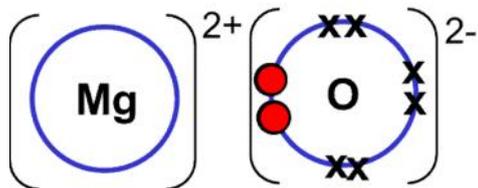


3 Bonding

Definition: An ionic bond 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.



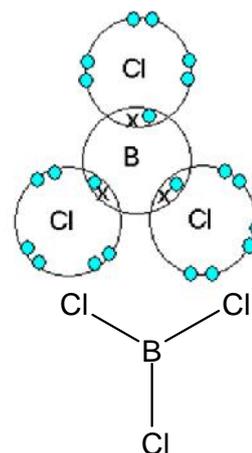
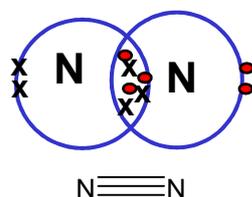
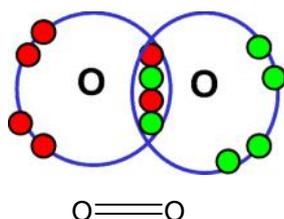
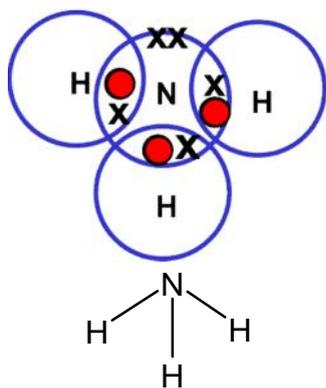
Mg goes from $1s^2 2s^2 2p^6 3s^2$
to $Mg^{2+} 1s^2 2s^2 2p^6$

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^-)

Covalent Bonding

Definition: covalent bond is the strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms



The term **bond energy** is used as a measurement of covalent bond strength. It is the energy needed to break one mole of bonds of (gaseous covalent) bonds into gaseous atoms

The larger the value of the average bond enthalpy, the stronger the covalent bond

Bond length measures the distance between the two nuclei in a covalent bond.

A shorter bond will have a higher bond energy as there will be a stronger force of attraction between the nuclei and the shared pair of electrons as they are closer together.

Effect of multiple bonds on bond strength and length.

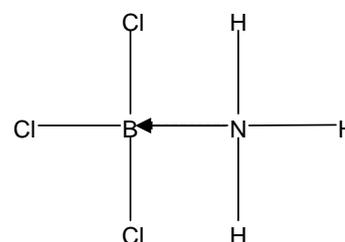
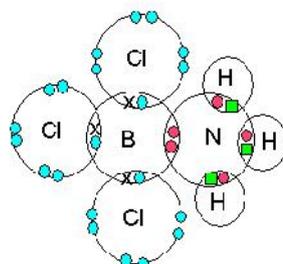
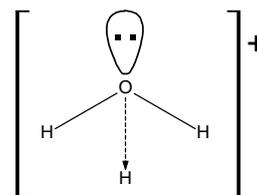
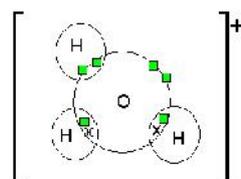
Nuclei joined by multiple (i.e. double and triple) bonds have a **greater electron density** between them.

This causes an **greater force of attraction** between the nuclei and the electrons between them, resulting in a **shorter bond length** and **greater bond strength**.

Dative Covalent bonding

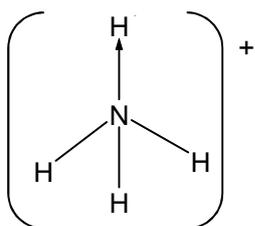
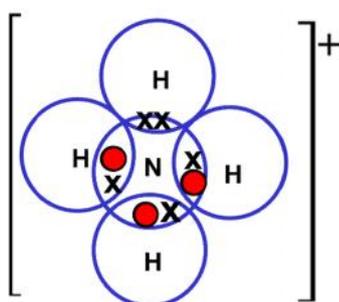
A **Dative covalent bond** forms when the **shared pair of electrons** in the covalent bond come from **only one of the bonding atoms**. A **dative covalent bond** is also called co-ordinate bonding.

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

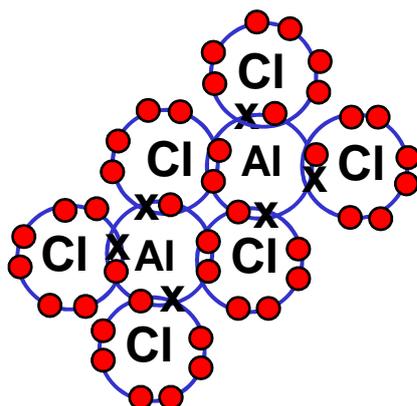


The direction of the arrow goes from the atom that is providing the lone pair to the atom that is deficient

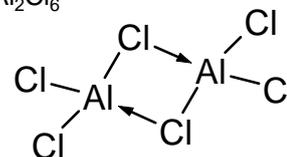
The ammonium ion NH_4^+



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

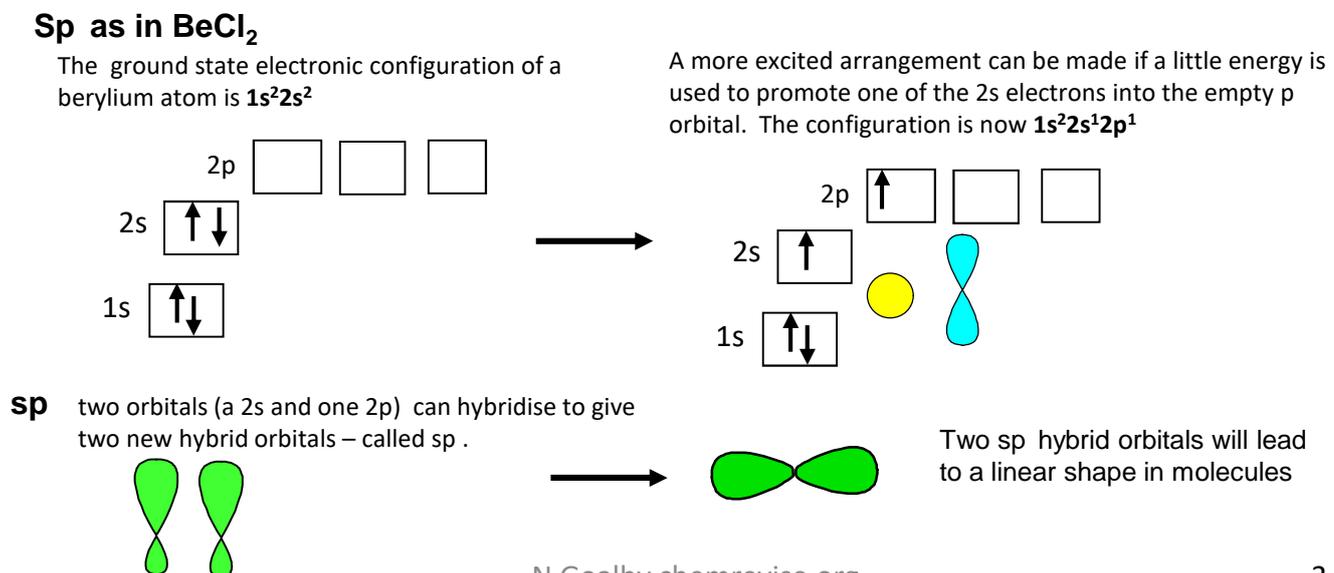
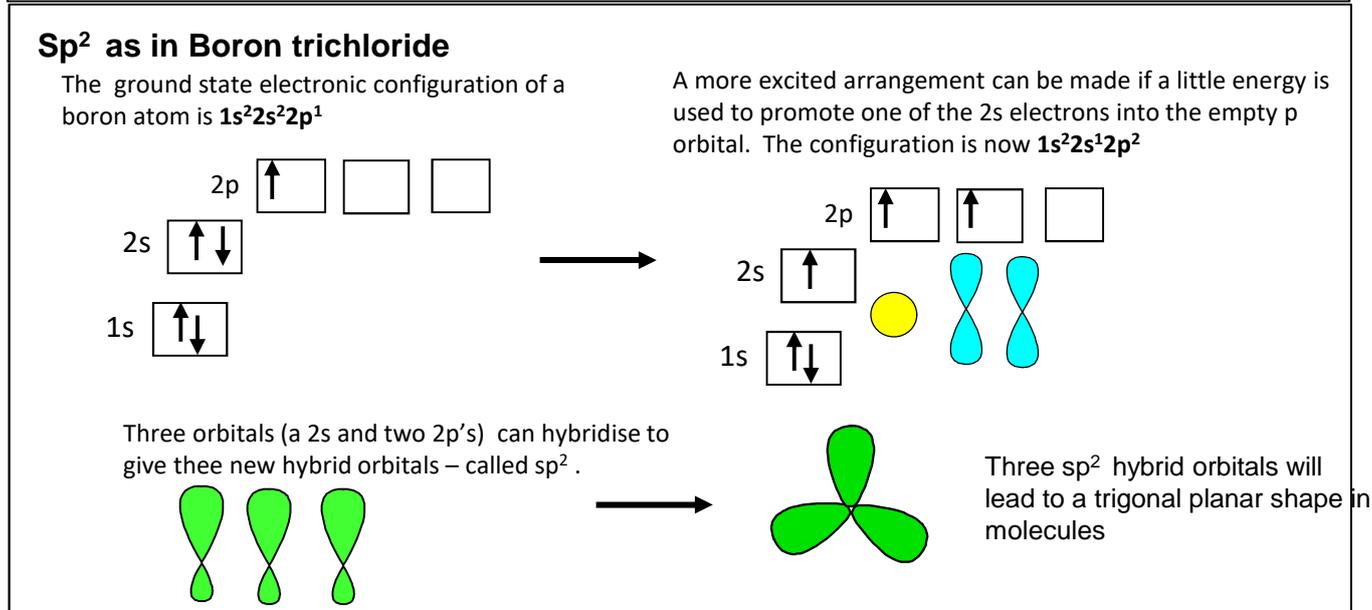
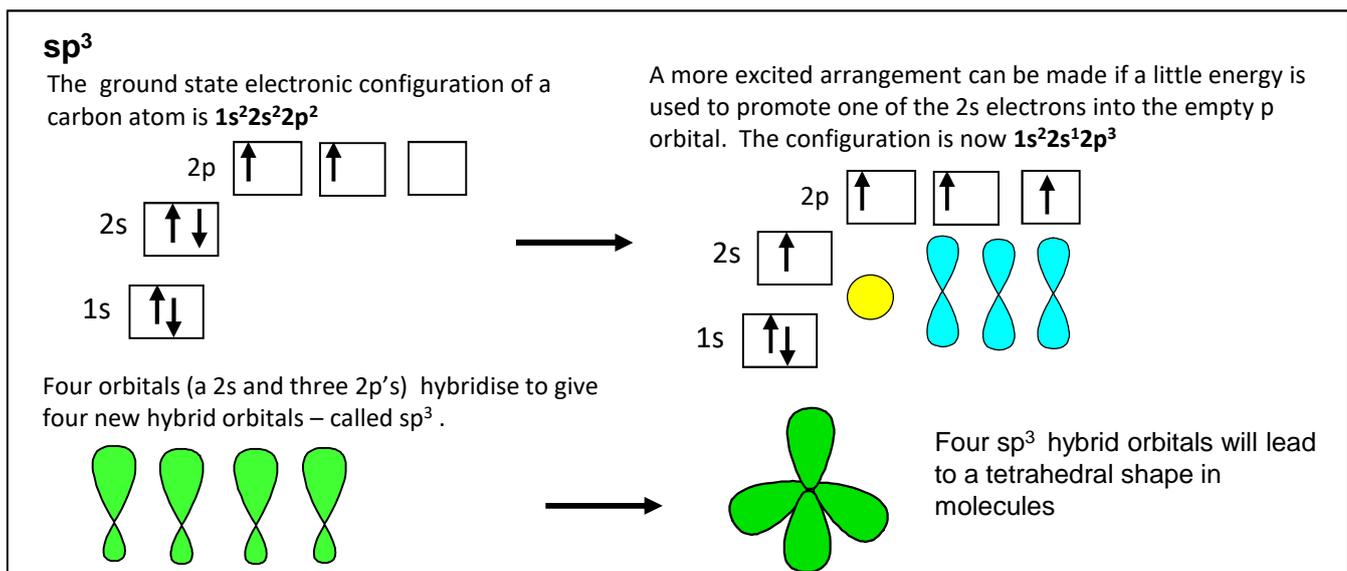


Two aluminium chloride (AlCl_3) molecules join together through two dative bonds to form the dimer Al_2Cl_6



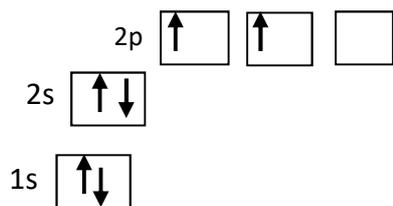
Covalent bonding in terms of orbital overlap

When atomic orbitals overlap they form molecular orbitals. The model can be used to show how single and double bonds form and can be used to show how why molecules have the shape they do.

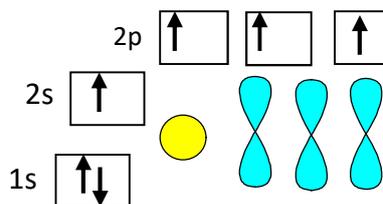


Forming multiple bonds. E.g the C=C bond

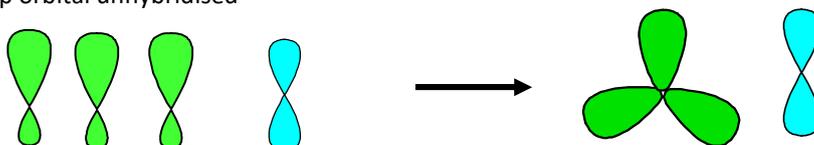
The ground state electronic configuration of a carbon atom is $1s^2 2s^2 2p^2$



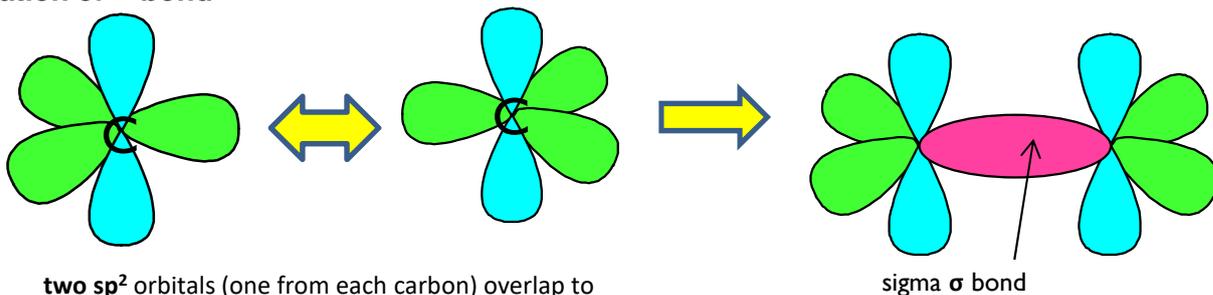
A more excited arrangement can be made if a little energy is used to promote one of the 2s electrons into the empty p orbital. The configuration is now $1s^2 2s^1 2p^3$



Three orbitals (a 2s and two 2p's) can hybridise to give three new hybrid orbitals – called sp^2 . Leaving one p orbital unhybridised



Formation of σ bond

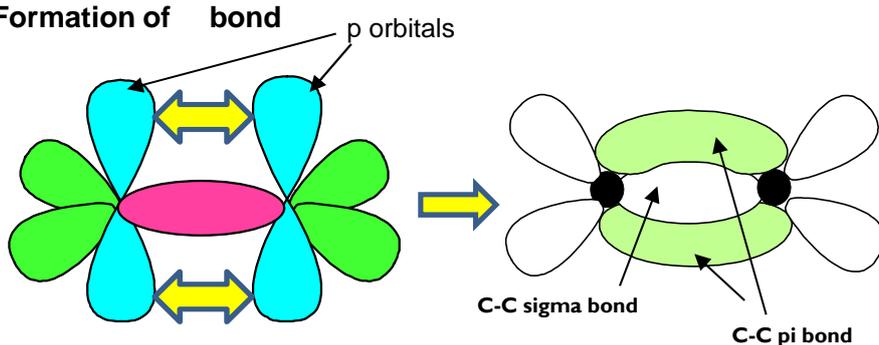


two sp^2 orbitals (one from each carbon) overlap to form a single C-C bond called a sigma σ bond

sigma σ bond

Rotation can occur around a sigma bond

Formation of π bond



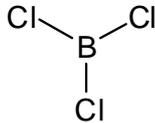
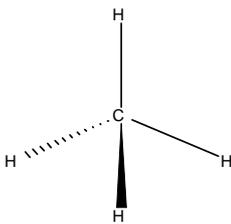
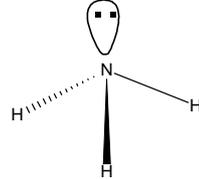
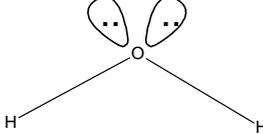
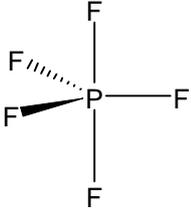
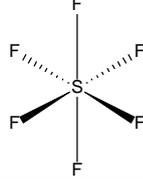
C-C sigma bond

C-C pi bond

The π bond is formed by sideways overlap of **two p orbitals** on each carbon atom forming a π -bond above and below the plane of molecule.

The π bond is weaker than the σ bond.

Shape of molecules

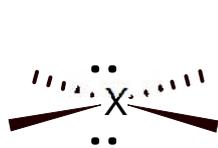
Name	No bonding pairs	No lone pairs	Diagram	Bond angle	Examples
linear	2	0	$\text{Cl}-\text{Be}-\text{Cl}$	180	CO_2 , CS_2 , HCN , BeF_2
Trigonal planar	3	0		120	BF_3 , AlCl_3 , SO_3 , NO_3^- , CO_3^{2-}
Tetrahedral	4	0		109.5	SiCl_4 , SO_4^{2-} , ClO_4^- , NH_4^+
Trigonal pyramidal	3	1		107	NCl_3 , PF_3 , ClO_3 , H_3O^+
Bent	2	2		104.5	OCl_2 , H_2S , OF_2 , SCl_2
Trigonal Bipyramidal	5	0		120 and 90	PCl_5
Octahedral	6	0		90	SF_6

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°

e.g. XeF_4

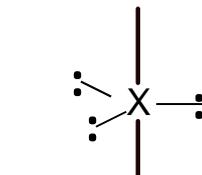


Xe has 8 electrons in its outer shell. 4 F's add 4 more electrons. This makes a total of 12 electrons made up of 4 bond pairs and 2 lone pairs. This means it is a variation of the 6 bond pair shape (octahedral)



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

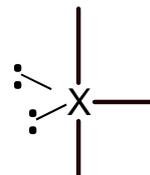
e.g. BrF_5



Bond angle 180°

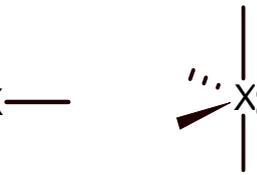
e.g. I_3^-

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. ClF_3



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

e.g. SF_4 & IF_4^+



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)

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**

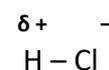
A compound containing elements of very different electronegativity and hence a very **large electronegativity difference** (> 1.7) will be **ionic**

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)** δ^+ δ^- ends. This charge separation is called bond polarity

The element with the larger electronegativity in a polar compound will be the δ^- 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.**

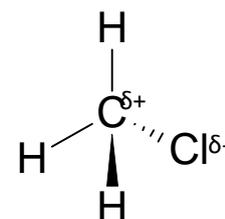
The individual dipoles on the bonds 'cancel out' due to the symmetrical shape of the molecule.

There is no **NET** dipole moment: the **molecule is NON POLAR**

e.g. CCl_4 will be non-polar whereas CH_3Cl will be polar



CO_2 is a symmetrical molecule and is a non-polar molecule



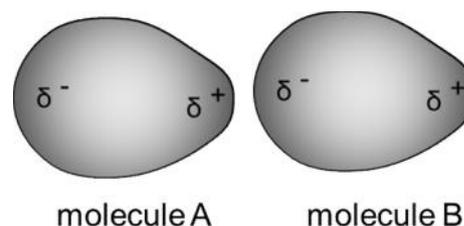
Intermolecular bonding

Induced dipole–dipole interactions

Induced dipole–dipole interactions occur between **all molecular substances** and noble gases. **They do not occur in ionic substances.**

Induced dipole–dipole interactions are also called **London forces**. 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 temporary 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 Induced dipole–dipole interactions

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

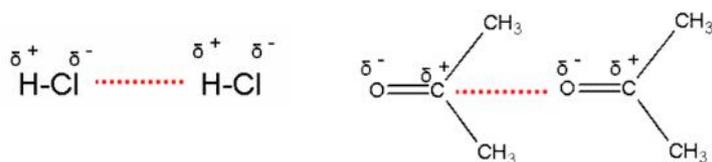
The increasing boiling points of the halogens down the group 7 series can be explained by the increasing number of electrons in the bigger molecules causing an increase in the size of the induced dipole–dipole interactions between the molecules. This is why I_2 is a solid whereas Cl_2 is a gas.

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 induced dipole–dipole interactions between molecules.

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

Permanent dipole-dipole forces

- Permanent dipole-dipole forces occur 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.

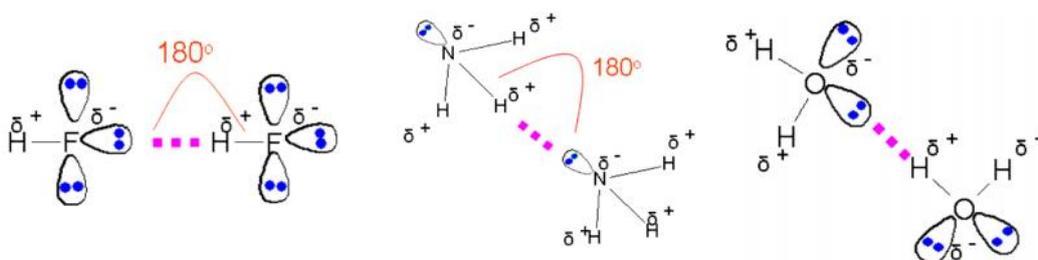


Permanent dipole forces occur in addition to induced dipole-dipole interactions

Permanent dipole-dipole and induced dipole-dipole interactions can **both** be referred to as 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**



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

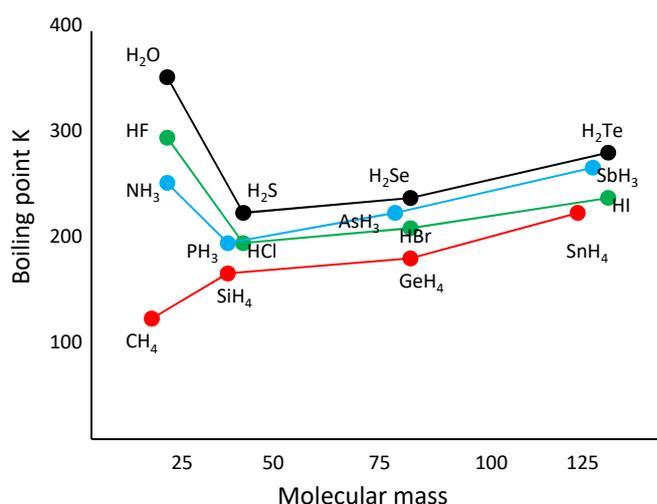
Hydrogen bonding occurs in addition to van der Waals forces

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

The **anomalously high** boiling points of H_2O , NH_3 and HF are caused by the hydrogen bonding between the molecules

The general increase in boiling point from H_2S to H_2Te is caused by increasing induced dipole-dipole interactions between molecules due to an increasing number of electrons.

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



Metallic bonding

Definition: A **metallic bond** is the **electrostatic force** of attraction between the **positive metal ions** and the **delocalised electrons**

The three main factors that affect the strength of a **metallic bond** 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.

Chemical reactions and bonds

Chemical reactions involve breaking and making bonds
Breaking bonds absorbs energy – is endothermic
Making bonds gives out energy – is exothermic

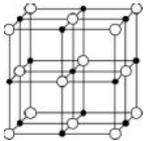
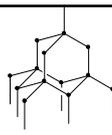
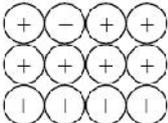
Whether a chemical reaction occurs can be dependent on the relative sizes of the bonds being broken and made.

If reactants have strong bonds they will be harder to break and this leads to a high activation energy.

In many organic reactions it is the weakest bond in the molecule that breaks.

The bond polarity of the bond can also have an effect. An attacking species might be attracted to an atom with a dipole of opposite charge

Bonding and Structure

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

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

Property	Ionic	Molecular (simple)	Macromolecular	Metallic
boiling and melting points	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
Solubility in water	Generally good	generally poor	insoluble	insoluble
conductivity when solid	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
conductivity when molten	good: ions can move	poor: no ions	poor	(good)
general description	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