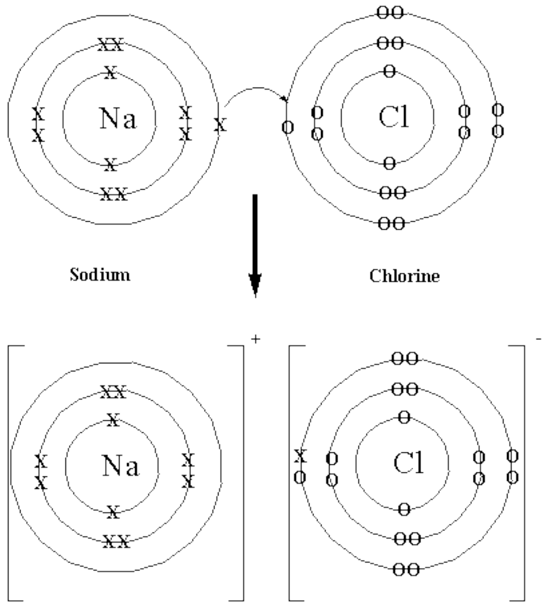
Bonding and Structure

Ionic Bonding

An ionic bond is the electrostatic attraction between oppositely charged ions.

Ions are formed when atoms lose or gain electrons. They have a noble gas electronic structure (a full outer electron shell). This makes them very stable because the effect of shielding is minimised.

Ionic bonding can be shown using dot and cross diagrams:



Sodium Chloride

An ionic lattice is a 3d structure with **cations** (positive ions) and **anions** (negative ions) closely packed together in a regular repeating pattern.

The ionic charge can be predicted from an element's position in the Periodic table:

Group 1 metals form ions with a single + charge, e.g. Na^+
Group 2 metals form ions with a double + charge, e.g. Mg^{2+}
Group 3 metals form ions with a triple + charge, e.g. Al^{3+}

Group 7 non-metals gain 1 electron to give ions with a single negative charge, e.g. Cl^-
Group 6 non-metals gain 2 electrons to give ions with a double negative charge, e.g. O^{2-}
Group 5 elements can gain 3 electrons to give ions with a triple negative charge, e.g. N^{3-} but this is rare.

Formulae of polyatomic ions

Nitrate NO_3^-
Carbonate CO_3^{2-}
Sulfate SO_4^{2-}
Ammonium NH_4^+

Covalent bonding and dative covalent bonding

A covalent bond is a shared pair of electrons. Covalent bonding occurs when two orbitals overlap. A σ (sigma) bond is made - there is a dense electron cloud between the nuclei.

Normally each orbital contains one electron, but in **dative covalent** bonding, an empty orbital overlaps with a full orbital.

A dative covalent bond is a shared pair of electrons, originating from one atom.

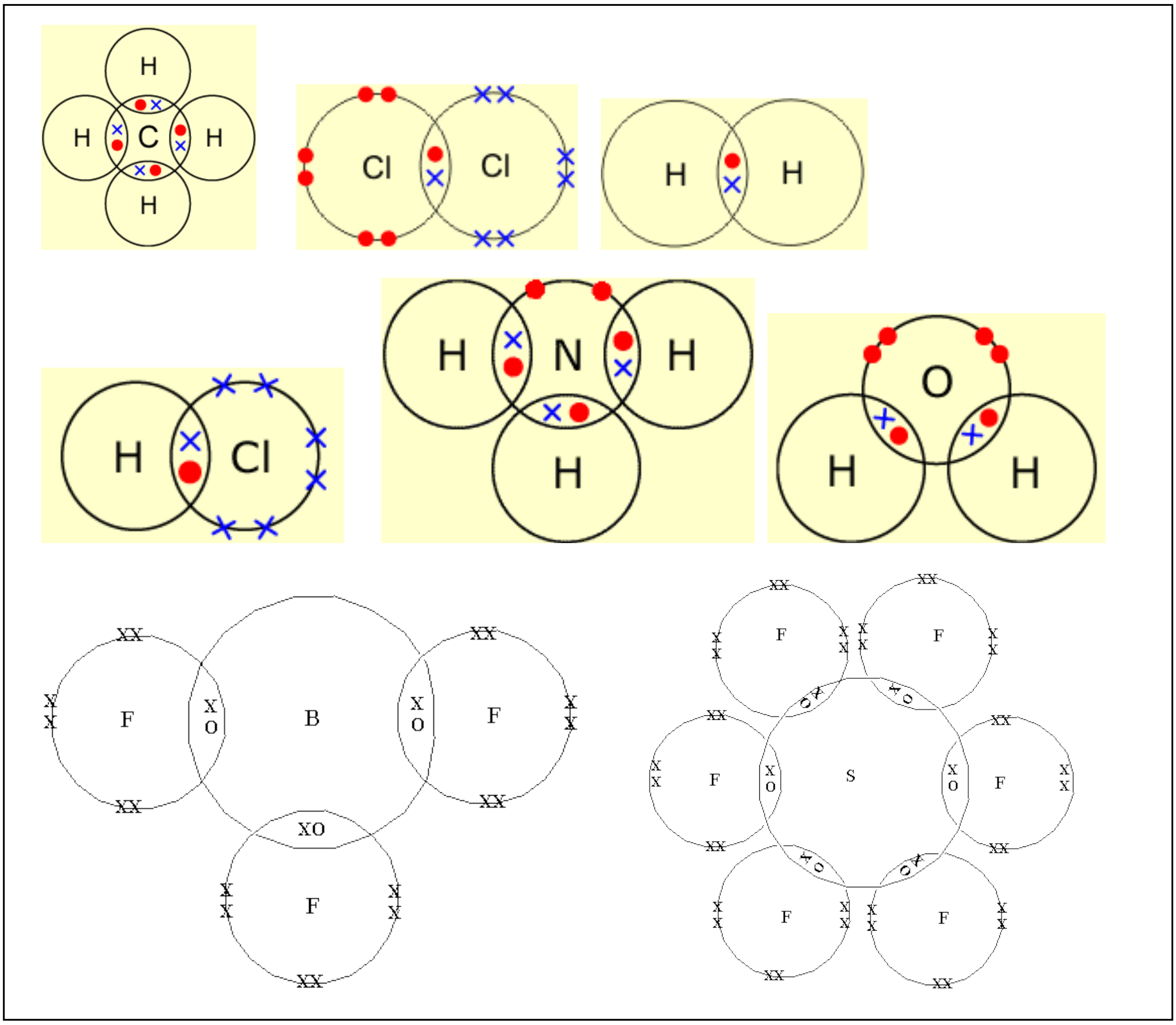
A molecule has two or more atoms linked together by covalent bonds.

A lone pair is a pair of outer electrons that does not take part in bonding.

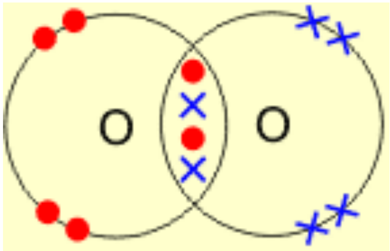
A double bond occurs when two pairs of electrons are shared between the same two atoms. The two bonds are a σ bond and a π bond.

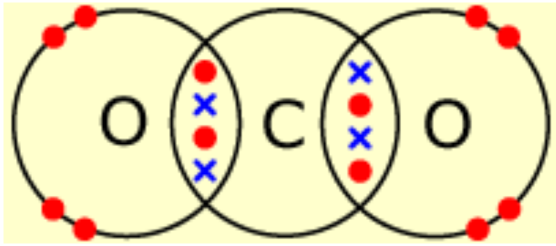
Dot and cross diagrams:

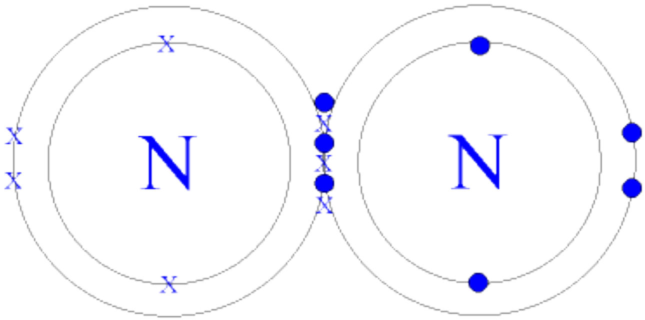
1. Single covalent bonding



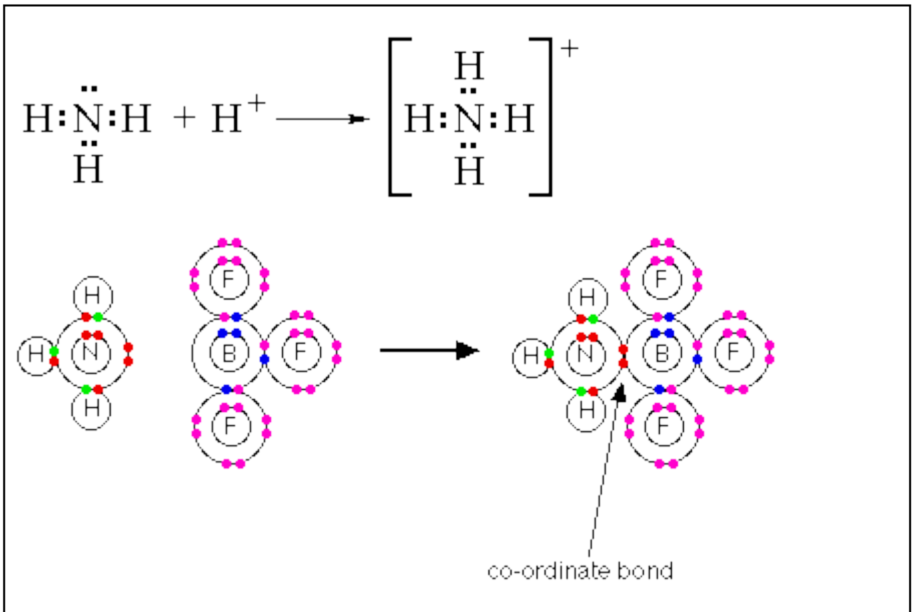
2. Multiple covalent bonding







3. Dative covalent bonding



The shapes of simple molecules and ions

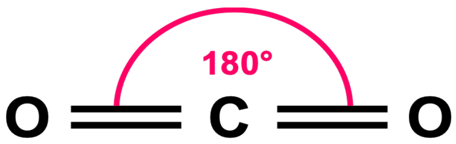
Shapes of molecules show the angles between the different bonds. **The electron pairs (bonds or lone pairs) around the central atom repel each other so they will be as far apart as possible.**

Linear molecules	angle 180º
Planar molecules	angle 120º
Tetrahedral molecules	angle 109.5º
Octahedral molecule	angle 90º

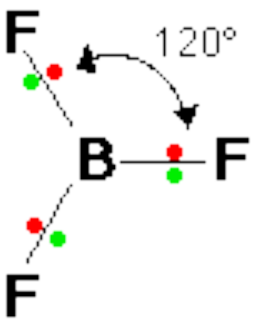
Lone pairs repel more than bonds so angles between the bonds are reduced. However, remember they are invisible, e.g. a tetrahedron with one lone pair is pyramidal.

Examples:

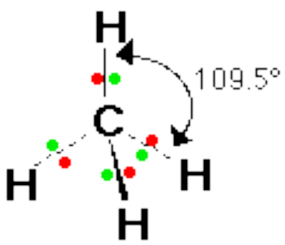
CO_2 – linear

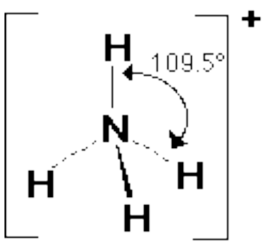


BF_3 – trigonal planar

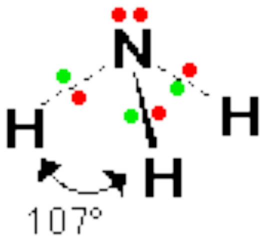


CH_4 and NH_4^+ - tetrahedral

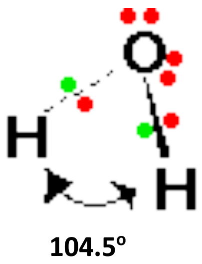




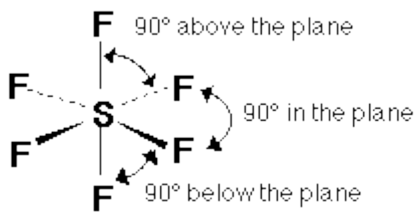
NH_3 – pyramidal



H_2O – non-linear



SF_6 – octahedral



You should be able to predict the shapes of, and bond angles in, molecules and ions analogous to those shown above.

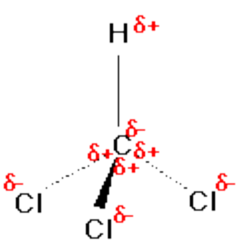
Electronegativity and Bond Polarity

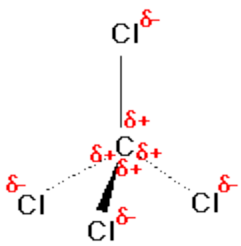
Electronegativity is the ability of an atom to attract the electrons in a covalent bond.

A **permanent dipole** may arise when covalently bonded atoms have different electronegatives, resulting in a polar bond.

Chlorine is more electronegative than hydrogen so the electron cloud of the bond in HCl is closer to the Cl than the H: the bond is **polar** - it has a **permanent dipole**. The charge distribution is not even so this covalent bond has some ionic character.

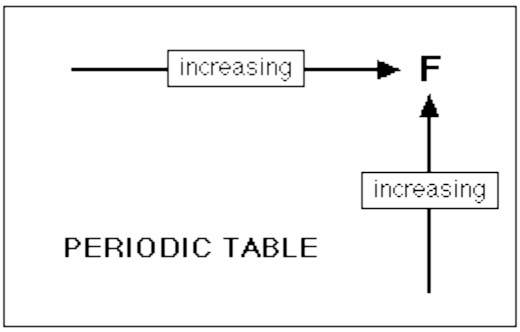
NB Not all molecules with polar bonds are polar, e.g.$CHCl_3$ is polar because it is not symmetrical so the dipoles do not cancel, the charge distribution is uneven.





CCl_4 is non polar because it is symmetrical so the dipoles cancel, there is an even charge distribution.

Trends in Electronegativities:



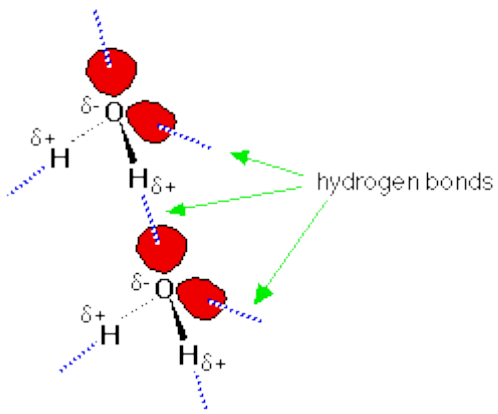
Intermolecular Forces

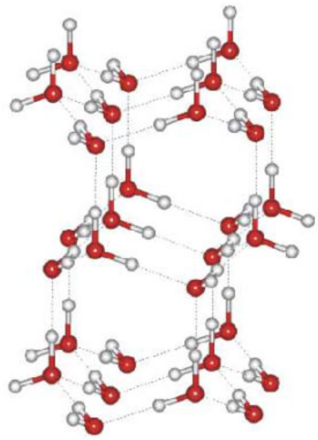
Intermolecular Forces are weak attractions between molecules: they occur when molecules have dipoles. HCl has a permanent dipole so there will be electrostatic attraction between the molecules - this is known as a **dipole - dipole attraction**

All molecules form temporary instantaneous dipoles which then induce dipoles in the surrounding molecules. This leads to electrostatic attractions called **van der Waals' forces**. These involve all the electrons in the molecule, so larger molecules, with more electrons have stronger Van der Waals' forces.

Noble gases- as relative mass increases, electron cloud becomes larger, VdW become stronger - boiling point increases as more energy is needed to overcome the attractive forces.

Hydrogen Bonds are stronger intermolecular forces but are much weaker than covalent bonds. They occur between hydrogen atoms which are connected to oxygen or nitrogen, and oxygen or nitrogen atoms in other molecules. They are particularly important in water which can have 4 hydrogen bonds for each molecule.



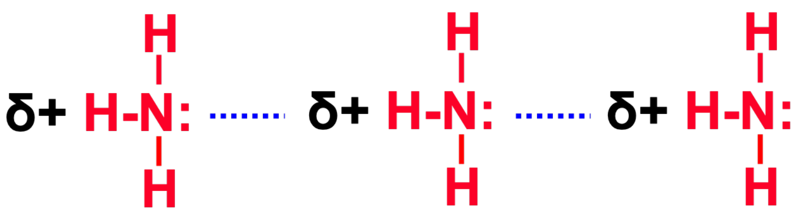


Water has a **high boiling and melting point** for such a small molecule (compare it with CH_4 or CO) because a large amount of energy is needed to overcome the hydrogen bonds.

Ice floats on water because the solid structure is held together by long hydrogen bonds: they give an open lattice which is less dense than the liquid. Most liquids become denser when they freeze.

You should be able to apply hydrogen bonding ideas to analogous molecules containing the groups O-H and N-H.

Hydrogen bonding in ammonia

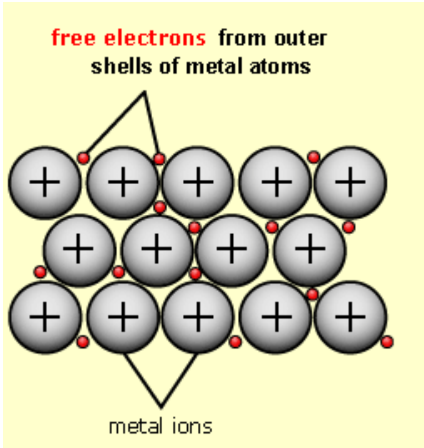


<u>Metallic Bonding</u>

Metallic Bonding occurs in all metals. It is the attraction between a lattice of positive ions and the surrounding sea of delocalised electrons. The structure is flexible because the lattice and the

delocalised electrons can move relative to each other - this explains the characteristic properties of metals:- conductivity, malleability etc.
The structure is described as a **Giant Metallic Lattice.** Metals conduct electricity because the electrons can move.

Metallic bonding can be thought of as a type of ionic bonding with electrons rather than negative ions.



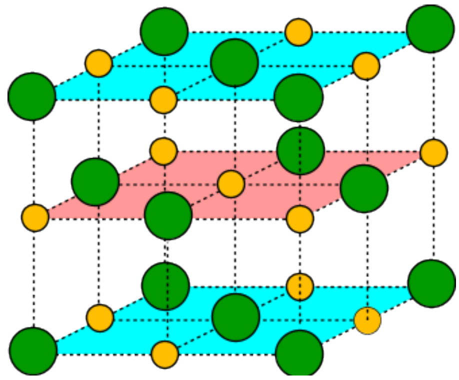
The metal is held together by the strong forces of attraction between the positive nuclei and the delocalised electrons and so metals generally have high melting and boiling points.

Metallic bond strength depends on

- The number of outer electrons donated
- The size of the metal ion

<u>Bonding and Physical Properties</u>

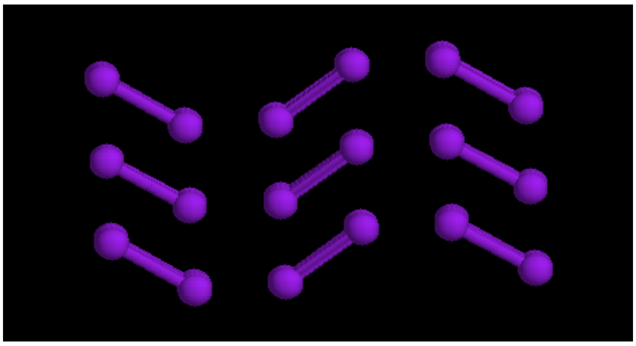
1. **Giant Ionic Lattices** have a large number of ions held together by strong ionic bonds in a 3d lattice. They are hard crystalline solids with high melting and boiling points because a great deal of energy is needed to break the strong forces. Ionic substances do not conduct electricity when solid because the ions cannot move. They can move when dissolved in water or when molten, so can then conduct electricity.
   Many, but not all, are soluble in polar solvents such as water because the polar water molecules break down the lattice by surrounding each ion to form a solution. Positive ions are attracted to the lone pairs on water molecules. Water molecules form hydrogen bonds with negative ions.



2. **Giant Metallic Lattices** – see above.

3. **Simple Molecular Lattice these** have molecules containing strong covalent bonds, linked together by weak forces (Van Der Waals', dipole-dipole or Hydrogen bonding).

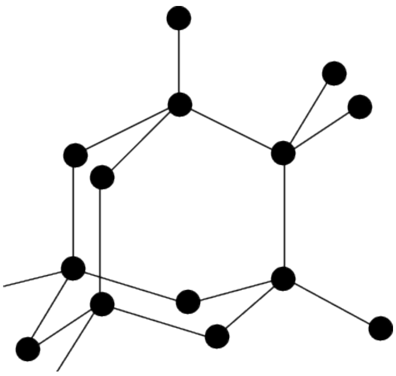
   These structures have low melting points and boiling points because only the weak forces are broken. They do not conduct electricity as they have no free electrons. Examples: simple molecular lattices as in I_2 and ice.
   They are soluble in non-polar solvents such as hexane because van der Waals' forces form between the molecules in the structure and the solvent.



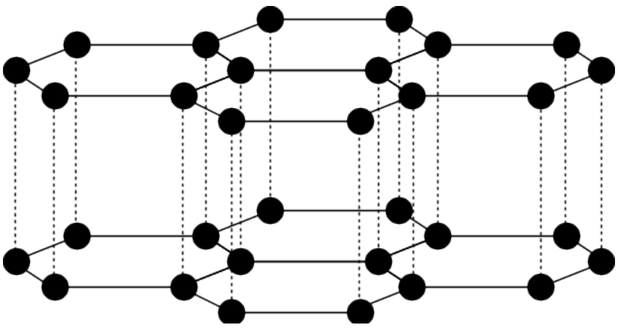
4. **Giant Covalent Lattices**

Diamond and Graphite are forms of carbon. They are both giant covalent lattices containing carbon atoms joined together by strong covalent bonds so they both have very high melting and boiling points. They are both insoluble in polar and non-polar solvents because the covalent bonds are too strong to be broken by either polar or non-polar solvent molecules.

Diamond has a 3 d structure with each atom bonded to four others. It does not conduct electricity as it has no free electrons.



Graphite has a layer structure, with strong covalent bonds within the layers but weak van der Waals' forces between the layers.



Graphite is much softer than diamond, the layers slide over each other quite easily (pencil lead). Graphite conducts electricity as it has electrons between its layers which can move.