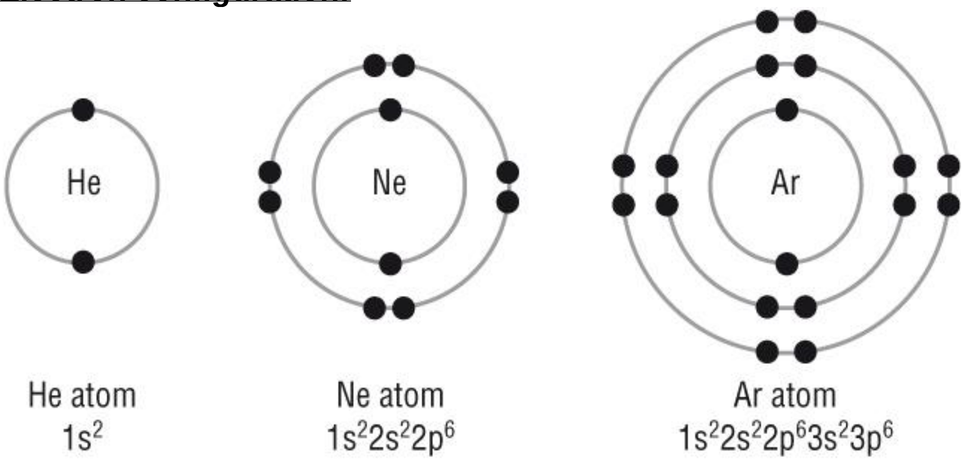
1.3 Bonding

Electron configuration:



- They have full outer shells and the electrons are paired with opposite spins fulfilling the '**octet rule**'.

Bonding:
- All other elements on the periodic table will combine by sharing or transferring electrons in order to fill their outer shell.
- The atoms, when combined **tend to** have the noble gas configuration - fulfilling the '**octet rule**', although there are some exceptions

Chemical bonding:
- The **Periodic Table** is made up of **2 types of elements**.

Metals + Non - metals
- This means there are **3 types of bonding** in which the atoms of the elements can combine:

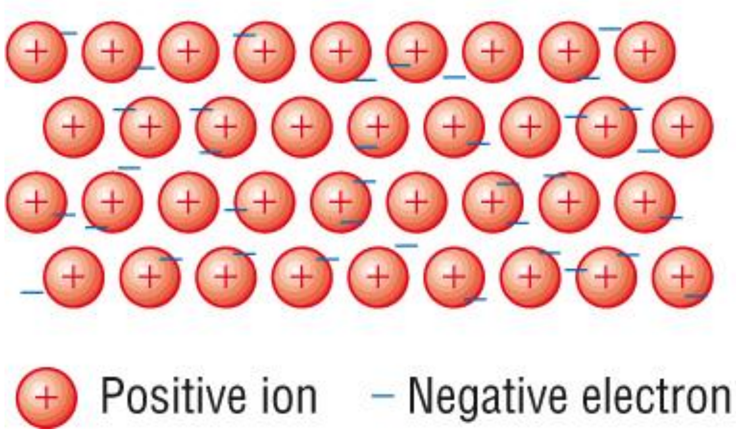
Type of Bonding	Which types of atoms combine	Electrons
Metallic	**Metal + Metal**	**Shared: Between all atoms**
Ionic	**Metal + Non-metal**	**Transferred: Metal → Non-metal**
Covalent	**Non-metal + Non-metal**	**Shared: Between the atoms**

Types of structure
- There are 2 types of structure:

Giant	Involves many atoms
Simple	Involves a few atoms

1) Metallic bonding

Giant Metallic Lattice Structure

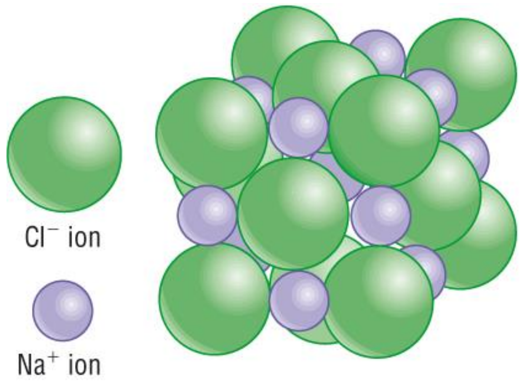


- Positive metal ions are in a fixed position while the outer shell electrons are **delocalised** between all the atoms in the metallic structure:
- **Attraction** occurs between the **positive metal ions** and the **negative delocalised electrons**.
- The metal ions pack together in a **lattice** arrangement.
- **Giant** as there are many metal ions in the arrangement.

2) Ionic Bonding

Giant ionic lattice structure

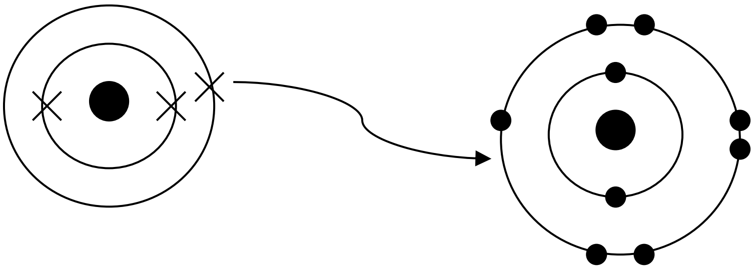
Ionic bonding results from the electrostatic force of attraction between oppositely charged ions.



- **Metals** form **positive** ions which we call **cations** when **electrons are lost**
- **Non - metals** form **negative** ions which we call **anions** when **electrons are gained**
- **Electrostatic forces of attraction** occur between ions holding them in place.
- **Giant** structure as **many ions** arranged in the structure

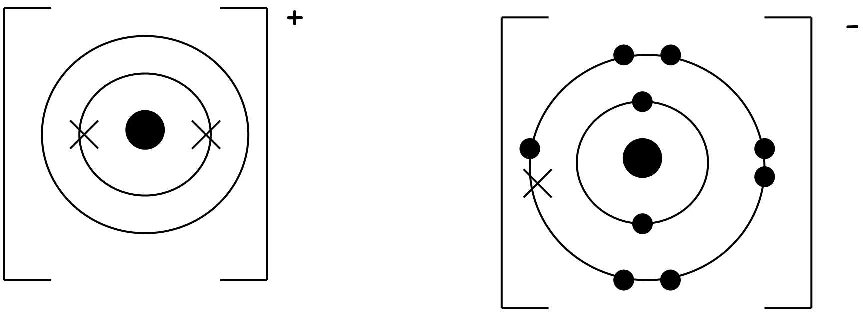
Lithium Fluoride:

- Dot and cross diagrams is a means of electron counting.



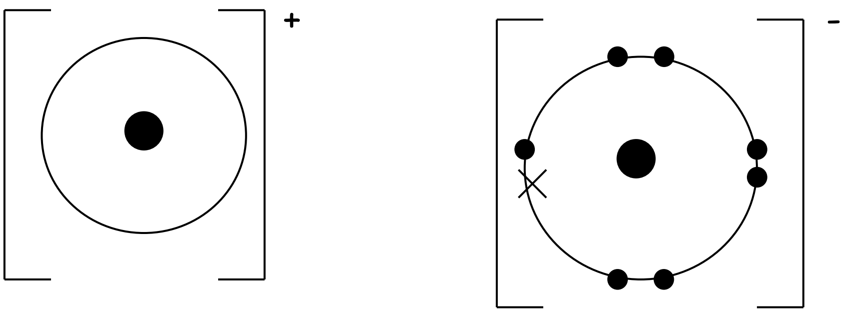
Li F

Becomes



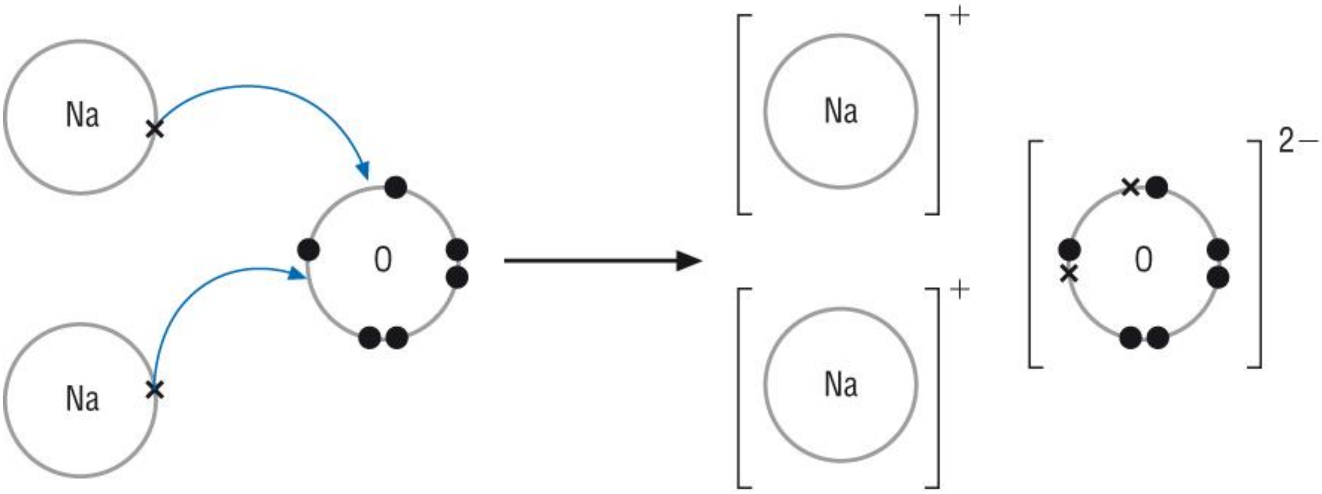
Li^+ F^-

However we only show the ions and only the outer shells:-



Li^+ F^-

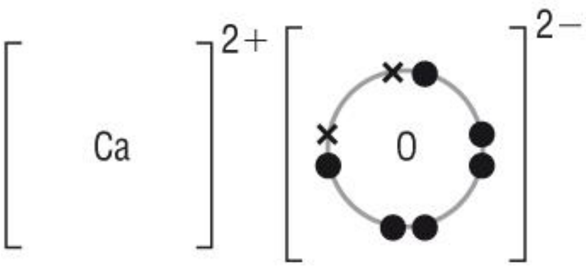
Sodium oxide:



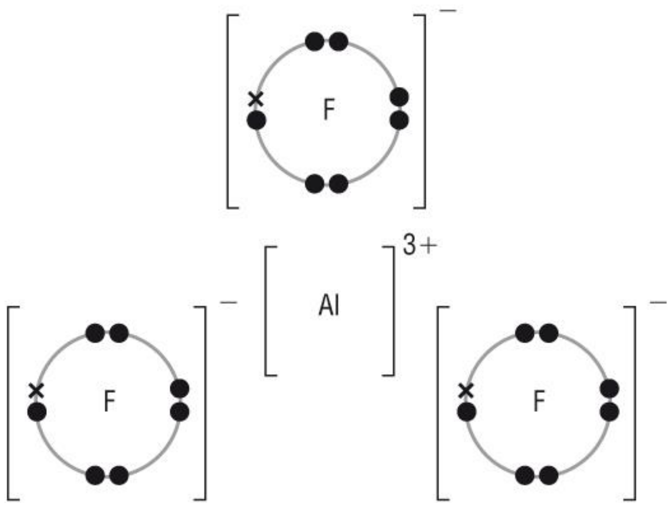
- Na atom loses 1e to empty its outer shell
- atom must gain 2e to fill its outer shell
- This means that 2 Na atoms must lose 1e each = 2e to give O its 2e's needed to fill its outer shell.

Further examples of ionic bonding:

CaO



AlF_3



Ions and the Periodic Table

Predicting charges:

- It is possible to predict the charges of elements using the Periodic Table:

Group	1	2	3	4	5	6	7	0
No e's in outer shell	1	2	3	4	5	6	7	Full
No electrons lost gained	Lose 1	Lose 2	Lose 3	x	Gain 3	Gain 2	Gain 1	x
Charge on ion	1+	2+	3+	x	3-	2-	1-	x

- Be, B, C and Si do not usually form ions as too much energy is required.

Transitions metals:

- These elements can usually form more than one ion:

Element							
Charge / Ox No		2+	3+	4+	5+	6+	7+
Roman numeral		II	III	IV	V	VI	VII

- The charge / oxidation number of the ion is written as a roman numeral:

Iron (II) Fe^{2+} Copper (II) Cu^{2+}

Compound ions (common ions)

- Groups of covalently bonded atoms can also gain and lose electrons.
- These are called molecular (or common ions)

1+		1-		2-		3-	
Ammonium	NH_4^+	Hydroxide	OH^-	Carbonate	CO_3^{2-}	phosphate	PO_4^{3-}
		Nitrate	NO_3^-	Sulphate	SO_4^{2-}		
		Nitrite	NO_2^-	Sulphite	SO_3^{2-}		
		Hydrogen carbonate	HCO_3^-	Dichromate	$Cr_2O_7^{2-}$		

Predicting ionic formula:

- The ions in a chemical formula must **add up to zero**.
- Use subscripts after an ion in a formula to double/triple that ion so the sum=0. eg. $CuCl_2$

For **F**ormulae, the numbers are on the **F**loor

- If you are double/tripling ions that consist of more than one element (molecular ions) brackets must be used. eg. Ca**(**OH**)**$_2$
- If Roman numeral numbers follow a transition metal ion in brackets, that tells you the positive charge on that transition metal ion.

Examples

Sodium Chloride -

Na^{+}	**Cl^{-}**	Write the ions with charges
1+	**1-**	Multiply up the charges to = 0
NaCl		Bring together / charges cancel

Copper (II) Chloride

Cu^{2+}	**Cl^{-}**	Write the ions with charges
2+	**1- x2**	Multiply up the charges to = 0
$CuCl_2$		Bring together / charges cancel

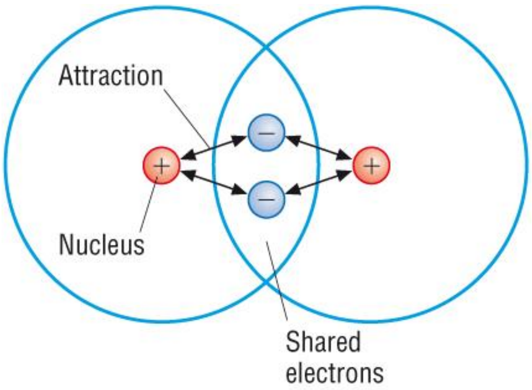
Calcium Hydroxide

Ca^{2+}	**(OH^{-})**	Write the ions with charges
2+	**1- x2**	Multiply up the charges to = 0
$Ca(OH)_2$		Bring together / charges cancel

In this example, brackets are required as we multiply up more than one element

3) Covalent Bonding

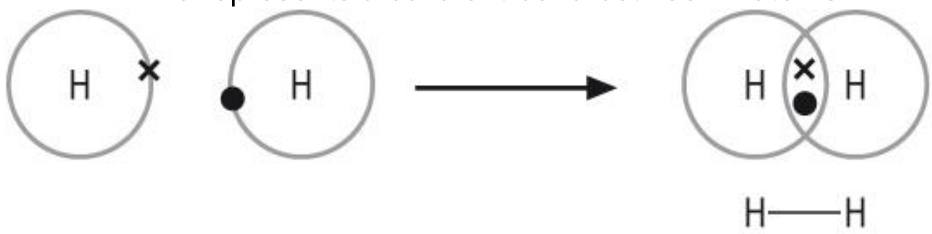
Covalent bond contains a shared pair of electrons



- These occur between 2 or more non metals.
- As neither will donate, they have to share.
- Attraction comes between the positive nucleus and the negative electrons that are shared between them.
- In each case the outer shell has to be filled (reach the noble gas configuration).

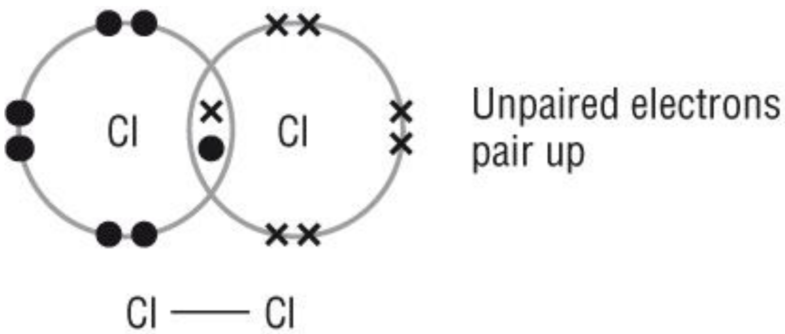
Full shell rule
Single covalent bonds

- The single outer shell electrons are involved in a covalent bond.
- Each hydrogen has 1 single unpaired electron in its outer shell.
- Each hydrogen contributes 1 electron to the covalent bond = 2 electron.
- A line represents a covalent bond between 2 atoms.

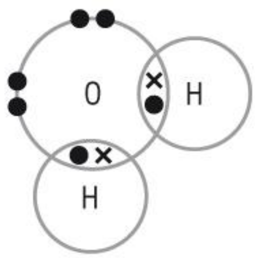


Chlorine:

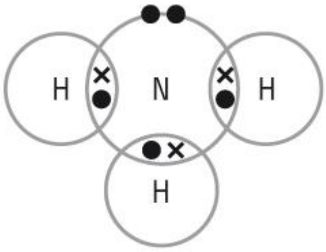
- Each chlorine has 7e in its outer shell.
- Each Cl has 1 unpaired electron in its outer shell.
- Each Cl contributes an electron each = 2e forming 1 single covalent bond:



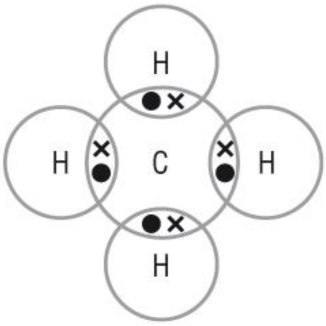
Other examples:



H_2O
2 unpaired electrons
pair up



NH_3
3 unpaired electrons
pair up



CH_4
4 unpaired electrons
pair up

Line diagrams (fill them in below)

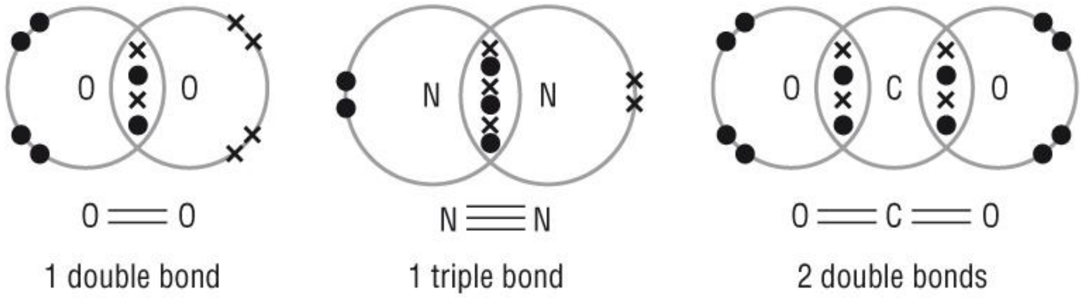
Group	4	5	6	7
Element	C	N	O	F
Electrons in outer shell	4	5	6	7
Lone pairs	0	1	2	3
Unpaired electrons in outer shell	**4**	**3**	**2**	**1**
No of covalent bonds	**4**	**3**	**2**	**1**

- **Lone pairs of electrons** are the paired electrons and are not used in a covalent bonding.
- They do however give a region of concentrated negative charge (later)

Multiple covalent bonds

Multiple covalent bonds contain multiple shared pair of electrons

- Some atoms can form more than a single bond, double and triple.
- This depends on how many single unpaired electrons there are in the outer shell:



- Each bond is represented by a line:

Dative covalent bonds (co – ordinate bonds)

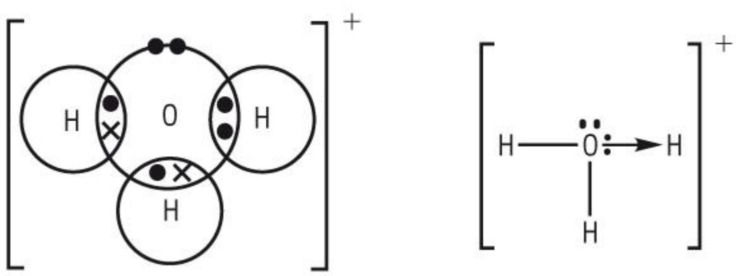
Dative covalent bond is where one atom provides both of the shared pair of electrons

- A covalent bond is when each atom provides 1e each to be shared between them.
- A dative covalent bond is when both of the bonding electrons are provided by the same atom.
- In order for this to happen:
  - An atom or atom in a molecule **must** have a **lone pair of electrons to donate**
  - An atom or atom in a molecule **must** have capacity to **accept a pair of electrons**

The hydroxonium ion:

- A lone pair of electrons from the water molecule is used to forma **dative covalent bond.**
- The bond is indistinguishable from the other covalent bonds but we use an **arrow** to show that one of the bonds is dative.

The arrow point in the direction of donation

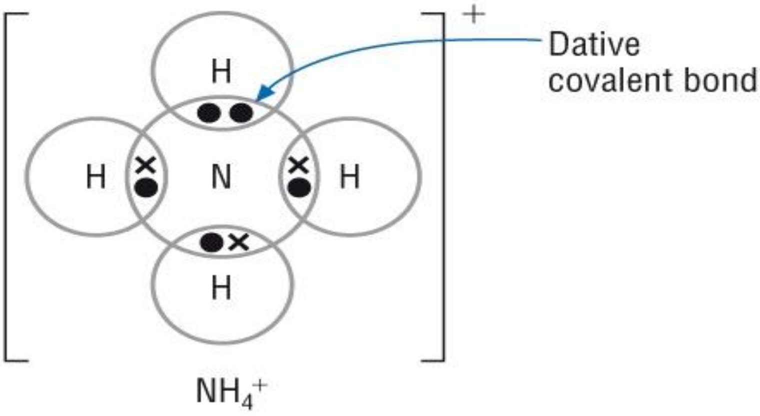


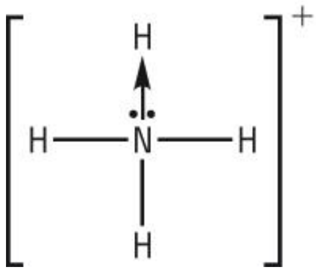
- The **hydroxonium ion** is responsible for acid reactions and is represented as:

H_3O^+ or $H^+_{(aq)}$

The ammonium ion:

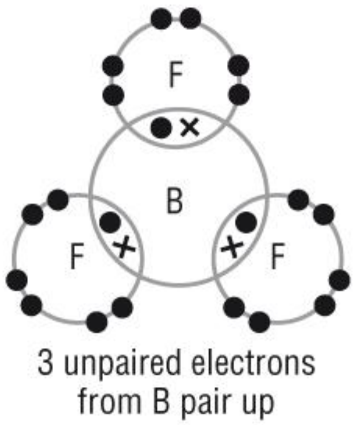
- The same can be done for ammonia. This forms the **ammonium ion:**





Breaking the full shell rule:

➢ **BF_3 Less than a full outer shell:**

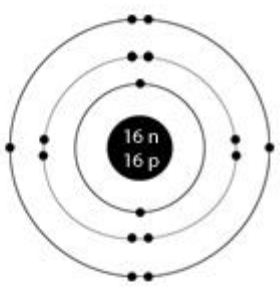
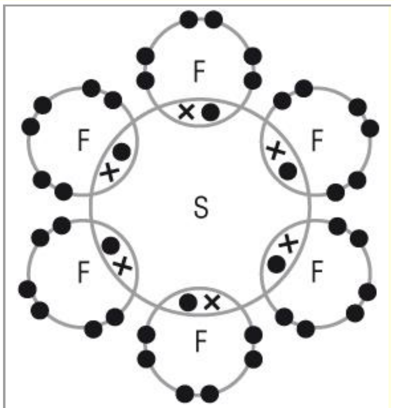


- Some atoms do not have enough electrons in their outer shell.
- Boron is in group 3 of the Periodic table. This means it has 3e in its outer shell.
- This means that Boron can only form 3 covalent bonds = 6e in its outer shell

➢ **SF_6 Expanding the full outer shell:**

From Period 3, the electron shell is much larger and can accommodate more covalent bonds.

- Some atoms are able to have more than 8 electrons in its outer shell.
- It depends on how many of their outer shell electrons / lone pairs of electrons are used in bonding

Sulphur		
		• If both of the lone pairs unpair then there are 6 unpaired electrons. • These will form **6 covalent bonds** = **12e**

Shapes of molecules and ions:

Valence Shell Electron Pair Repulsion Theory (VSEPR):

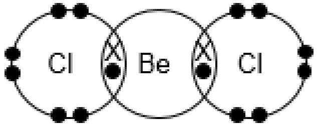
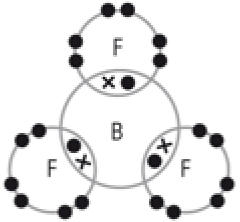
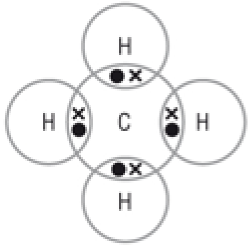
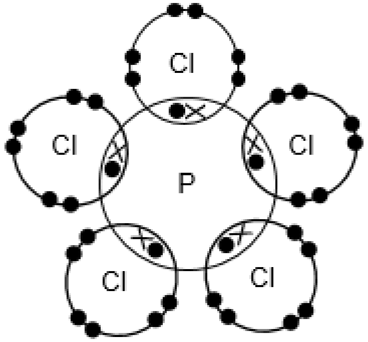
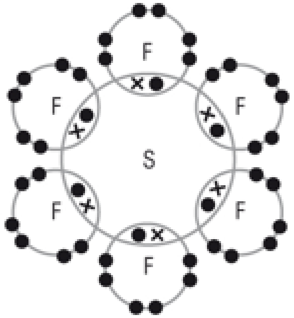
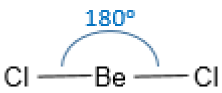
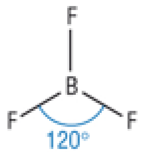
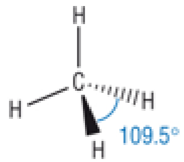
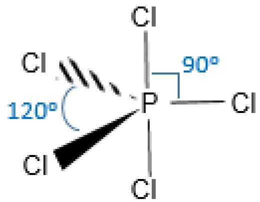
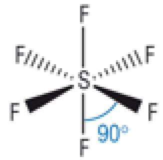
- The shape of a molecule is wholly determined by the number of electron pairs (or charge clouds) around a central atom. These electron pairs can be:

  - **Bonding pairs**
  - **Lone pairs**

Charge Clouds

- A charge cloud is an area where you will find an electron.
- As electrons move, they move around in this charge cloud.
- As electrons are negative, this gives a region of negativity.
- Electron pairs repel each other as far as possible around the central atom.
- This repulsion determines the shape of the molecule and bond angle.

Bonding pairs of electrons

Molecule	$BeCl_2$	BF_3	CH_4	PCl_5	SF_6
Dot and Cross Diagram					
Number of electron pairs	**2**	**3**	**4**	**5**	**6**
Shape and bond angle					
Name of shape	**Linear**	**Trigonal planar**	**Tetrahedral**	**Trigonal bipyramidal**	**Octahedral**

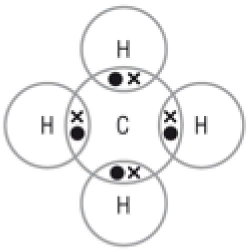
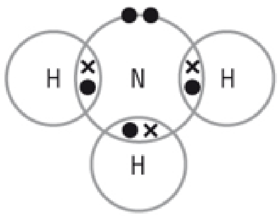
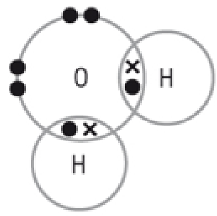
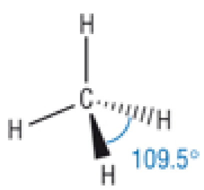
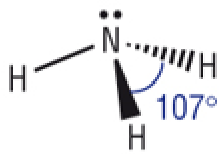
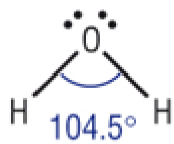
Lone pair electrons:

- Lone pairs are closer to the central atom than bonding pair electrons.
- Lone pair electrons will repel other pairs of electrons more than bonding pairs of electrons.
- Each lone pair of electrons reduces the bond angle by 2.5_o.
- Count how many bonding pairs (or areas with double bonds) and lone pair electrons
- Pairs of electrons repel other pairs of electrons as far as possible and this determines the shape
- Order of repulsion:

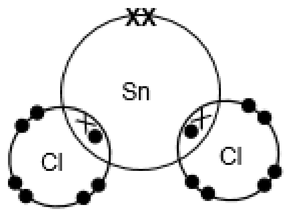
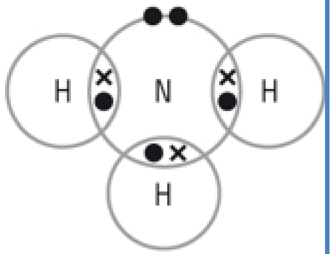
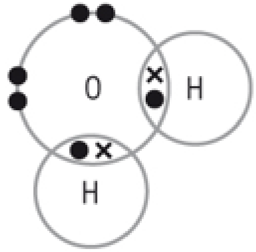
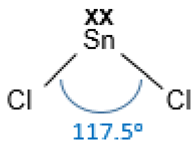
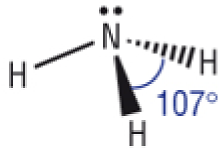
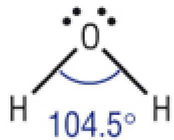
Lone Pair – Lone Pair > Lone Pair – Bonding Pair > Bonding Pair – Bonding Pair

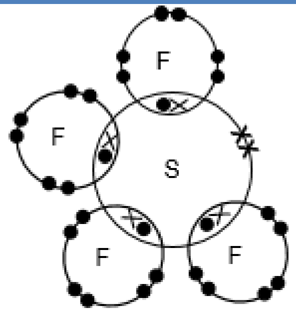
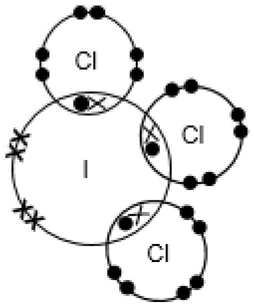
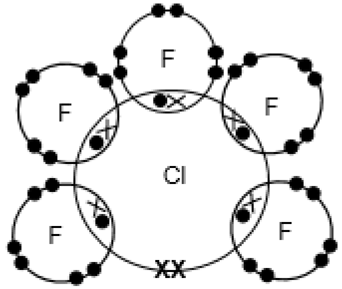
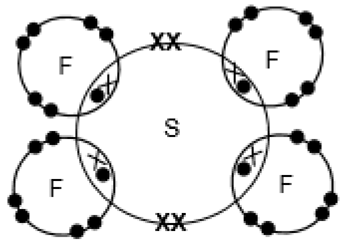
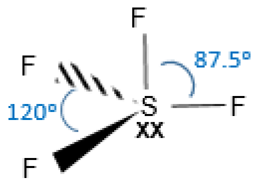
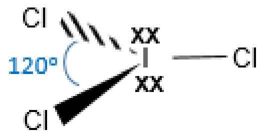
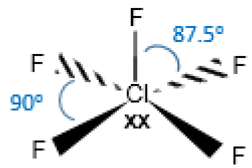
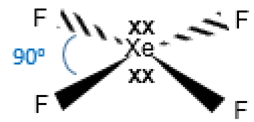
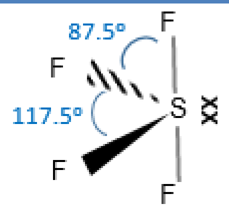
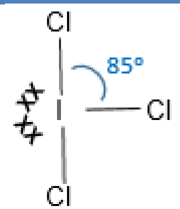
Example:

- All 3 of the molecules below are based on 1 parent shape.
- The parent shape is **Tetrahedral** as this is the standard shape with all 4 pairs of electrons around the central atom being used as **bonding pairs**.
- As you go down the table (across the Period) one of the bonding pairs is now a lone pair.
- The Bond angle reduces by 2.5_o for each lone pair electrons.
- The shape also changes due to one less atom bonded to the central atom.

Dot and Cross diagram			
Shape and Bond angle			
Numbers of pairs of electrons	**4**	**4**	**4**
Number of Bonding pairs	**4**	**3**	**2**
Number of Lone pairs	**0**	**1**	**2**

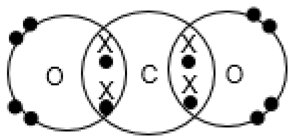
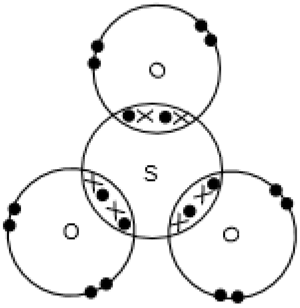
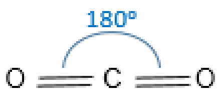
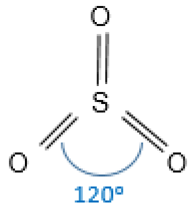
Lone pairs of electrons

Molecule		$SnCl_2$	NH_3	H_2O
Dot and Cross Diagram				
Number of electron pairs	Total	3	4	4
	Bonding pairs	2	3	2
	Lone pairs	1	1	2
Shape and bond angle				
Name of shape		**Non - Linear**	**Trigonal Pyramidal**	**Non - Linear**
		Parent shape – Trigonal planar	*Parent shape – Tetrahedral*	*Parent shape – Tetrahedral*

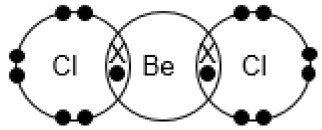
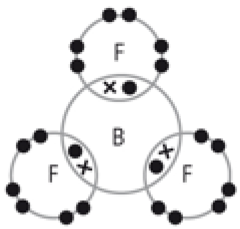
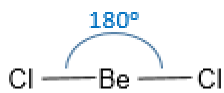
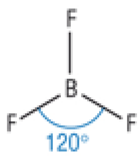
Molecule		SF_4	ICl_3	ClF_5	XeF_4
Dot and Cross Diagram					
Number of electron pairs	Total	5	5	6	6
	Bond pairs	4	3	5	4
	Lone pairs	1	2	1	2
Shape and bond angle					
Name of shape		**Trigonal Pyramidal**	**Trigonal planar**	**Square Pyramidal**	**Square planar**
		Parent shape – Trigonal Bipyramidal	*Parent shape – Trigonal Bipyramidal*	*Parent shape – Octahedral*	*Parent shape – Octahedral*
Alternative Shape and bond angle					
Name of shape	(alternative)	**Seesaw**	**T - Shape**		

Molecules with double bonds:

- Double bonds have 4e in one region.
- To work out the shape we have to think in terms of **bonding regions** instead of bonding pairs:

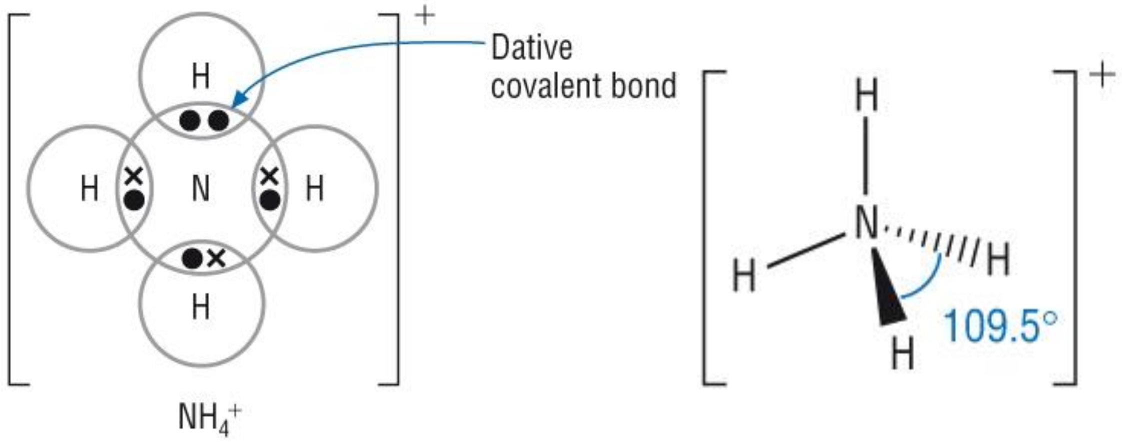
Molecule	**CO_2**	**SO_3**
Dot and Cross Diagram		
Number of electron pairs	**2 regions (4 pairs in 2 regions)**	**3 regions (6 pairs in 3 regions)**
Shape and bond angle		
Name of shape	**Linear**	**Trigonal planar**

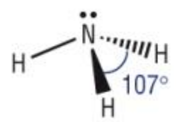
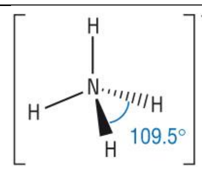
- CO_2 has **2 bonding regions and no lone pair regions.** The **2 bonding regions** will repel each other as far as possible.
- This will give a **linear** molecule with a bond angle of **180**$_o$.

Molecule	**$BeCl_2$**	**BF_3**
Dot and Cross Diagram		
Number of electron pairs	**2**	**3**
Shape and bond angle		
Name of shape	**Linear**	**Trigonal planar**

Shapes of ions:

- Molecular ions follow the same rules as any molecule, remember the ammonium ion:



3BP / 1LP		4BP
	**The addition of a datively bonded H^+ ion uses the lone pair converting it to a bonding pair (similar for water)**	
Trigonal pyramidal		**Tetrahedral**

- 4 bonding pairs therefore it is basically the same shape as CH_4 - tetrahedral.

Summary:

State the number of bonding and lone pairs of electrons

Pairs of electrons repel as far as possible

This determines the shape

Lone pairs repel more than bonding pairs as closer to central atom

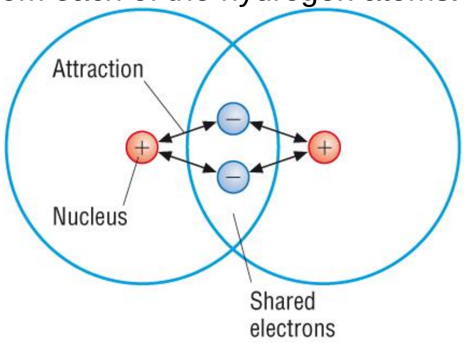
Each lone pair reducing the bond angle by 2.5$_{o}$ as it is closer to the central atom

A - Polar bonds

Polar and non-polar bonds:

1) Non-Polar bonds:

- A covalent bond shares an electron pair. In a hydrogen molecule, the electrons are attracted by the nucleus from each of the hydrogen atoms.



- As the 2 atoms are identical, they will have the same number of protons and shielding.
- This means that the electrons are being 'pulled' equally by both of the hydrogen atoms.
- We say that the H - H bond is **non-polar**
- **The same is true for any diatomic molecule where the atoms are identical**

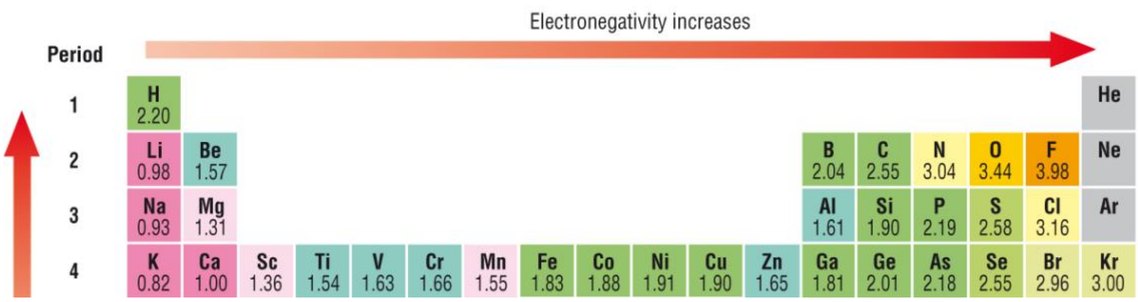
2) Polar bonds and Electronegativity

- A covalent bond is between 2 different atoms means attraction will be unequal.
- One atom will have more protons in the nucleus / less shielding.
- This means that that atom will attract the bonding pair of electrons more than the other.
- This power of attraction is called **Electronegativity**

Electronegativity: **The ability of an atom to attract the bonding pair of electrons in a covalent bond.**

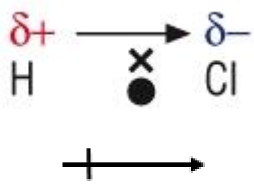
How is electronegativity measured?

- Linus Pauling came up with the Pauling scale in 1932.
- Basically as you go towards the top right hand side of the Periodic table, the elements become more electronegative:
  - ➢ Across a Period the number of protons increase therefore attraction increases.
  - ➢ Up a Group the number of shells / shielding decreases therefore attraction increases



- **Cl, N, O and F** are the **most electronegative** elements.
- Reactive metals, **Na - K** are the **least electronegative** elements
- Generally, for a bond to be polar, the differences have to be greater than 0.4

Hydrogen chloride:



- The attraction for the bonding electrons will be different.
- Chlorine is **more electronegative**.
- This means the bonding electrons will be closer to the chlorine atom.
- **Electron distribution is unsymmetrical**
- This covalent bond is **Polar**.
- The molecule has a **permanent dipole** δ**-** δ**+**

Di - ' 2 ' **Pole - poles (positive and negative)**

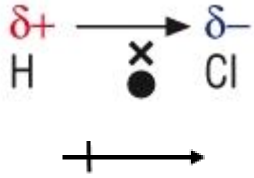
- The chlorine atom has the bonding electrons nearer to it, therefore it has a **small negative charge,** δ**-**.

δ is used to mean ' a little bit of '

- The hydrogen doesn't have its fair share of the bonding electrons it will have a **small positive charge,** δ**+**.
- The greater the difference in electronegativities, the greater the permanent dipole and the bigger the δ**-.**

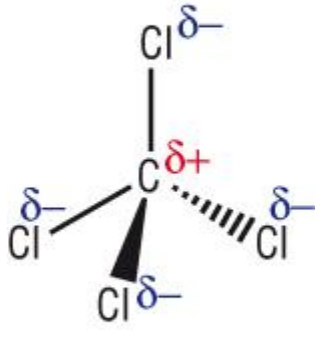
<u>B - Polar molecules:</u>

- Molecules like HCl are called polar molecules.
- This is because **over the whole molecule** there are ' **2 poles '**.



This represents the direction of the dipole

- Some molecules can have polar bonds without being polar.
- It comes down to the shape of the molecule and symmetry:

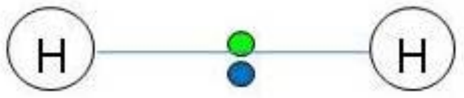


- Basically there is not a positive end and a negative end.
- So CCl_4 in classed as non - polar even though all of the bonds are polarised.

Electronegativity and bonding type

1) Covalent

- Elements of very similar electronegativities have their bonding electrons shared equally between the 2 atoms.



- This is clearly a covalent bond.

2) Polar covalent

- Elements with a slight difference in electronegativity will still share their bonding electrons.
- The electrons are not evenly shared:

- The covalent bond will however be polar.

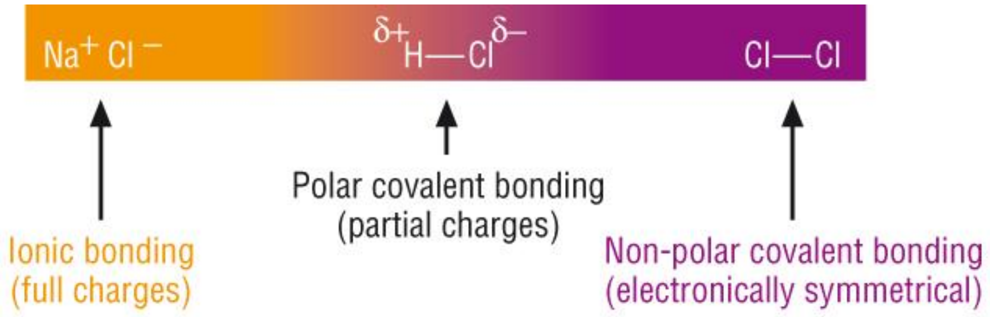
3) Ionic

- Elements with very different electronegativities however have a different type of bonding
- The more electronegative element will attract the bonding electrons to itself so much that that element takes both of the bonding electrons.

- The more electronegative element has now gained an extra electron to become a 1- ion.
- The lesser electronegative element has now lost an electron to become a 1+ ion.

Covalent to ionic:

- As a bond becomes **more polar**, there is a movement from **covalent to ionic bonding**.



Intermolecular forces

Strengths of bonds and forces:

Ionic - When ionic compounds melt/boil, the forces of attraction are overcome and the ions separate.

$$NaCl_{(s)} \longrightarrow Na^{+}_{(g)} + Cl^{-}_{(g)}$$

Molecular - When molecular compounds melt/boil, the covalent bonds remain intact.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)}$$

- The molecules do not break up therefore there must be forces of attraction between the molecules.
- These are called **intermolecular forces**.
- Since molecular substances can exist as solids, liquids and gases intermolecular forces must always exist although they may be very weak.
- There are 3 types of intermolecular forces of attraction:

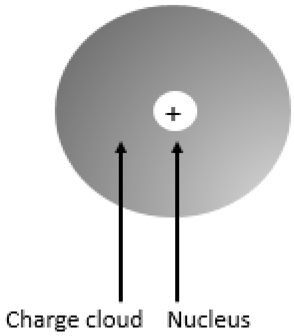
1) Van der Waals' forces

2) Permanent dipole - dipole forces

3) Hydrogen bonding

1) Van der Waals' forces (induced dipole - dipole interactions)

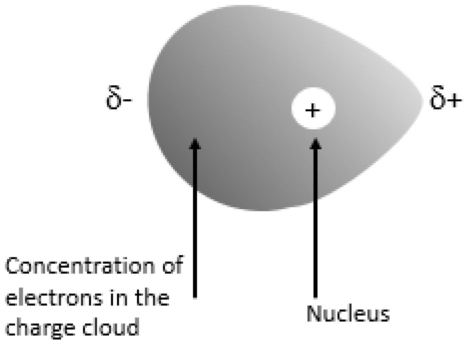
- Helium does not form ionic or covalent bonds but it is possible to condense it to a liquid.
- This very weak force of attraction is known as **Van der Waals** forces.
- It is due to the **continually moving electrons in charge clouds**.
- At any instant the electrons may be on one side more than another (an uneven distribution of electrons)
- This sets up an **instantaneous dipole**.
- This will **induce a dipole** in a neighbor.
- The **2 dipole are now attracted to each other**.
- This continues throughout the **neighbor's neighbor and so on**.
- As the **electrons continually move, the dipoles are continually changing**
- **Overall the atoms to experience attraction.**



No Dipole

- The (+)ve nucleus is in the middle.
- Electrons in charge cloud is evenly distributed

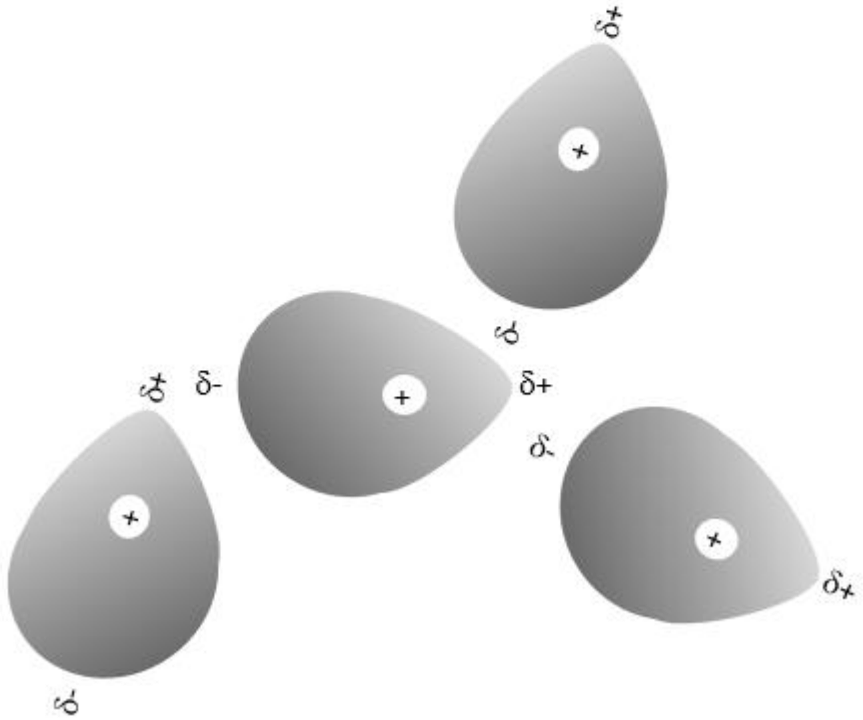
Unlikely



Instantaneous dipole

- The (+)ve nucleus is **not** in the middle.
- Electrons in charge cloud is **unevenly** distributed

Most likely



Induced dipole

- The instantaneous dipole **induces a dipole** in neighbors.
- **δ+ and δ- attracted** to each other.
- **Instantaneous dipole changes – repeats process**

Overall - attraction

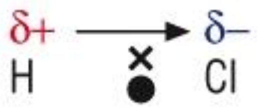
Hydrocarbons and Diatomic elements

- The greater the number of electrons, the greater the charge cloud, the greater the uneven distribution of electrons in the charge cloud.
- This will give stronger dipole and therefore greater VDW forces of attraction.
- If you consider the alkanes – The boiling point increases as the Mr increases (number of electrons).
- As you go down a Group, the Boiling point increases due to the increasing Ar (number of electrons).
- This is because of the increased number of electrons increases the VDW attraction.

NOTE:
All atoms / molecules have VDW due to the fact that all atoms / molecules have electrons

<u>2) Permanent dipole - dipole interactions:</u>

- **Only occurs in polar molecules**
- When we have 2 atoms in a covalent bond with different electronegativities, the bond is polarised.
- If the molecule is unsymmetrical (ie has a **δ+** end and a **δ-** end), the molecule is said to have a **permanent dipole**:



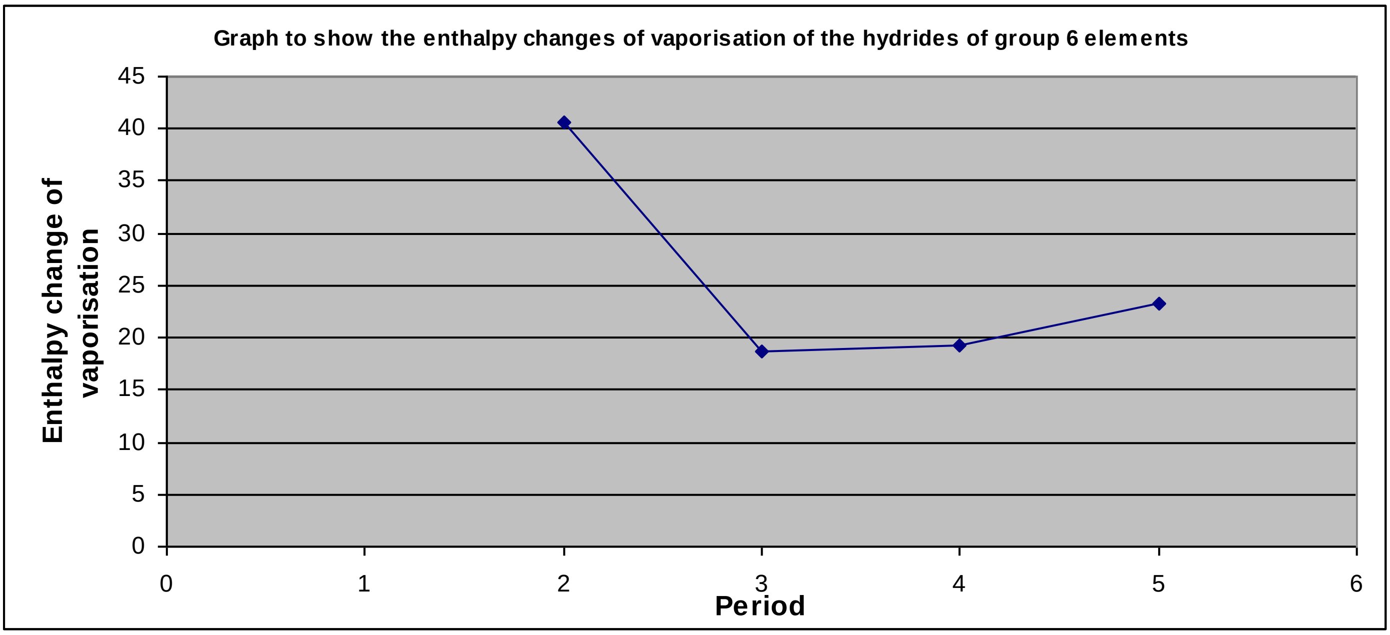
- The **δ+** end of one molecule will be attracted to the **δ-** end of a neighboring molecule.
- This attraction is called a **permanent dipole - dipole force** of attraction.
- They are stronger than VDW so these molecules tend to have higher boiling points.

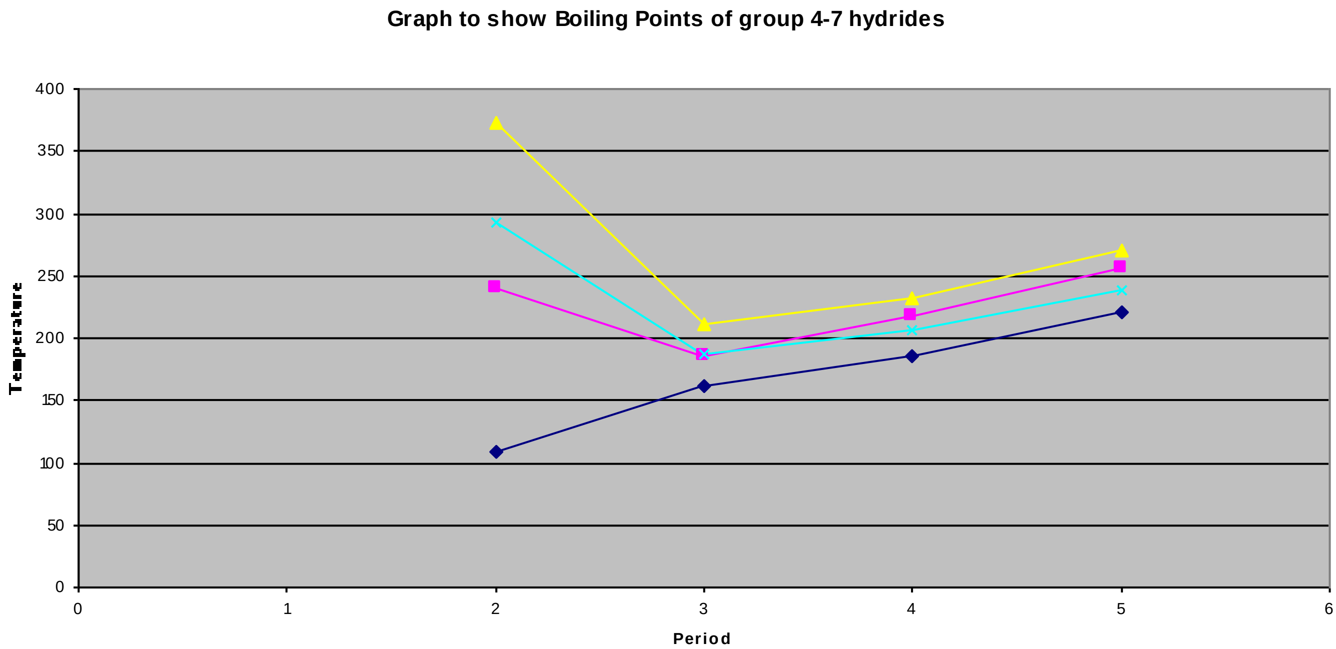
- Typically, the liquids of these are deflected charged rods.

NOTE:
These forces of attraction are stronger than VDW and occur in addition to VDW forces of attraction.

3) Hydrogen Bonding:

Water is peculiar



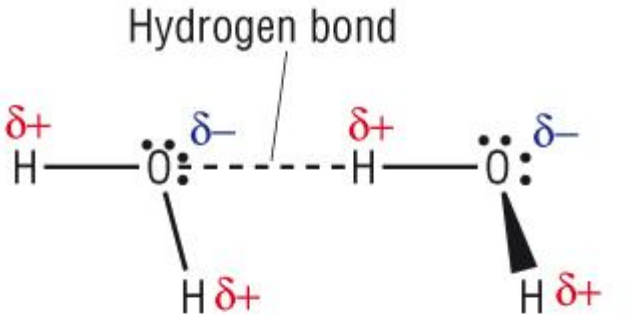


- Hydrogen bonding explains these observations.
- It is the strongest of the intermolecular forces.

NOTE:
These forces of attraction are stronger than Permanent dipole – dipole and VDW forces of attraction. They occur in addition to VDW and permanent dipole - dipole forces of attraction.

Why are hydrogen bonds so strong?

- The atoms **O,N,F** are so strongly **electronegative** that the bonding pair of electrons are so far from the **H** that they are almost able to be donated.
- This along with the small size of the H atom means that the H in the molecule is very positive.
- The pair of bonding electrons are very near the **O,N,F**. With their small sizes they are very negative.
- With its own **lone pair**(s) of electrons a strong force of attraction is able to occur between the H and the lone pair of electrons on neighbouring molecules.
- This force of attraction is known as a **Hydrogen bond** and is represented by a dotted line:



A hydrogen bond is formed by attraction between δ+ and δ– charges on different water molecules.

- Maximum bond strength is when the bond angle O-H-O is 180_o.
- The strength of a hydrogen bond is typically ~30Kj.
- Compare this with the strength of a covalent bond ~300Kj. A Hydrogen bond is ~$1/10_{th}$ a covalent bond.
- Similarly with VDW attraction ~3Kj.

Summary

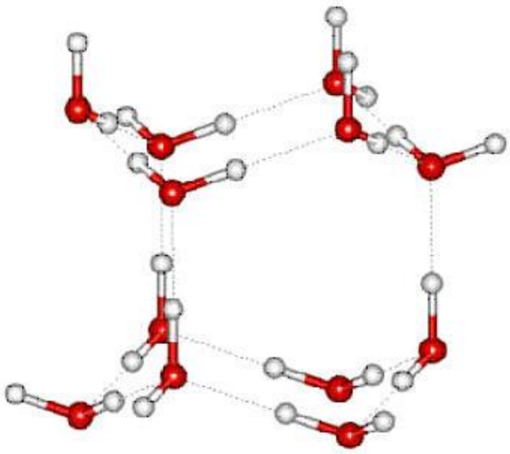
1) Within a molecule, the hydrogen is highly polarised (very positive)
2) Within a molecule, the atom joined to the hydrogen is very electronegative, O,N,F.
3) Within a molecule, the atom joined to the hydrogen must also have a lone pair of electrons.
4) The covalent – Hydrogen – Covalent bond MUST be 180_o

- H
- O,N,F
- O,N,F must have a lone pair
- C – HB – CB must be 180_o

- Hydrogen bonding is strong enough to **change physical properties** but **not chemical properties**.
- Water would be a **gas** at room temperature and pressure **if it was unable to hydrogen bond**.
- Other molecules with hydrogen bonding **tend to be soluble** in water as they can form Hydrogen bonds with water.

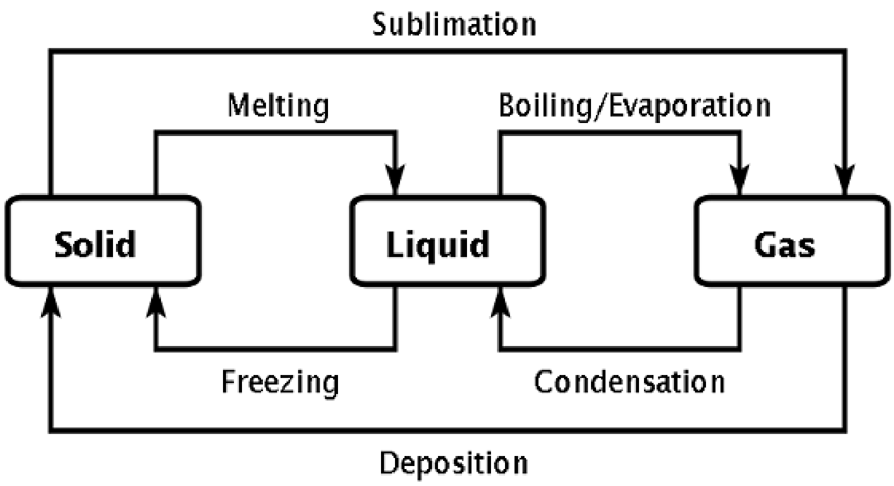
Ice is less dense than water:

- The 2 hydrogen bonds per water molecule of water sets up a 3D structure.
- The Hydrogen bonds are longer than the O - H bonds.
- This means they are held further apart than in water making ice less dense.



- Hydrogen bonds give water a skin effect and this contributes to the high surface tension.
- Other liquids with hydrogen bonding also have a surface tension.

Physical properties – Rules:



1) **Melting and Boiling points**
    - Due to the strength of attraction between particles.
    - Order of strengths:

Strongest:	**Giant Molecular (covalent)**	Highest Mpt
	Ionic	
	Metallic	
	Simple Molecular:	
	- Hydrogen bonding (covalent)	
	- Permanent Dipole – dipole (covalent)	
Weakest:	**- Van der Waals (covalent)**	Lowest Mpt

    - Energy must go in to overcome bonds / forces of attraction – Melting / Boiling
    - Energy is given out when bonds / forces of attraction are made – Condensing / freezing

2) **Conductivity**
    - **There must be charged particles that are free to move**

3) **Solubility**
    - The interactions between the solute and solvent **must** be about the same:

      Polar will dissolve in **polar** (water)
      Non – polar will dissolve in **non – polar**
      Ionic will dissolve in **polar (ie partially charged)**

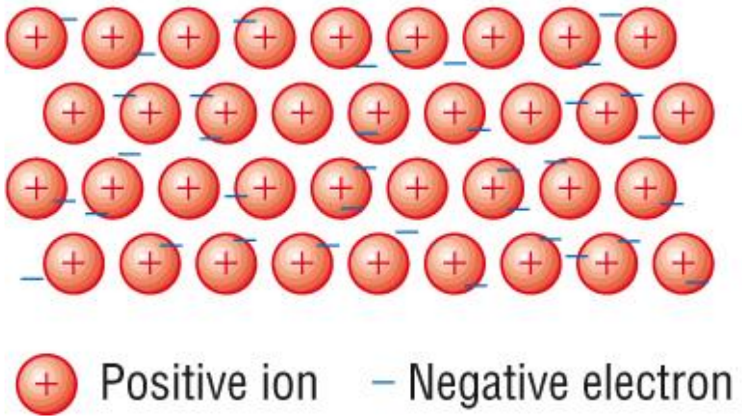
Now we know the rules, we can look at the physical properties of:

1) **Metals**
2) **Ionic compounds**
3) **Covalent compounds**
    a. **Simple molecular**
    b. **Giant molecular (macromolecules)**

1) Metallic bonding, structure and physical properties

Giant metallic lattice structure

- Positive metal ions are in a fixed position while the outer shell electrons are **delocalised** between all the atoms in the metallic structure:



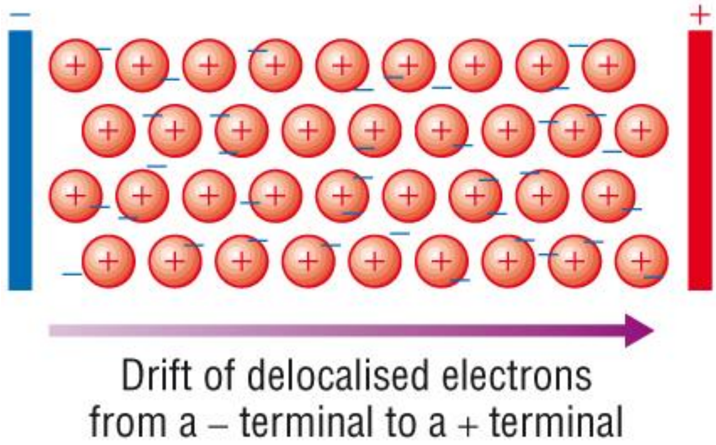
- The model consists of **metal ions** surrounded by '**Mobile sea of electrons**'.

High Melting points

- **Strong electrostatic forces of attraction** occurs between the ions and the delocalised electrons.

High electrical conductivity

- The sea of electrons are charges that are free to move.



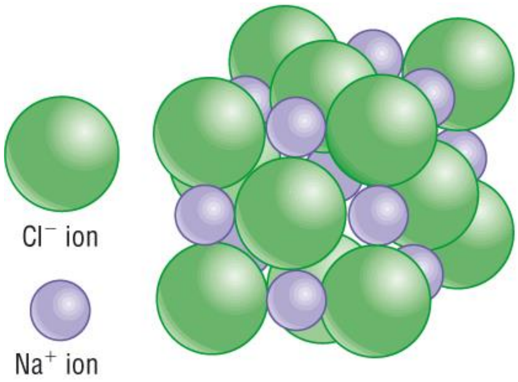
- The sea of electrons also pass on kinetic energy making them **good thermal conductors**

Malleability and ductile -

- Malleable - can be hammered or pressed into shape.
- Ductile - can be drawn / stretched into wire.
- Due to the delocalised electrons, the metallic structure has a degree of 'give' which allow layers to slide past each other.

2) Ionic bonding, structure and physical properties

Giant ionic lattices



- Each sodium ion is surrounded by 6 chloride ions.
- Each chloride ion is surrounded by 6 sodium ions.
- This continues in all directions and is describes as a **Giant Ionic Lattiice**

High melting points

- There are **strong electrostatic forces of attraction between the ions**.
- The higher the charges between the ions, the stronger the electrostatic forces of attraction.
- The stronger the forces of attraction, the more heat energy is required to overcome those forces and hence melt / boil.

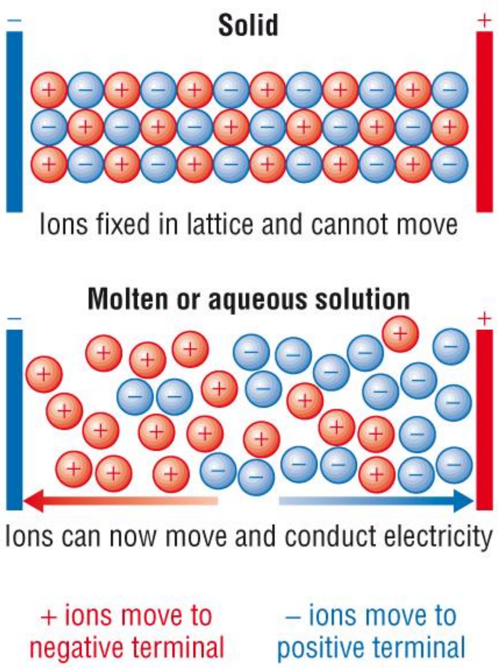
High electrical conductivity:

Solid – does not conduct

- In a **solid giant ionic lattice**, the ions are held in a **fixed position**. This means that the **ions cannot move**.
- This is why they **do not conduct electricity as a solid**.

Molten or dissolved – does conduct

- When the ions are **molten or dissolved - the ions are now free to move.**
- This means they **do conduct electricity:**

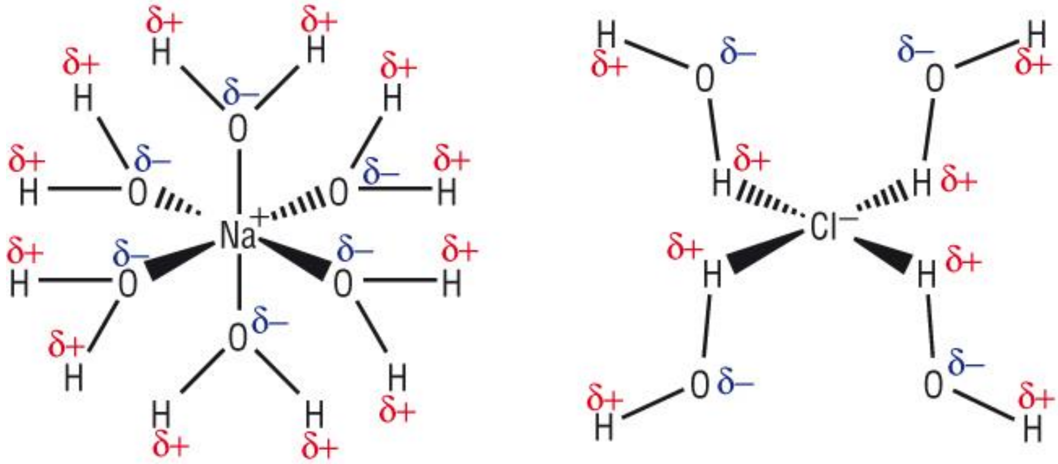


Solubility:

- When an ionic solid dissolves, the ions in the lattice separate.
- The energy required to separate ions within a lattice is large. ie high melting points.
- The energy required on dissolving must be equal and opposite to the energy required to separate ions (as dissolving separates the ions).

Where does the energy come from?

- Where does this large amount of energy come from if all we are doing is dissolving the solid in water?
- There must be some process during dissolving that can release enough energy to separate the ions from the lattice.
- If energy is being released there must be some type of attractive interaction to release energy:-



- The force of attraction comes between the δ+ / δ- end of the **polar** water molecule and the opposite charges on the ion.
- This attraction releases energy (as all 'bond forming' reaction do).
- Many water molecules surround the anion as shown.
- This releases energy which is used to break up the lattice structure:

3) Covalent bonding, structure and physical properties

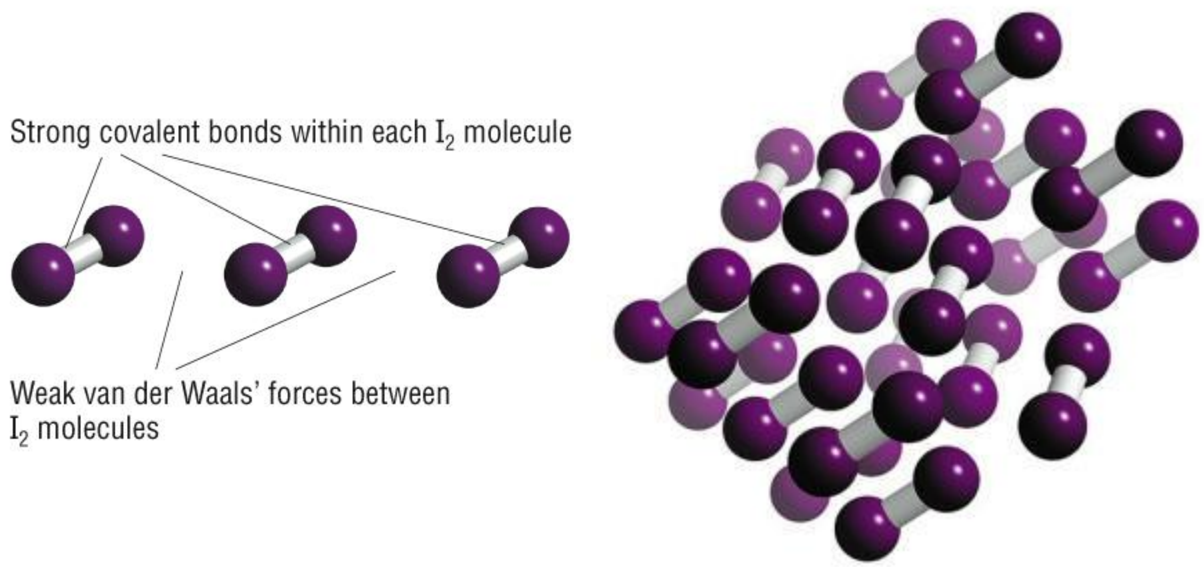
- Covalent compounds fall into 2 categories:

a) Simple molecular structures:

- These are made up from simple (small) molecules such as: CO_2, N_2, O_2, I_2 and H_2O.
- The molecules are held together by **weak intermolecular forces**:

Strongest:	**Hydrogen bonding**	**Highest Mpt**
	Permanent Dipole – dipole	
Weakest:	**Van der Waals**	**Lowest Mpt**

- The atoms within the molecules are made up from strong covalent bonds.



Low melting points:

- These have **weak IMF of attraction between** the molecules.
- When you melt or boil these molecules, you do not break up the molecule, only overcome the forces of attraction between them.

Does not conduct electricity:

- Simple molecules have no free moving charges.
- This means they do not conduct electricity.

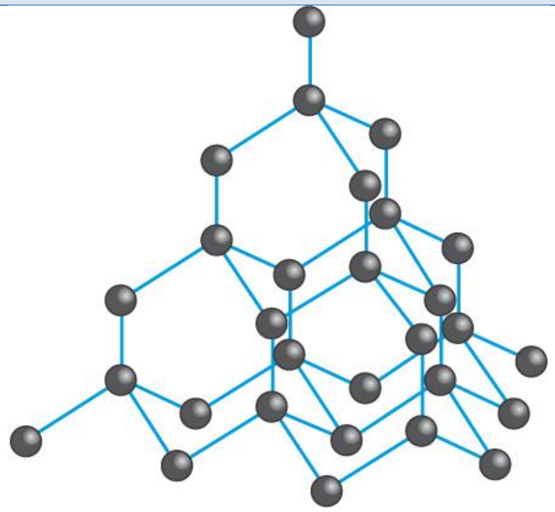
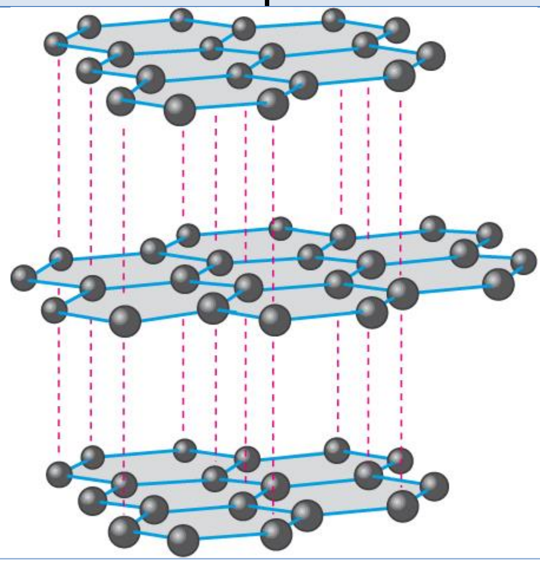
Solubility:

- The interactions between the solute and solvent **must** be about the same:

Polar will dissolve in **polar** (water)
Non – polar will dissolve in **non – polar**

b) Giant molecular structures (macromolecules)

- These are structures that have extensive covalent bonded atoms in a giant lattice structure.
- Diamond and graphite are examples of these:

Diamond	Graphite
	
• Each carbon has 4 covalent bonds arranged in an extended tetrahedral structure.	• Each carbon has 3 covalent bonds arranged in a hexagonal layer. • The 4^{th} **electron is delocalised** between the hexagonal layers. • **Weak VDW forces** of attraction hold the layers in place.
High Melting point • Extensive **strong covalent bonds** throughout the whole structure.	**High Melting point** • Extensive strong covalent bonds within the layers of the structure. • However the layers are able to slide over each other easily. • Used in pencils and dry lubricant.
Does not conduct electricity • **No free moving electrons or charges**: **non - conductors of electricity.**	**Does conduct electricity** • The 4th **electron is delocalised and free moving** between the layers. • This means it is **able to conduct** electricity.
Insoluble	**Insoluble**
Hardness • Extremely hard	**Hardness** • Very hard although the layers can slide over each other.
Good thermal conductor • As vibrations can carry through the extensive structure	**Low density** • As the layers are far apart compared with the covalent bond length – sports equipment
Sublimes	