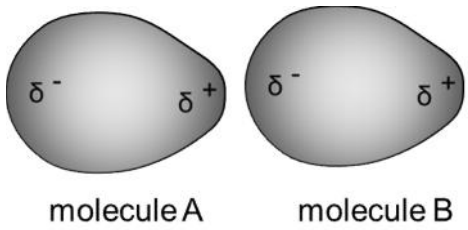
7: Intermolecular Forces

London Forces

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

London Forces are also called **instantaneous, 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 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 London Forces
The **more electrons** there are in the molecule the higher the chance that temporary dipoles will form. This makes the **London forces stronger between the molecules** and more energy is needed to break them so boiling points will be greater.

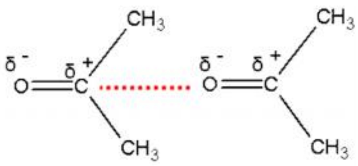
The increasing boiling points of the halogens down the group 7 series can be explained by the increasing number of electrons in the bigger molecules causing an increase in the size of the London forces between the molecules. This is why I_2 is a solid whereas Cl_2 is a gas.

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 London forces between molecules.

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

Permanent dipole-dipole forces

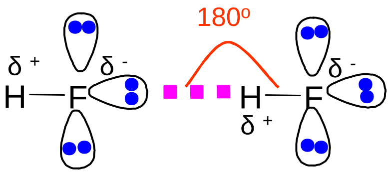
- Permanent dipole-dipole forces occurs between polar molecules
- It is stronger than London forces and so the compounds have higher boiling points
- Polar molecules have a permanent dipole. (commonly compounds with C-Cl, C-F, C-Br H-Cl, C=O bonds)
- Polar molecules are asymmetrical and have a bond where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

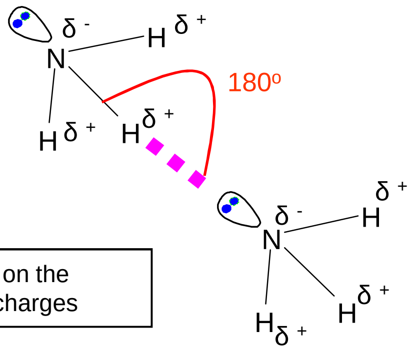


Permanent dipole forces occur in addition to London forces

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

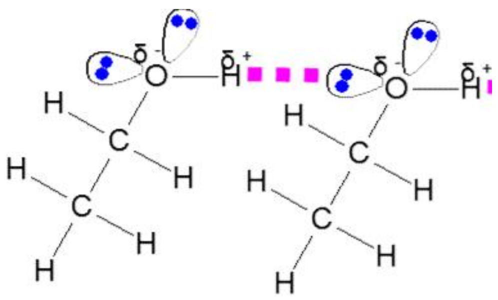
Hydrogen bonding occurs in addition to London forces

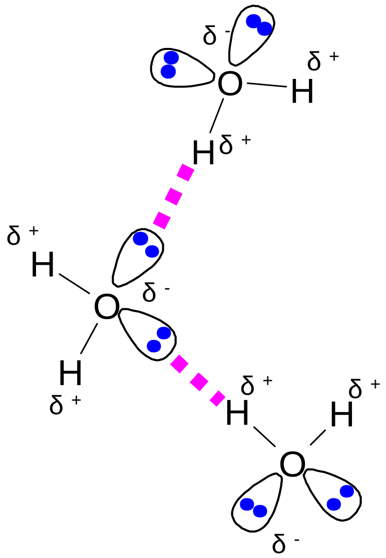
The hydrogen bond should have an bond angle of 180° with one of the bonds in one of the molecules

The bond angle is 180° around the H atom because there are two pairs of electrons around the H atom involved in the hydrogen bond. These pairs of electrons repel to a position of minimum repulsion, as far apart as possible.

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

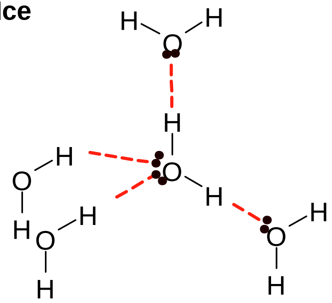
Alcohols form hydrogen bonds. This means alcohols have higher boiling points and relatively low volatility compared to alkanes with a similar number of electrons





Water can form two hydrogen bonds per molecule, because the electronegative oxygen atom has two lone pairs of electrons on it.
It can therefore form stronger hydrogen bonding and needs more energy to break the bonds, leading to a higher boiling point.

Ice

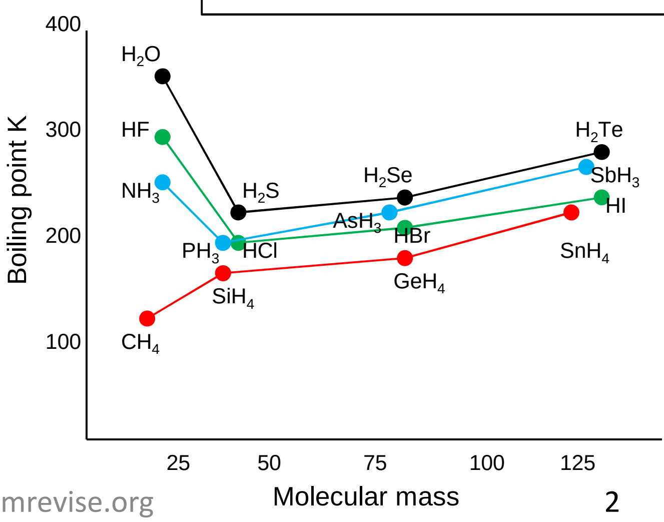


In ice the molecules are held further apart by the hydrogen bonds than in liquid water and this explains the lower density of ice

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 these molecules in addition to their London forces. The additional forces require more energy to break and so have higher boiling points

The general increase in boiling point from H_2S to H_2Te or from HCl to HI is caused by increasing London forces between molecules due to an increasing number of electrons.



Solvents and Solubility

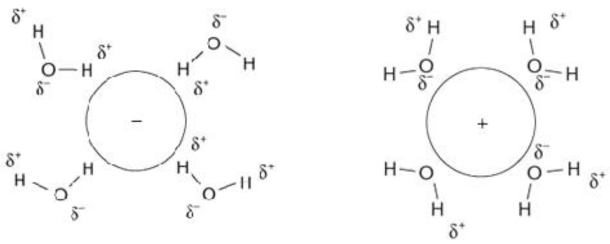
Solubility of a solute in a solvent is a complicated balance of energy required to break bonds in the solute and solvent against energy given out making new bonds between the solute and solvent.

Ionic substances dissolving in water

When an ionic lattice dissolves in water it involves breaking up the bonds in the lattice and forming new bonds between the metal ions and water molecules.

The **negative** ions are attracted to the **δ⁺ hydrogens** on the **polar water** molecules and the positive ions are attracted to the δ⁻ oxygen on the polar water molecules.

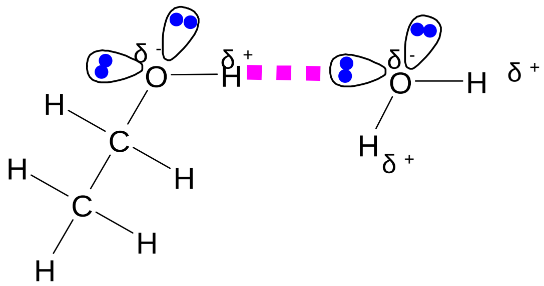
hydration of the ions



The higher the **charge density** the greater the hydration enthalpy (e.g. **smaller ions** or **ions with larger charges**) as the ions attract the water molecules more strongly.

Solubility of simple alcohols

The smaller alcohols are soluble in water because they can form hydrogen bonds with water. The longer the hydrocarbon chain the less soluble the alcohol.



Insolubility of compounds in water

Compounds that cannot form hydrogen bonds with water molecules, e.g. polar molecules such as halogenoalkanes or non polar substances like hexane will be insoluble in water.

Solubility in non-aqueous solvents

Compounds which have similar intermolecular forces to those in the solvent will generally dissolve

Non-polar solutes will dissolve in non-polar solvents. e.g. Iodine which has only London forces between its molecules will dissolve in a non polar solvent such as hexane which also only has London forces.

Propanone is a useful solvent because it has both polar and non polar characteristics. It can form London forces with some non polar substances such as octane with its CH_3 groups. Its polar C=O bond can also hydrogen bond with water.