

Bonding

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.

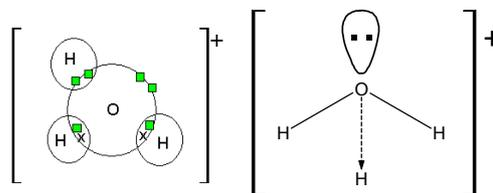
Mg goes from $1s^2 2s^2 2p^6 3s^2$
to $Mg^{2+} 1s^2 2s^2 2p^6$

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

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

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)

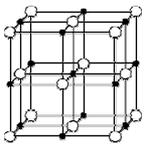
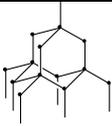
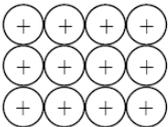


A **metallic bond** is the **electrostatic force** of attraction between the **positive metal ions** and the **delocalised electrons**

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.

Bonding and Structure

Bonding	Structure	Examples
Ionic : electrostatic force of attraction between oppositely charged ions	 Giant Ionic Lattice	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. 	Diamond Graphite Silicon dioxide Silicon
Metallic: electrostatic force of attraction between the metal positive ions and the delocalised electrons	 Giant metallic lattice	Magnesium (all metals)

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

Property	Ionic	molecular	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)	high- because of strong covalent forces in giant 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

Electronegativity and intermediate bonding

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 increases across a period as the **number of protons increases** 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

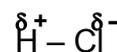
A compound containing elements of similar electronegativity and hence a **small electronegativity difference** will be **covalent**

A **covalent bond** with elements of **different electronegativities** will be **polar covalent bond**. This results in an **unequal distribution of electrons** in the bond and produces a **charge separation, (dipole)** δ^+ δ^- ends.

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

If all the bonds in a compound are the same polar bond and there are no lone pairs then the dipoles cancel out and the substance will be non polar.

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

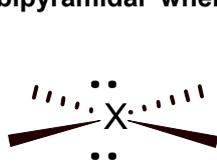


A compound containing elements of very different electronegativity and hence a very **large electronegativity difference** will be **ionic**

Shape of molecules

Name	No bonding pairs	No lone pairs	Diagram	Bond angle	Examples
linear	2	0		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 ₃ O ⁺
Bent line	2	2		104.5	OCl ₂ , H ₂ S, OF ₂ , SCl ₂
Trigonal Bipyramidal	5	0		120 and 90	PCl ₅
Octahedral	6	0		90	SF ₆

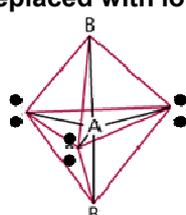
Occasionally more complex shapes are seen that are variations of octahedral and trigonal bipyramidal where some of the bonds are replaced with lone pairs.



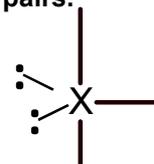
e.g XeF₄



e.g. BrF₅



e.g I₃



e.g. ClF₃



e.g. SF₄

Remember lone pairs repel more than bonding pairs and so reduce bond angles

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.

Intermolecular bonding

Van der Waals' Forces

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.

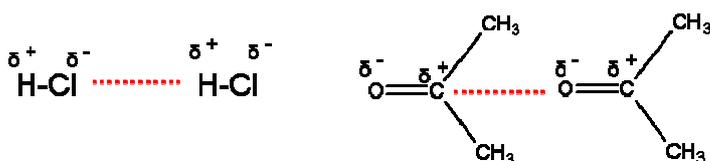
These instantaneous dipoles can cause dipoles to form in neighbouring molecules. These are called induced dipoles. The induced dipole is always the opposite sign to the original one.

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

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

Permanent dipole-dipole bonding

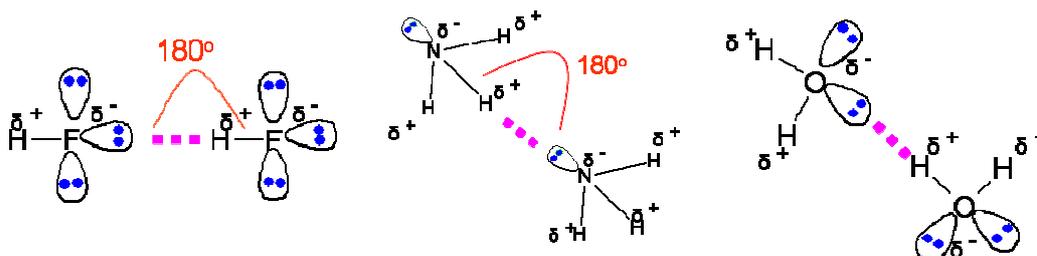
- Permanent dipole-dipole bonding occurs between polar molecules (stronger than van der waals)
- Polar molecules have a permanent dipole. (commonly compounds with C-Cl, C=O bonds)
- Polar molecules are asymmetrical and have a bond where there is a significant **difference in electronegativity** between the atoms.



Permanent dipole bonding 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



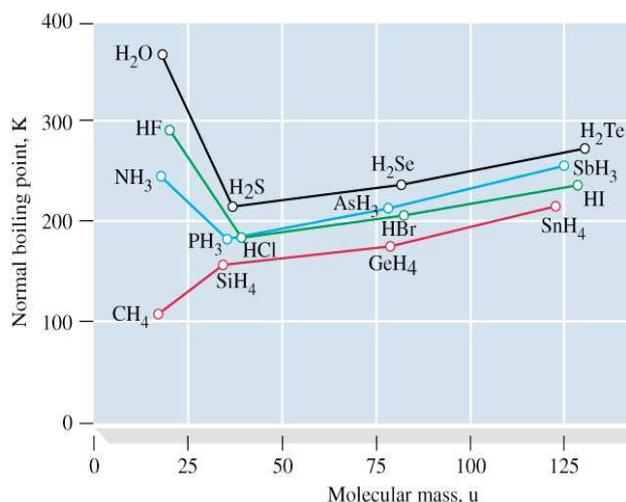
Always show the **lone pair** of electrons on the O,F,N and the dipoles

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.

The general increase in boiling point from H_2S to H_2Te is caused by increasing van der Waals due to an increasing number of electrons.

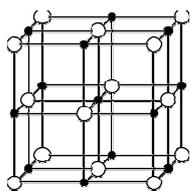
Alcohols, carboxylic acids, proteins, amides all hydrogen bond



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

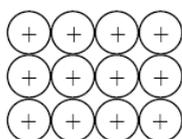
You should be able to draw the following diagrams or describe the structure in words to show the four different types of crystal. You should also be able to explain the properties of these solids. The tables earlier in the revision guide explain these properties

Ionic: sodium chloride



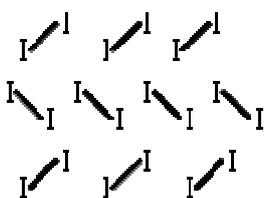
Giant Ionic lattice showing alternate Na^+ and Cl^- ions

Metallic: magnesium



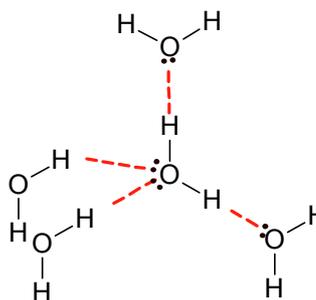
Giant metallic lattice showing close packing magnesium **ions**

Molecular: Iodine



Regular arrangement of I_2 molecules held together by weak van der Waals forces

Molecular: Ice

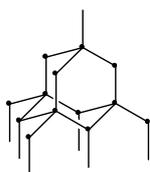


This is a difficult diagram to draw.

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

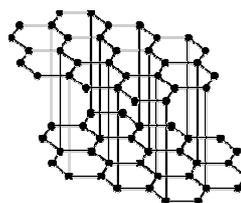
The molecules are held further apart than in liquid water and this explains the lower density of ice

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. 4th outer electron per atom is delocalised. Delocalised electrons between layers.

Both these 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