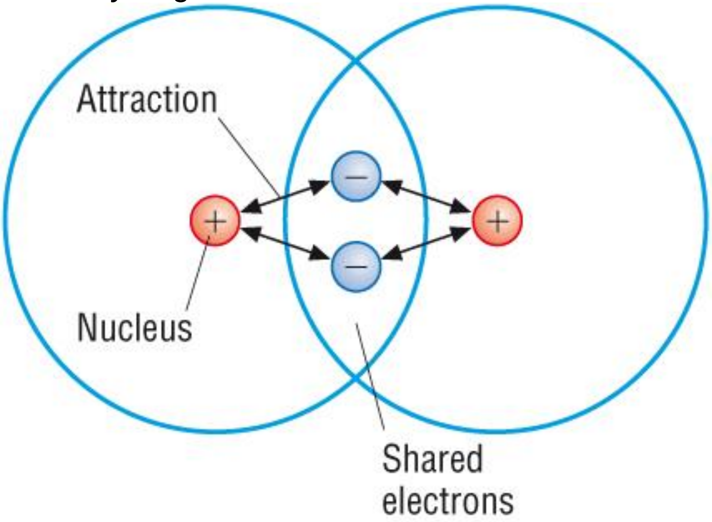
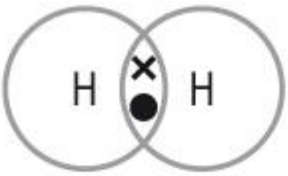
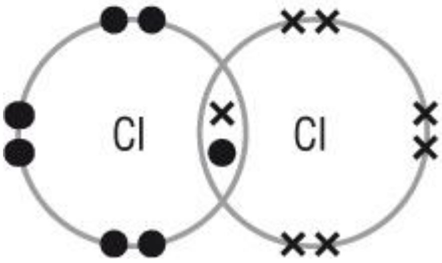
2C – Intermolecular forces, structure and properties:

Electronegativity and polarity
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.





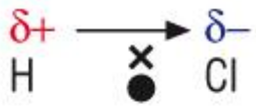


- As the 2 atoms are identical, they will have the same number of protons.
- This means that the electrons are being 'pulled' equally by both of the hydrogen atoms.
- We say that the H - H bond is **non-polar**
- **The same is true for any diatomic molecule where the atoms are identical:**

2) Polar bonds:

- If a covalent bond is between 2 different atoms then the attraction from each is more likely to 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.
- **The power of an atom to attract bonding pair electrons to itself is called electronegativity/**

Hydrogen chloride:



The bonded electron pair is attracted towards the Cl atom.

- The attraction from the nucleus for the bonding electrons will be different.
- Chlorine is the more electronegative atom.
- This means the bonding electrons will not be closer to the chlorine atom.
- This covalent bond is **Polarised**.
- Because chlorine atom has the bonding electrons nearer to it, the chlorine atom will have a **small negative charge, δ-**.

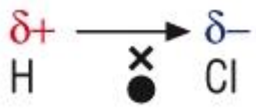
δ is used to mean ' a little bit of '

- Because the hydrogen doesn't have its fair share of the bonding electrons it will have a **small positive charge, δ+**.
- The molecule has a small negative charge at one end and a small positive charge at the other end.
- We say the molecule has a **permanent dipole**:

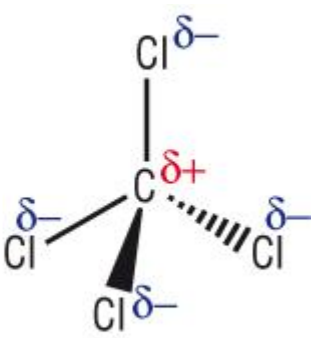
Di - ' 2 '
Pole - poles (positive and negative)

Polar molecules:

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



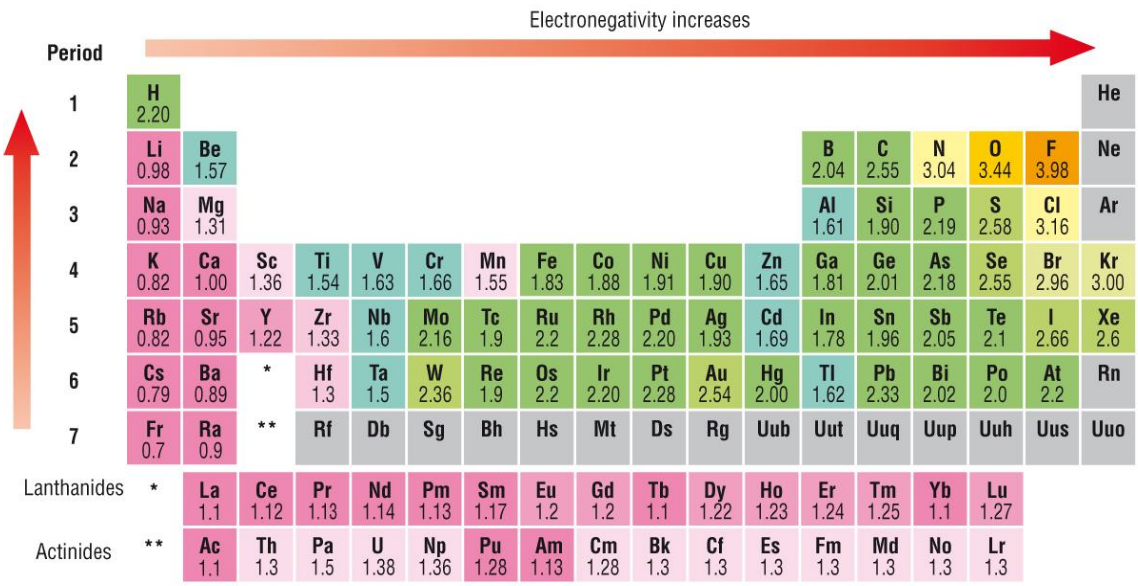
- Some molecules can have polar covalent bonds without being polar.
- It comes down to symmetry:



- 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.

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:



Pauling definition of electronegativity:-

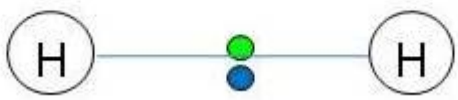
The electronegativity of an atom represents the power of an atom in a molecule to attract electrons to itself.

- **Cl, N, O and F** are the **most electronegative** elements.
- Reactive metals, **Na - K** are the **least electronegative** elements
- The greater the difference in electronegativities, the greater the permanent dipole and the bigger the δ**-**.

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.

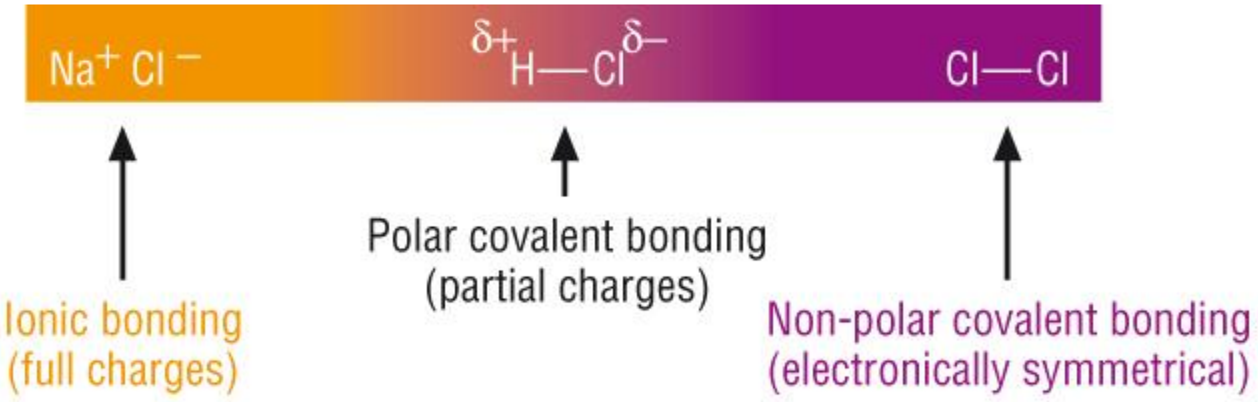
3) Ionic

- 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**.



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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.

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

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

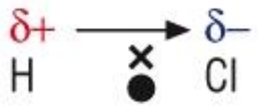
1) Van der Waals' forces

2) Permanent dipole - dipole forces

3) Hydrogen bonding

Permanent dipole - dipole interactions:

- When we have 2 atoms in a covalent bond with different electronegativities, the bond is polarised.
- If the molecule 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:

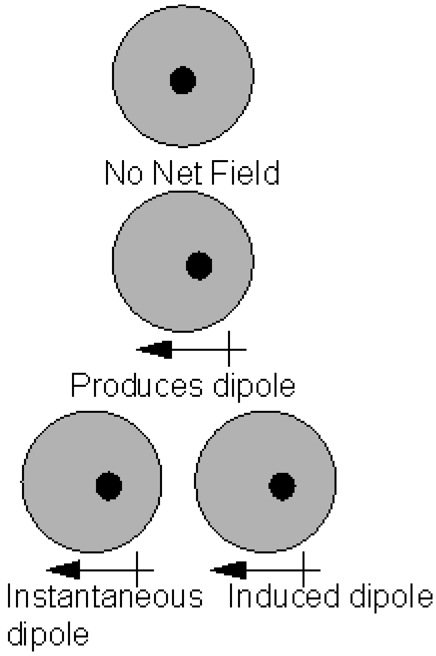
Dipole–dipole interaction between a δ+ atom of one molecule and a δ– atom of another molecule.

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 then to a solid.
- Energy is released when a change of state occurs. 0.105$KjMol^{-1}$
- This very weak force of attraction is known as **Van der waals** forces.
- It is due to the continually changing electric charge interactions between atoms, called **induced dipole - dipole forces of attraction**.

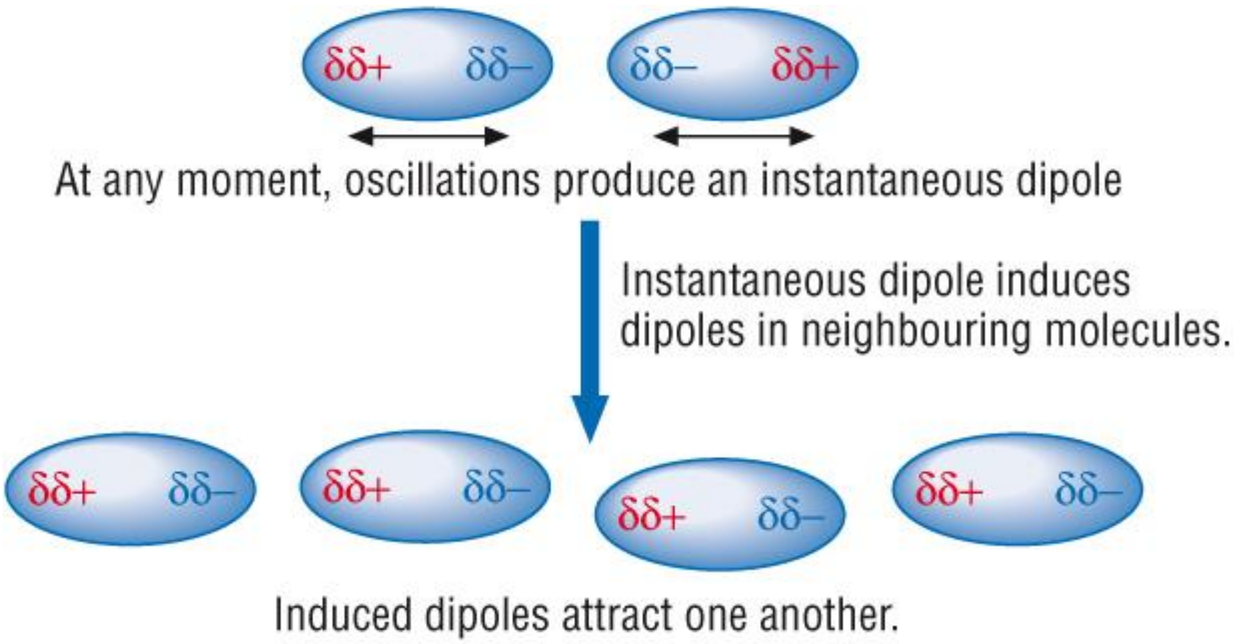
What causes Van der Waals' forces:

- These are present in all molecules but are the only forces of attraction present in non - polar molecules
- The electrons in shells are **continually moving**.
- In the turmoil we get an **uneven distribution of electrons / charge**.
- At any moment or snap shot in time there would be an **instantaneous dipole** across the whole atom / molecule.
- The negative end of the dipole **induces a dipole of opposite charge** in neighbouring atoms
- A force of **attraction results**.
- These induced dipole – dipole interactions produces a **cohesive force**.



- Imagine an atom to be like a large spherical jelly with a golf ball at the center.
- The golf ball is the nucleus, the jelly is the cloud of electrons whizzing about this.
- The net average field will be zero because the (+)ve nucleus field will be exactly balanced by the electron cloud.
- Atoms vibrate, at any instant the cloud is likely to be slightly off center. This creates an **instantaneous dipole**.
- If we have another atom next to it, this atom will be affected by the **instantaneous dipole**.
- This will **induce a dipole** in the neighbouring atom.
- The 2 dipoles attract one another – producing an attractive interaction.

- The forces of attraction are so weak that we use δδ to represent extra small charges:



- Molecules with very similar electronegativities will also only have Van der Waals' forces of attraction, eg:

Hydrocarbons Diatomic elements

- The greater the number of electrons (and protons) in an atom / molecule, the greater these fluctuations are and the greater the fluctuations
- This will give greater VDW attraction.
- If you consider the alkanes – The boiling point increases as the Mr increases.
- This is because of the increased number of electrons, which increase the VDW attraction.

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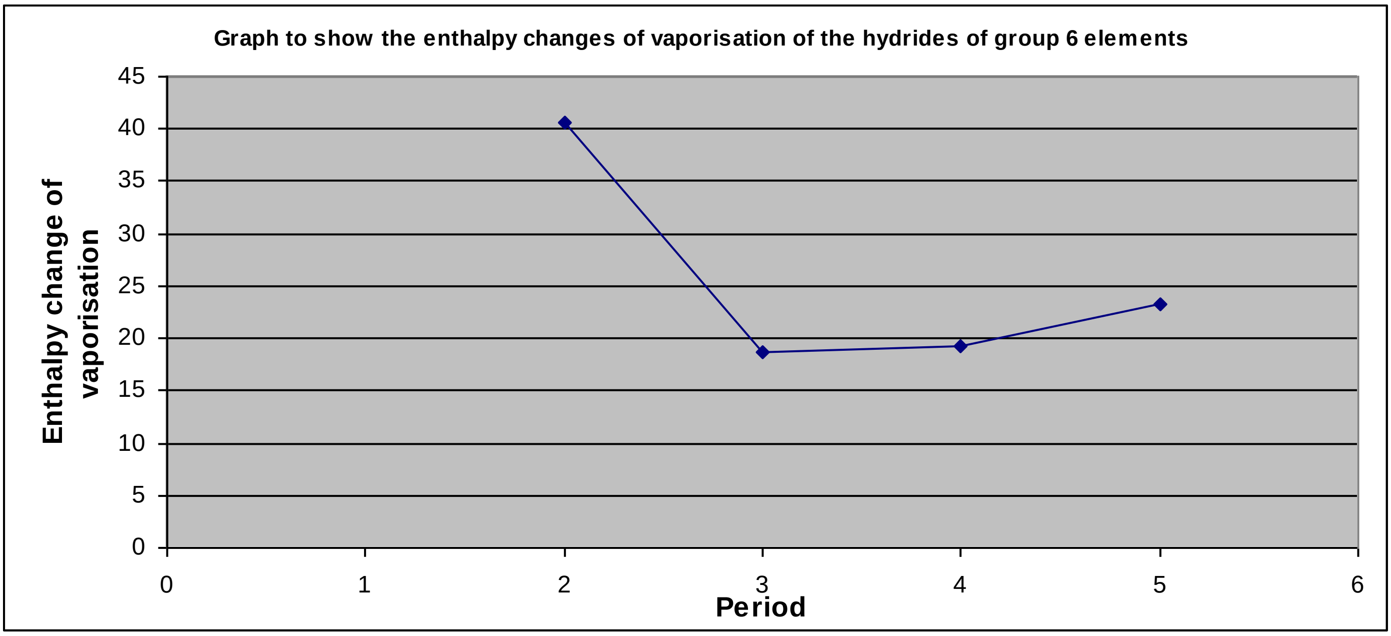
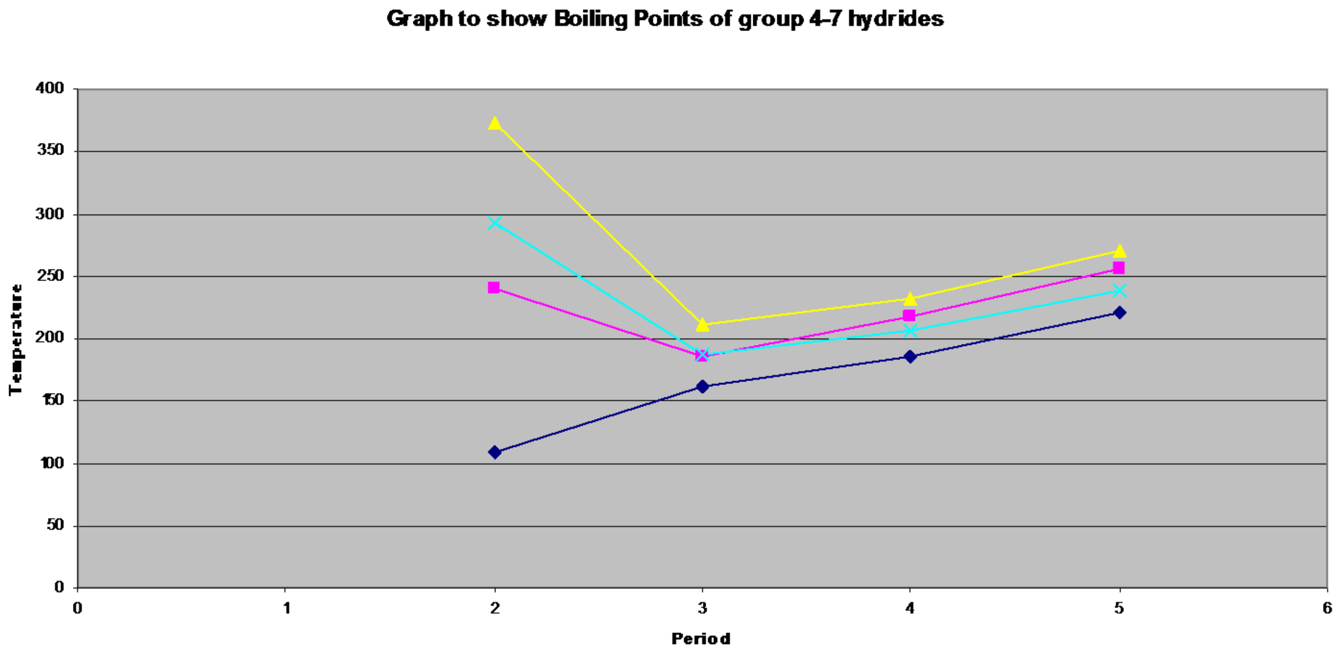
<u>Hydrogen Bonding:</u>

Water is peculiar

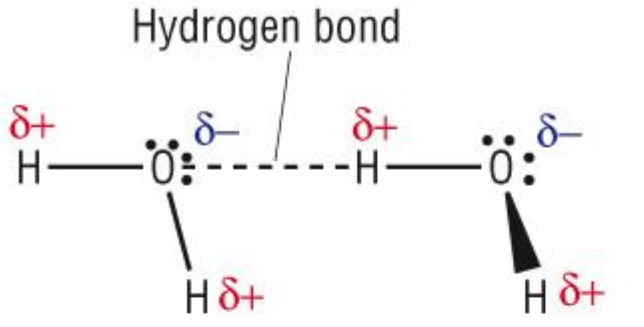




- Hydrogen bonding explains these observations.
- It is the strongest of the intermolecular forces.

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 ~1/10th a covalent bond.
- Similarly with VDW attraction ~3Kj.

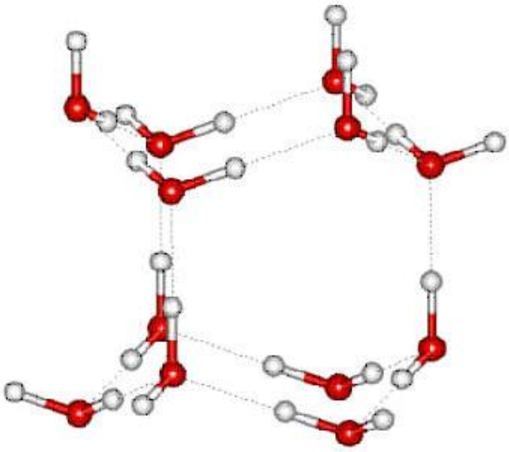
Summary

1) **Within a molecule, the hydrogen must be highly polarised (very positive)**
2) **Within a molecule, the atom joined to the hydrogen must be very electronegative, O,N,F.**
3) **Within a molecule, the atom joined to the hydrogen must also have a lone pair of electrons.**
   - **H**
   - **O,N,F**
   - **O,N,F must have a lone pair**

- 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**.

Ice is less dense than water:

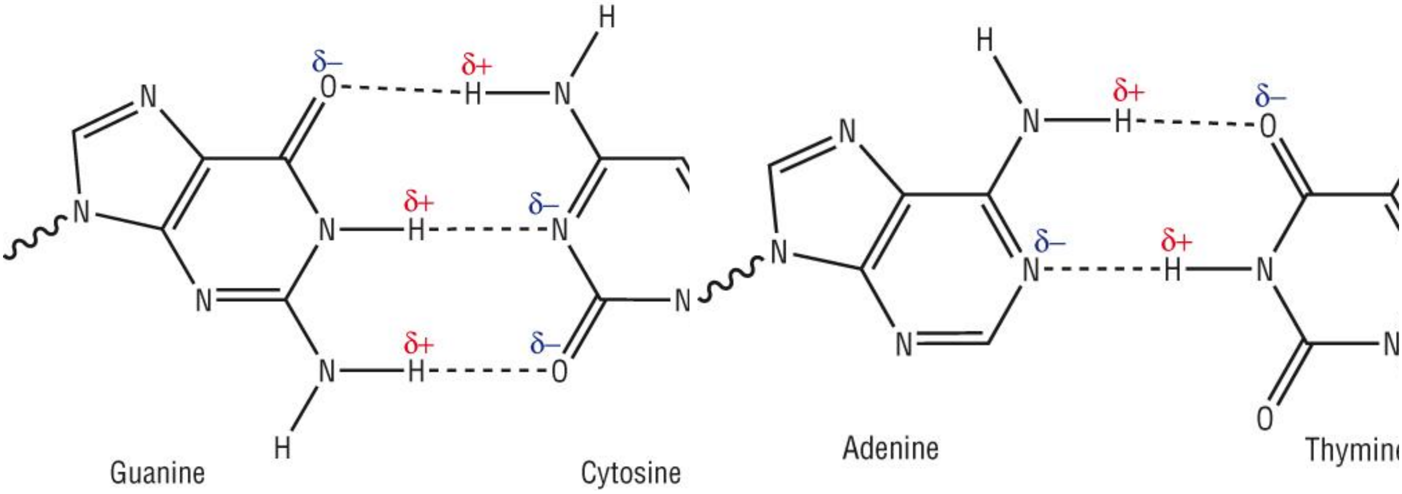
- The 2 hydrogen bonds per water molecule of water sets up a 3D structure.
- The Hydrogen bonds are longer than the O - H bonds.
- This means they are held further apart than in water making ice less dense.

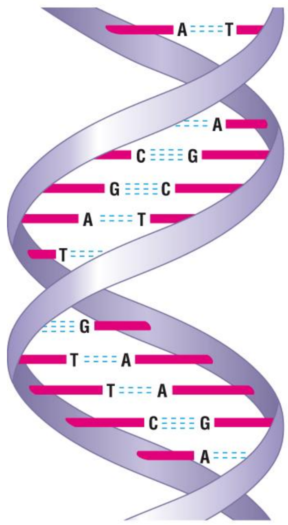


- Hydrogen bonds give water a skin effect and this contributes to the high surface tension.
- Other liquids with hydrogen bonding also have a surface tension.

Hydrogen bonding in biological molecules:

- Hydrogen bonds exist in biological molecules due to NH_2 and OH groups.
- Hydrogen bonding is responsible for the shapes of proteins and the double helix in DNA:





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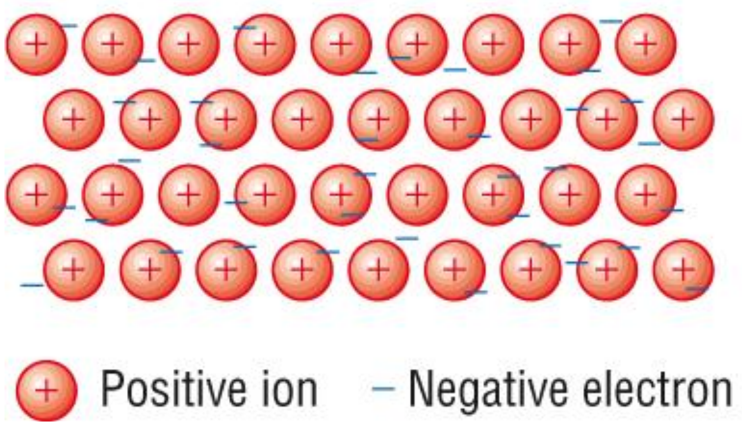
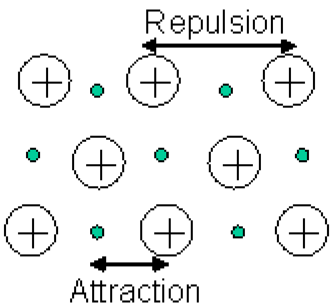
Metallic bonding and structure

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





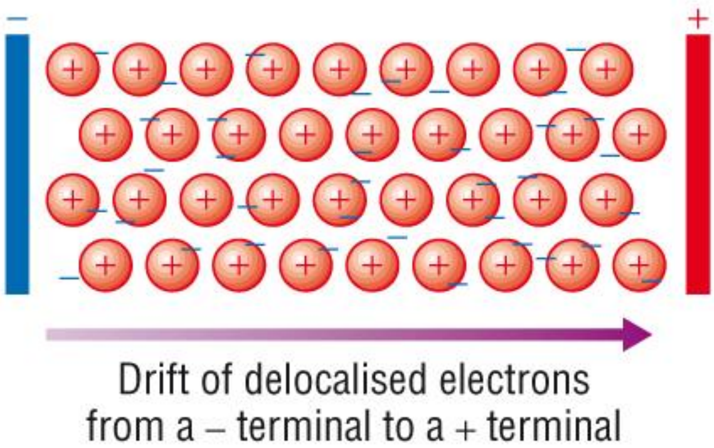
- The model consists of metal ions surrounded by '**Mobile sea of electrons**'.
- Attraction occurs between the ions and the delocalised electrons
- The sea of electrons explains its electrical conductivity and its thermal conductivity.
- The sea of electrons bonds the metal ions tightly into the lattice.
- This explains its high melting point.
- Since strong forces of attraction exists even in the liquid phase, metals tend to have a wide temperature range over which they remain liquid.
- The giant metallic lattice is often referred to as

Properties of giant metallic lattices:

1) High melting and boiling points -

- Attraction occurs between the fixed ions and the delocalised electrons.
- The attraction between the positive ions and the 'sea of electrons' is strong
- The sea of electrons bonds the metal ions tightly into the lattice.
- This explains its high melting point and boiling points.

2) Good electrical conductors -



- Mobile electrons will be attracted to a positive terminal.
- Electrons are replaced from the negative terminal.

3) 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.

Alloys

- These are mixtures of metals. This do not however form compounds.
- Compounds have a definite ratio of the elements.
- Metals can be mixed in different proportions.
- In the structure one metal ion is replaced with another.
- Often one ion is bigger than the original metal ion.
- This acts as a barrier preventing the layers from sliding past each other.
- This makes the alloy harder than the original metal.

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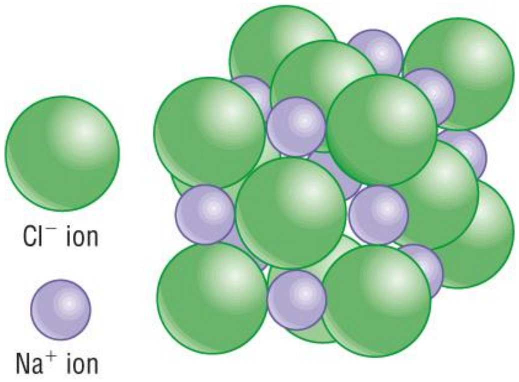
<u>Structure of ionic compounds</u>

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**

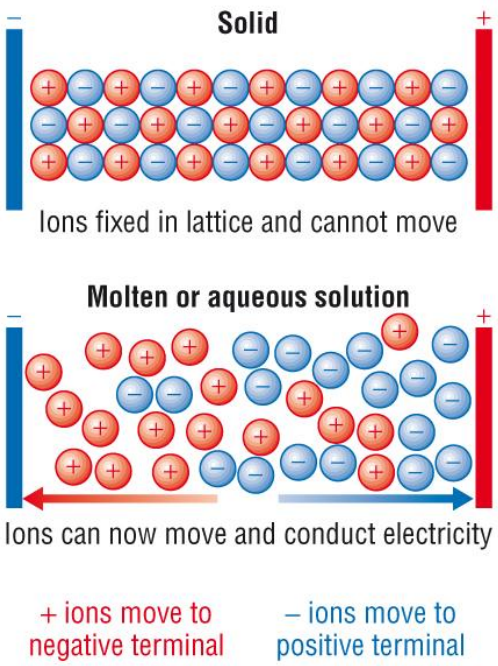
Properties of ionic compounds

1) High melting and boiling points:

- There are strong electrostatic forces of attraction between the ions.
- This means that they are not easily broken which is why they have high melting and boiling points.
- 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.

2) Electrical conductivity:

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

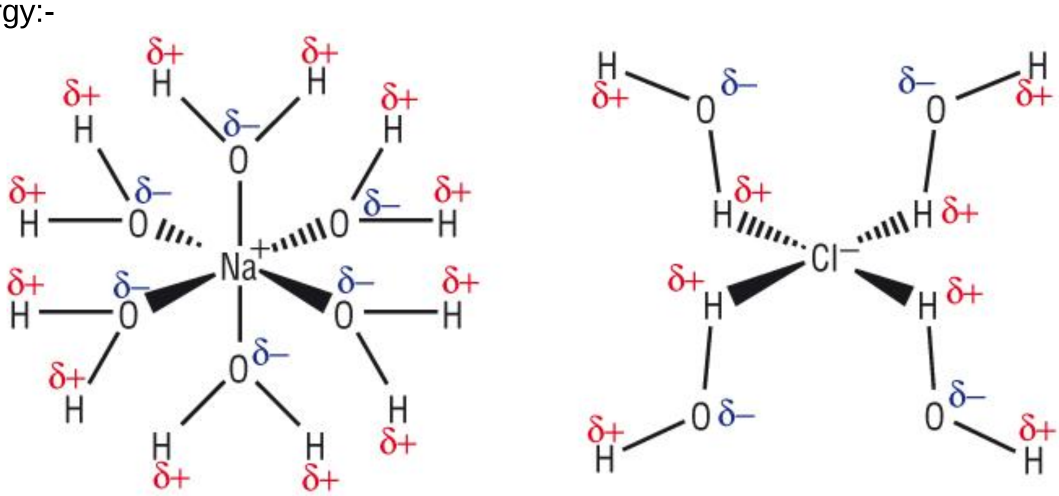


Solubility:

- 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:

Qu 1-3 P69

Structures of covalent compounds

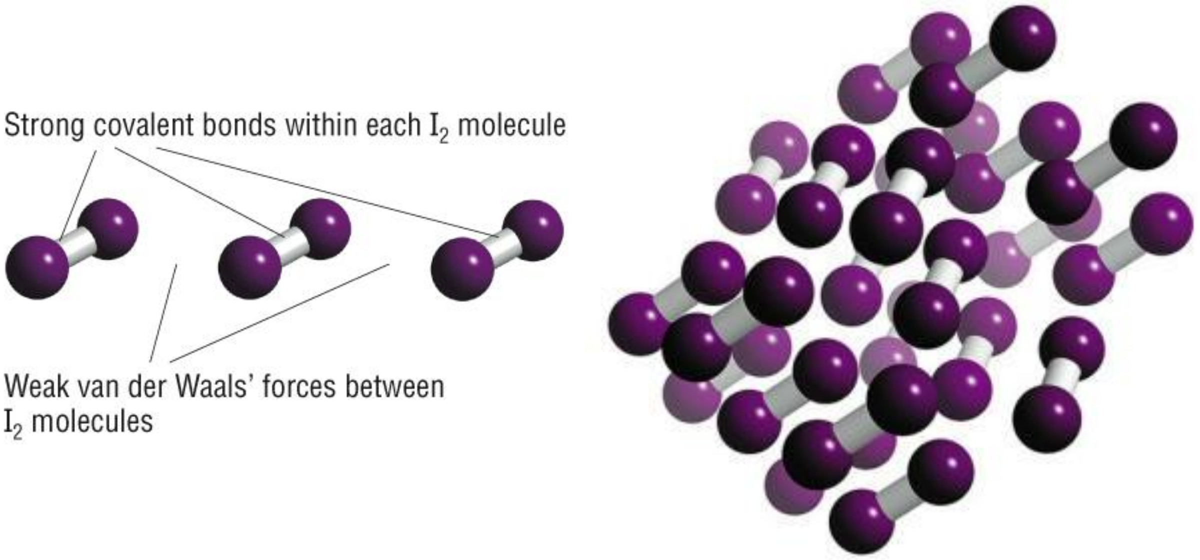
- Covalent compounds fall into 1 of 2 categories:

A) Simple molecular lattice

B) Giant covalent lattice

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.
- In its solid forms, the molecules are held together by weak intermolecular forces (VDW / Dipole / H Bonding).
- The atoms within the molecules are made up from strong covalent bonds.



Properties of simple molecular structures:

1) Low melting and boiling points:

- Simple molecular molecules have low melting and boiling points due to weak forces 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.
- From the animation - can you explain why water is a liquid at room temperature while carbon dioxide is a gas?

2) Electrical conductivity:

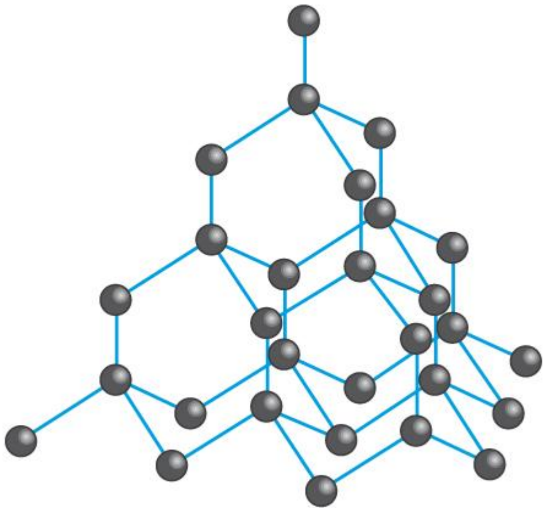
- Simple molecules have no free moving electrons or charges.
- This means they do not conduct electricity.

3) Solubility:

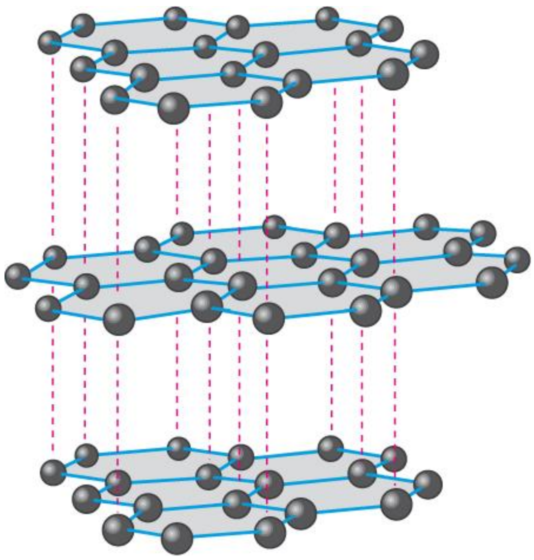
- Simple molecules are only soluble in non - polar solvents like hexane.
- This is because both molecule and solvent have weak forces of attraction between their molecules.

B) Giant covalent structures

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



Diamond



Graphite

Properties of giant covalent structures

1) High melting and boiling points:

- Due to the extensive covalent lattice structure of these types of compounds, a lot of energy is required to break these covalent bonds.
- This gives them very high melting and boiling points.
- Diamond is the hardest natural substance due to its covalent lattice structure.

2) Electrical conductivity:

- As there are no free moving electrons or charges, they are **non - conductors of electricity.**
- **Graphite** however is the **exception** to the rule.
- Each carbon in graphite has 3 covalent bonds.
- The 4th electron from each carbon atom is delocalised and free moving between the layers.
- This means it is able to conduct electricity.

3) Solubility:

- Giant covalent structures are **insoluble in both polar and non - polar solvents**.
- The covalent bonds are too strong to be over come by any solvent.

4) Hardness:

- Due to the extensive covalent lattice structure, all giant covalent structures are **hard**.
- **Graphite** however is the **exception** to this.
- Due to its layered structure and that the delocalised electrons are between the layers.
- This allows the layers to slide over and past each other.

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