-dipole forces occurs between polar molecules
- It is stronger than Van der Waals and so the compounds have higher boiling points
- Polar molecules have a permanent dipole. (commonly compounds with C-Cl, C-F, C-Br H-Cl, C=O bonds)
- Polar molecules are asymmetrical and have a bond where there is a significant **difference in electronegativity** between the atoms.



Permanent dipole-dipole forces occurs in addition to Van der Waals forces

### Hydrogen bonding

It occurs in compounds that have a **hydrogen atom attached to** one of the three **most electronegative** atoms of **nitrogen, oxygen and fluorine**, which must have an available lone pair of electrons. e.g. a **-O-H -N-H F- H** bond. There is a **large electronegativity difference** between the **H and the O,N,F**



Hydrogen bonding occurs in addition to van der waals forces

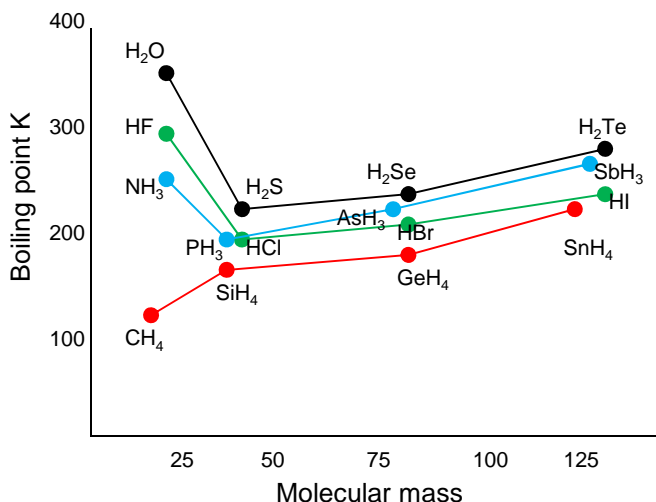
Always show the **lone pair** of electrons on the O,F,N and the dipoles and all the  $\delta^-$   $\delta^+$  charges

Hydrogen bonding is stronger than the other two types of intermolecular bonding.

The **anomalously high** boiling points of  $\text{H}_2\text{O}$ ,  $\text{NH}_3$  and  $\text{HF}$  are caused by the hydrogen bonding between the molecules

The general increase in boiling point from  $\text{H}_2\text{S}$  to  $\text{H}_2\text{Te}$  is caused by increasing Van der Waals forces between molecules due to an increasing number of electrons.

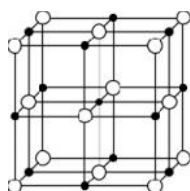
Alcohols, carboxylic acids, proteins, amides all can form hydrogen bonds



### Four types of crystal structure: ionic, metallic, molecular and giant covalent (macromolecular).

You should be able to draw the following diagrams or describe the structure in words to show the four different types of crystal. You should also be able to explain the properties of these solids. The tables earlier in the revision guide explain these properties.

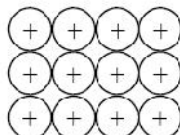
#### Ionic: sodium chloride



Giant ionic lattice showing alternate  $\text{Na}^+$  and  $\text{Cl}^-$  ions

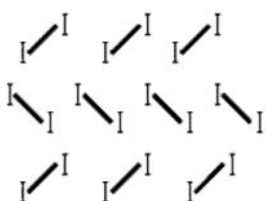
#### Metallic: magnesium or sodium

Use this diagram for any metal



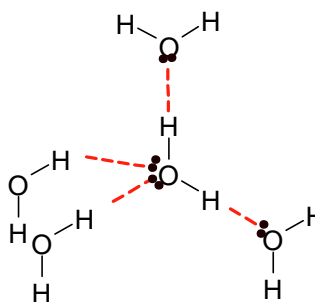
Giant metallic lattice showing close packing magnesium **ions**

#### Molecular: iodine



Regular arrangement of  $\text{I}_2$  molecules held together by weak van der Waals forces

#### Molecular: Ice

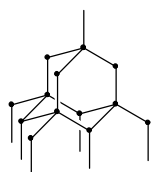


This is a difficult diagram to draw.

The main point to show is a central water molecule with two ordinary covalent bonds and two hydrogen bonds in a tetrahedral arrangement

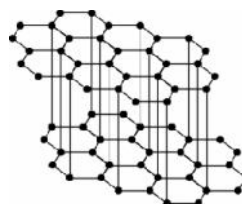
The molecules are held further apart than in liquid water and this explains the lower density of ice

#### Macromolecular: diamond



Tetrahedral arrangement of carbon atoms. 4 covalent bonds per atom

#### Macromolecular: Graphite



Planar arrangement of carbon atoms in layers. 3 covalent bonds per atom in each layer. 4<sup>th</sup> outer electron per atom is delocalised. Delocalised electrons between layers.

Both these macromolecular structures have very high melting points because of strong covalent forces in the giant structure. It takes a lot of energy to break the **many strong** covalent bonds.