

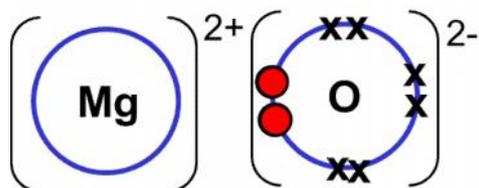
## 2.2.2 Bonding and Structure

### Ionic 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.

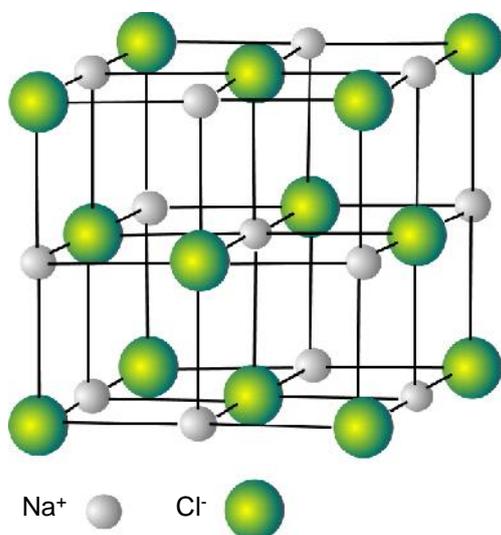
Non-metal atoms gain electrons to form -ve ions.



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,  $Na^+$  &  $Cl^-$ )



The ions in an ionic solid are arranged in a regular 3D pattern called a giant ionic lattice

The sticks in this diagram are there to help show the arrangements of the ions. They do not represent the ionic bonds.

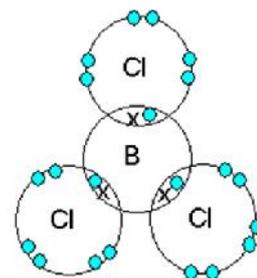
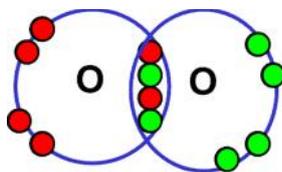
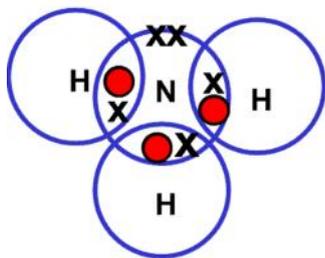
Ionic bonding is between ions and all their surrounding oppositely charged ions. Each sodium ion in this structure is surrounded and equally attracted by six chloride ions. The ionic bond is the attraction between all these ions

### Typical Physical properties of Ionic Compounds

- High melting points - There are strong electrostatic attractive forces between the oppositely charged ions in the lattice
- Non conductor of electricity when solid- The ions are held together tightly in the lattice and can not move so no charge is conducted
- Good conductor of electricity when in solution or molten – The ions are free to move when in solution and molten. Charge can be carried
- They are **usually** soluble in aqueous solvents.

## Covalent Bonding

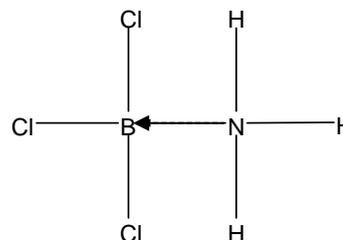
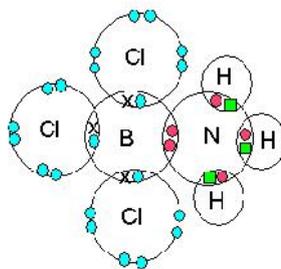
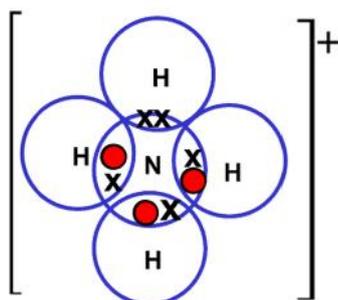
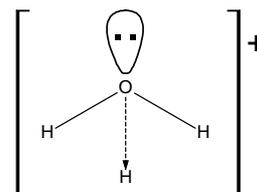
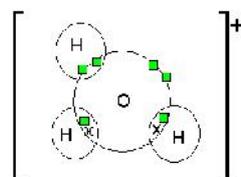
Definition: covalent bond is the strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms



## 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.  $\text{NH}_4^+$ ,  $\text{H}_3\text{O}^+$ ,  $\text{NH}_3\text{BF}_3$ )

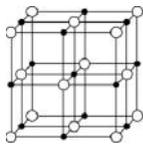


The dative covalent bond acts like an ordinary covalent bond when thinking about shape so in  $\text{NH}_4^+$  the shape is tetrahedral

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

The term *average bond enthalpy* is used as a measurement of covalent bond strength. The larger the value of the average bond enthalpy, the stronger the covalent bond

## 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 (Induced dipole–dipole, permanent dipole-dipole, hydrogen bonds) between molecules	<b>Iodine</b> <b>Ice</b> Carbon dioxide Water Methane

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

Property	Giant Ionic	Molecular (simple)
<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 Induced dipole–dipole/hydrogen bond)
<b>Solubility in water</b>	Generally good	generally poor
<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)
<b>conductivity when molten</b>	good: ions can move	poor: no ions
<b>general description</b>	crystalline solids	mostly gases and liquids

## 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>
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)**

## 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

### 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

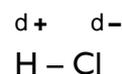
A compound containing elements of similar electronegativity and hence a **small electronegativity difference** will be purely **covalent**

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

A **polar covalent bond** forms when the elements in the bond have **different electronegativities**.

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



### 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**.

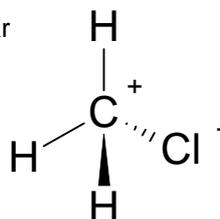
The individual dipoles on the bonds '**cancel out**' due to the symmetrical shape of the molecule.

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



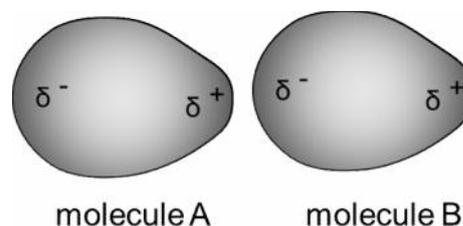
A compound containing elements of very different electronegativity and hence a very **large electronegativity difference** will be **ionic**

## 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.

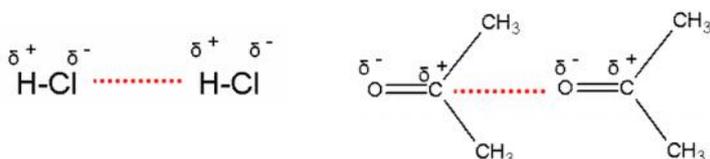
The increasing boiling points of the halogens down the group 7 series can be explained by the increasing number of electrons in the bigger molecules causing an increase in the size of the induced dipole–dipole interactions between the molecules. This is why  $I_2$  is a solid whereas  $Cl_2$  is a gas.

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 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.

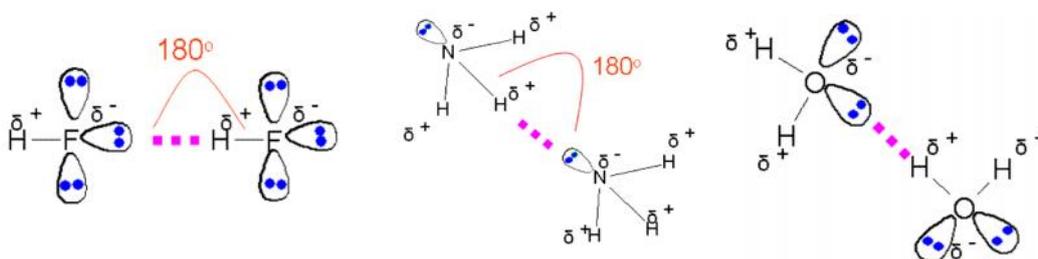


Permanent dipole forces occur in addition to induced dipole–dipole interactions

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  $\delta^-$   $\delta^+$  charges

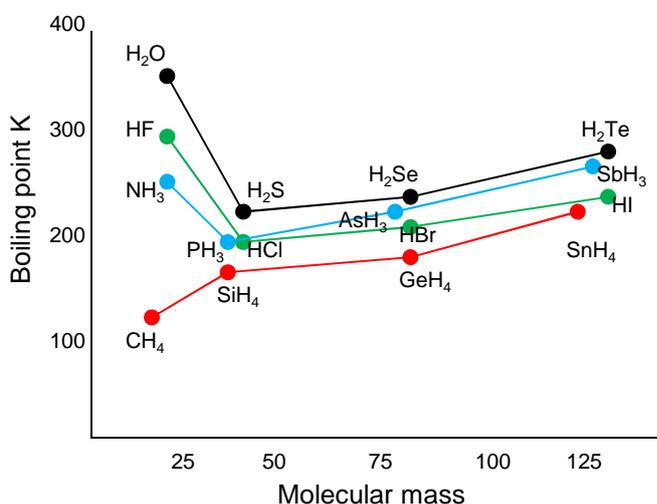
Hydrogen bonding occurs in addition to van der waals forces

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

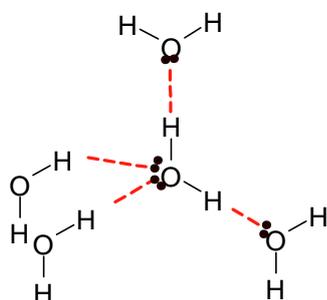
The **anomalously high** boiling points of  $\text{H}_2\text{O}$ ,  $\text{NH}_3$  and  $\text{HF}$  are caused by the hydrogen bonding between the molecules

The general increase in boiling point from  $\text{H}_2\text{S}$  to  $\text{H}_2\text{Te}$  is caused by increasing induced dipole-dipole interactions between molecules due to an increasing number of electrons.

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



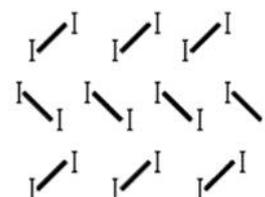
### Molecular: Ice



Water can form two hydrogen bonds per molecule, because oxygen is very electronegative, and it has two lone pairs of electrons.

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

### Molecular: Iodine



There are covalent bonds between the iodine atoms in the  $\text{I}_2$  molecule

The crystals contain a regular arrangement of  $\text{I}_2$  molecules held together by weak induced dipole-dipole interactions intermolecular forces