

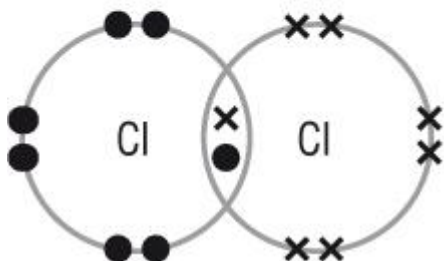
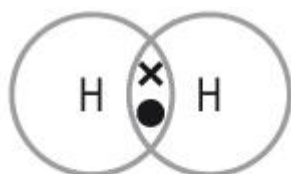
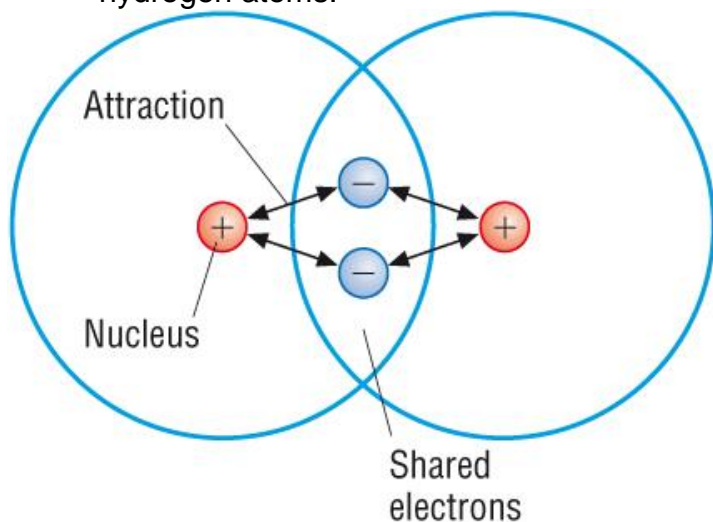
## 2C – Intermolecular forces, structure and properties:

### 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.

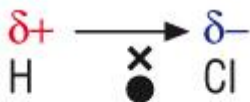


- As the 2 atoms are identical, they will have the same number of protons.
- This means that the electrons are being 'pulled' equally by both of the hydrogen atoms.
- We say that the H - H bond is **non-polar**
- **The same is true for any diatomic molecule where the atoms are identical:**

## 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 hydrogen 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,  $\delta^-$** .

### $\delta$ 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,  $\delta^+$** .
- 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 '

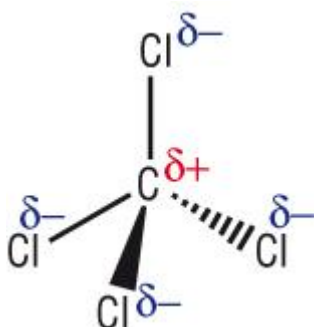
### Pole - poles (positive and negative)

### 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  $\text{CCl}_4$  is 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:

Electronegativity increases

Period																			He
1	H 2.20																		He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne	
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar	
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00	
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6	
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn	
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo	
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27			
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3			

## 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  $\delta^-$ .

## 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

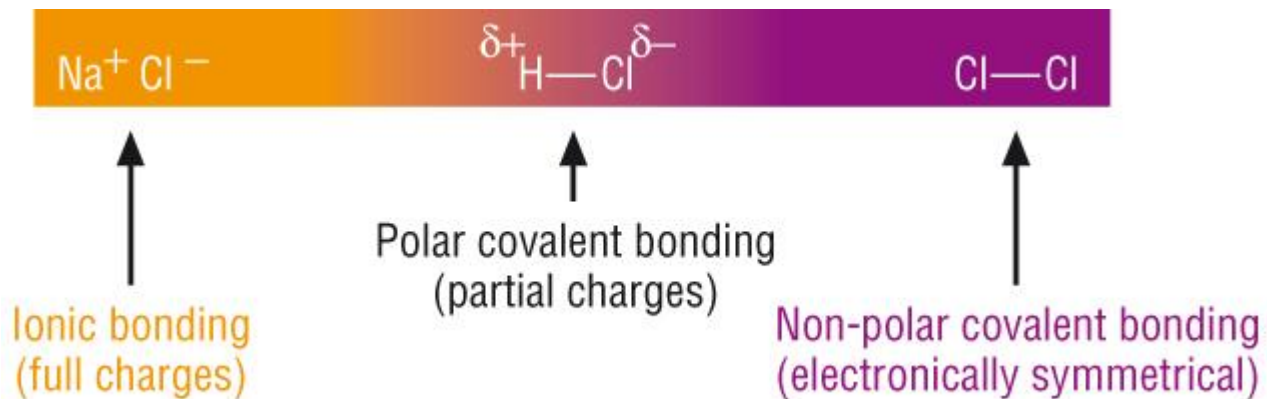
- Elements with very different electronegativities however have a different type of bonding
- The more electronegative element will attract the bonding electrons to itself so much that that element takes both of the bonding electrons.



- 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**.

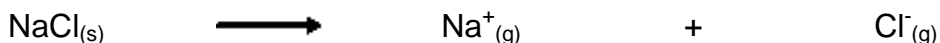


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## Intermolecular forces

### Strengths of bonds and forces:

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



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



- 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:

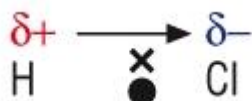
#### 1) Van der Waals' forces

#### 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  $\delta+$  end and a  $\delta-$  end, the molecule is said to have a **permanent dipole**:



- The  $\delta+$  end of one molecule will be attracted to the  $\delta-$  end of a neighboring molecule.
- This attraction is called a **permanent dipole - dipole force** of attraction:



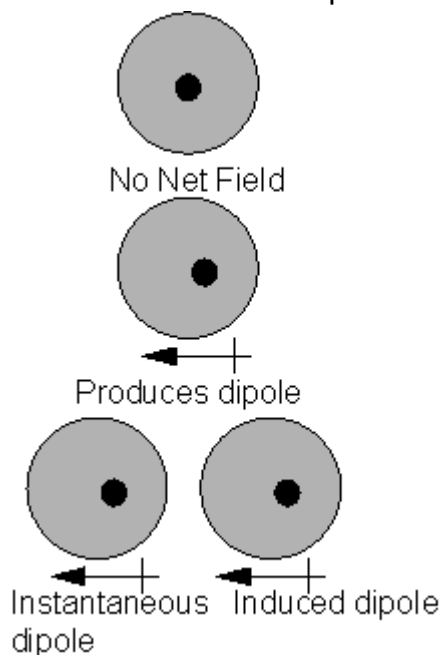
Dipole-dipole interaction between a  $\delta+$  atom of one molecule and a  $\delta-$  atom of another molecule.

#### 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.  $0.105\text{KJmol}^{-1}$
- 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**.

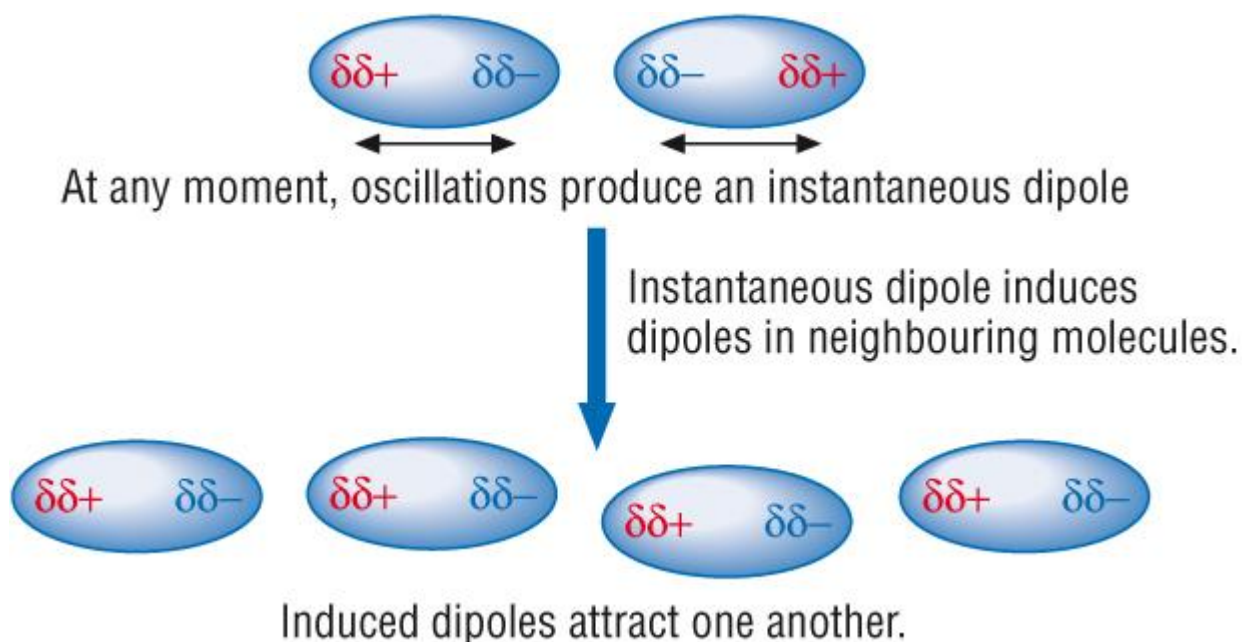
## What causes Van der Waals' forces:

- 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  $\delta\delta$  to represent extra small charges:



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

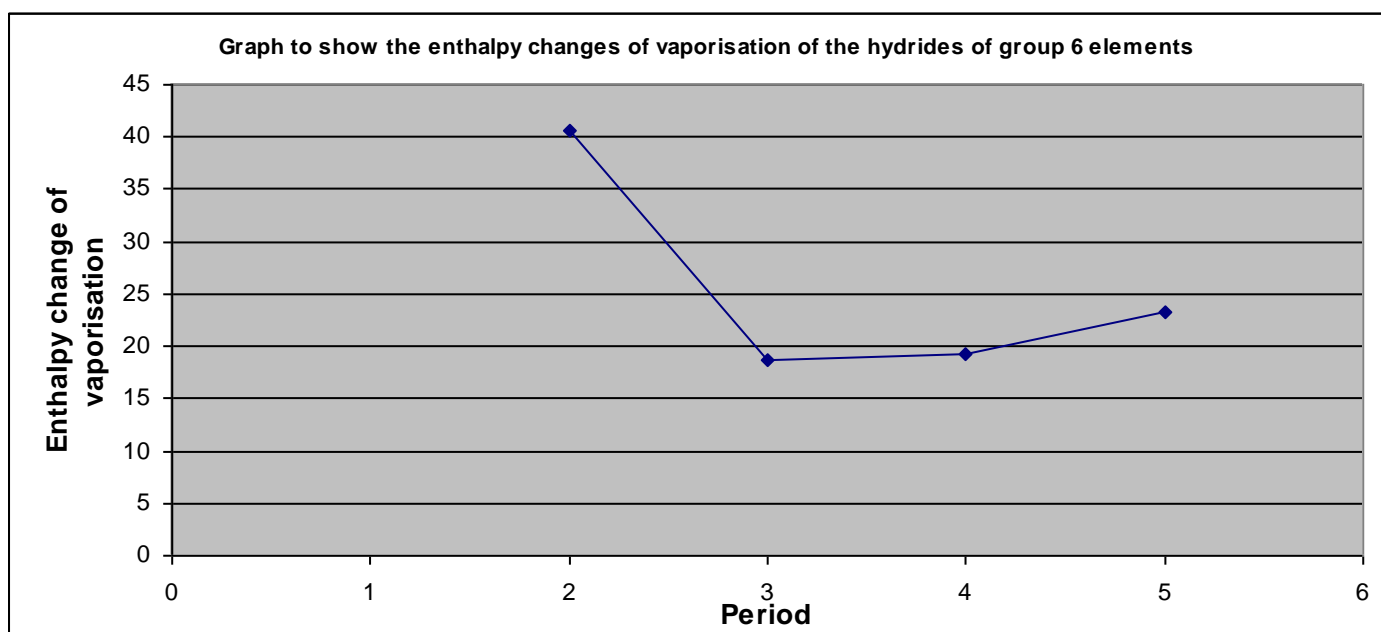
## Hydrocarbons Diatomic elements

- 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.

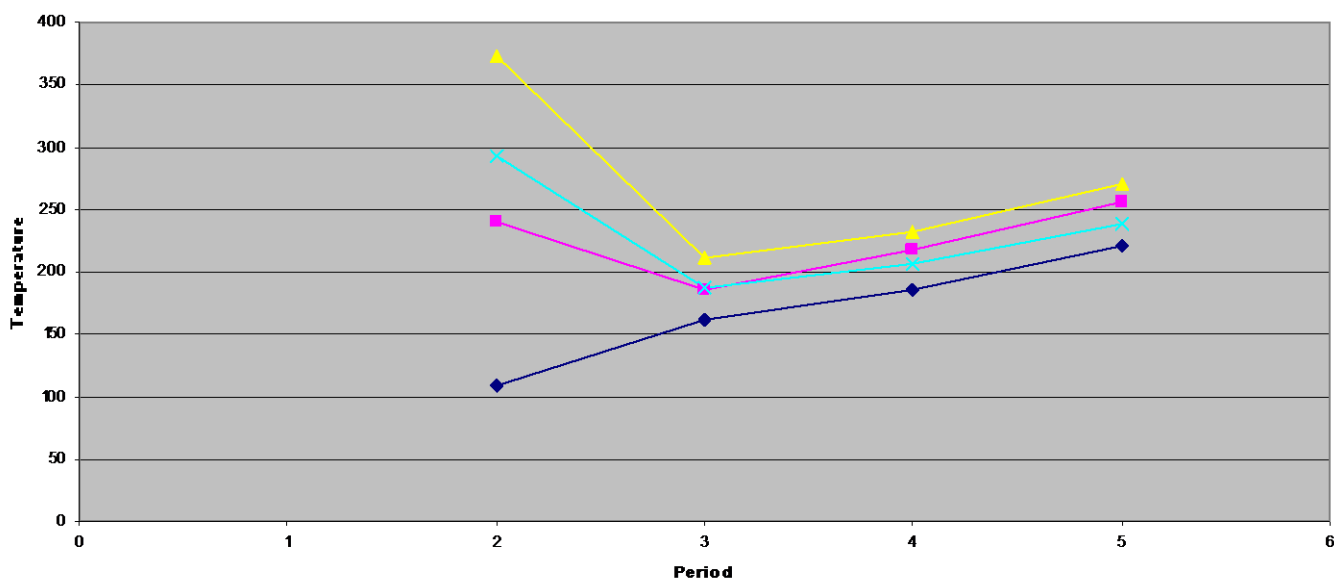
## Questions 1-2 P63

### Hydrogen Bonding:

### Water is peculiar



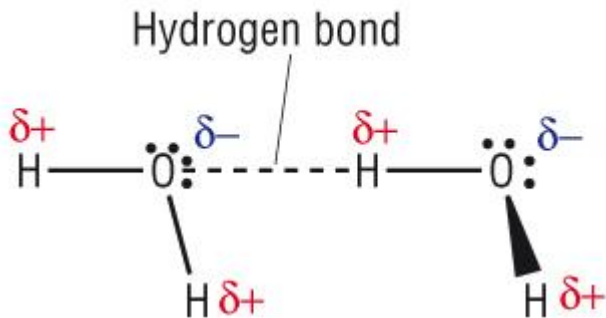
Graph to show Boiling Points of group 4-7 hydrides



- 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  $\delta^+$  and  $\delta^-$  charges on different water molecules.

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

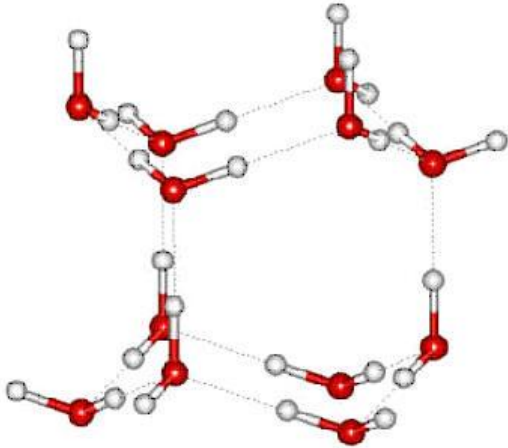
## Summary

- 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:

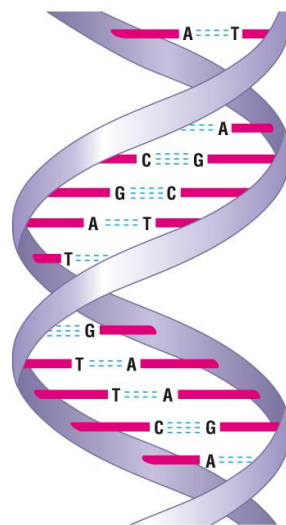
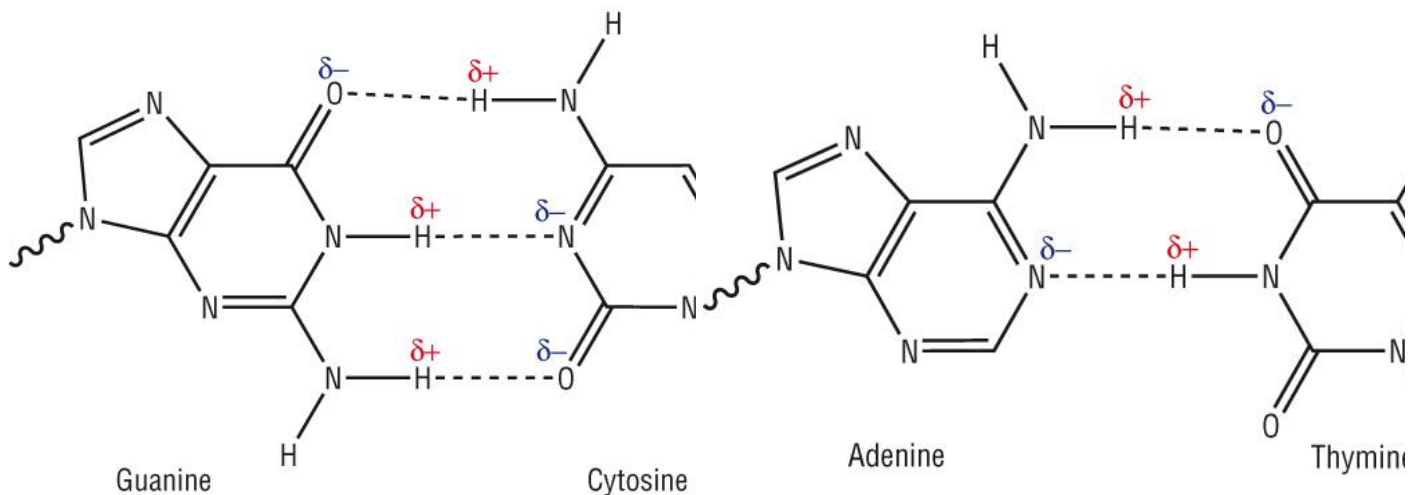
- The 2 hydrogen bonds per water molecule of water sets up a 3D structure.
- The Hydrogen bonds are longer than the O - H bonds.
- This means they are held further apart than in water making ice less dense.



- 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.

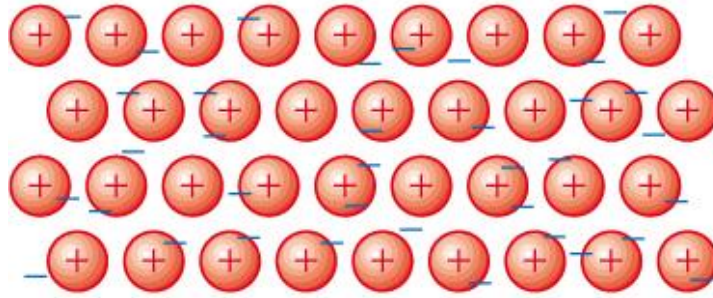
### Hydrogen bonding in biological molecules:

- Hydrogen bonds exist in biological molecules due to  $\text{NH}_2$  and  $\text{OH}$  groups.
- Hydrogen bonding is responsible for the shapes of proteins and the double helix in DNA:

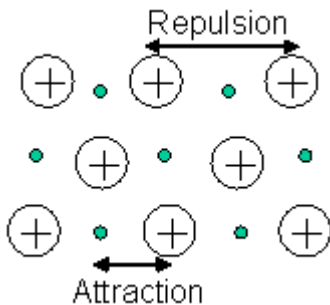


## Metallic bonding and structure

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



 Positive ion     Negative electron



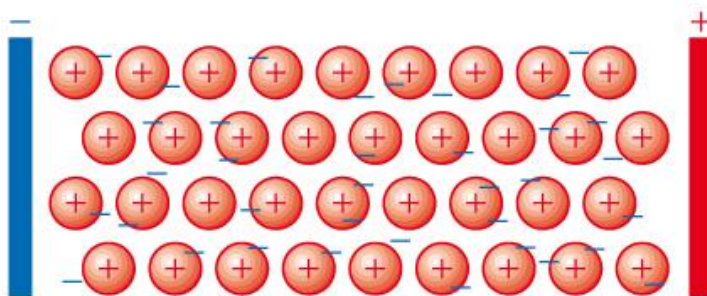
- The model consists of metal ions surrounded by '**Mobile sea of electrons**'.
- Attraction occurs between the ions and the delocalised electrons
- The sea of electrons explains its electrical conductivity and its thermal conductivity.
- The sea of electrons bonds the metal ions tightly into the lattice.
- This explains its high melting point.
- Since strong forces of attraction exist 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.

Drift of delocalised electrons  
from a - terminal to a + 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.

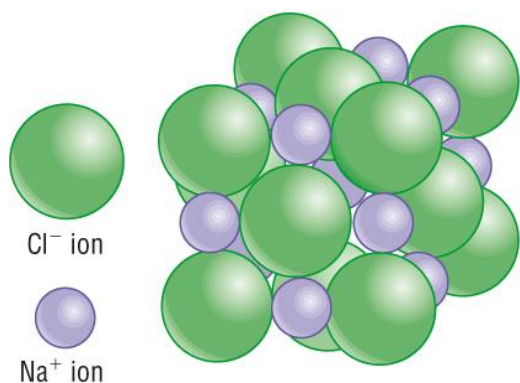
### Alloys

- 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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### Structure of ionic compounds

#### 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 Lattice**

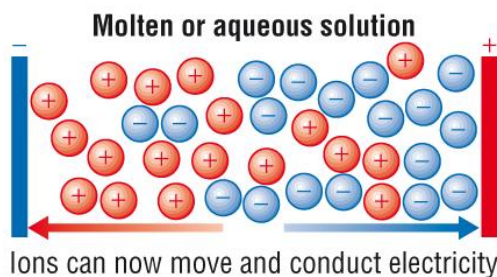
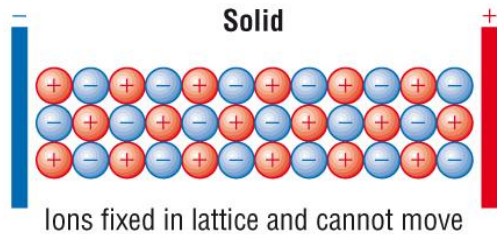
#### Properties of ionic compounds

##### 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:



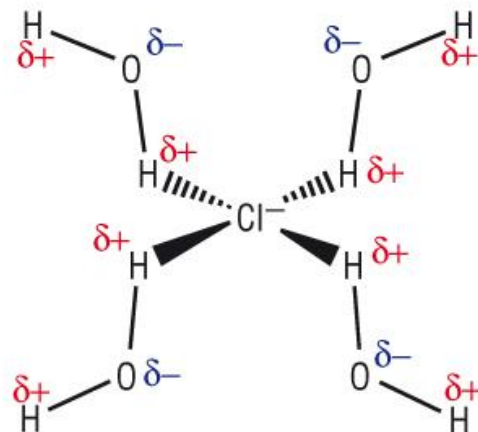
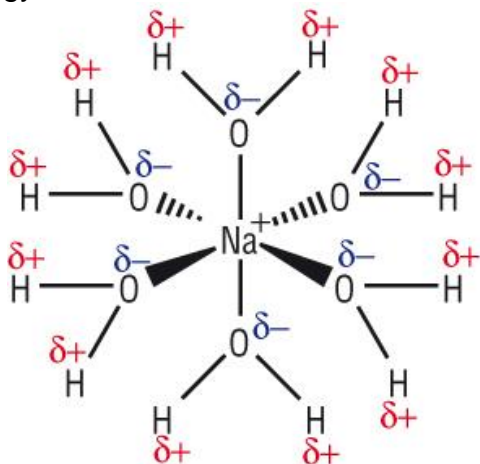
+ ions move to negative terminal      - ions move to positive terminal

## 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  $\delta+$  /  $\delta-$  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

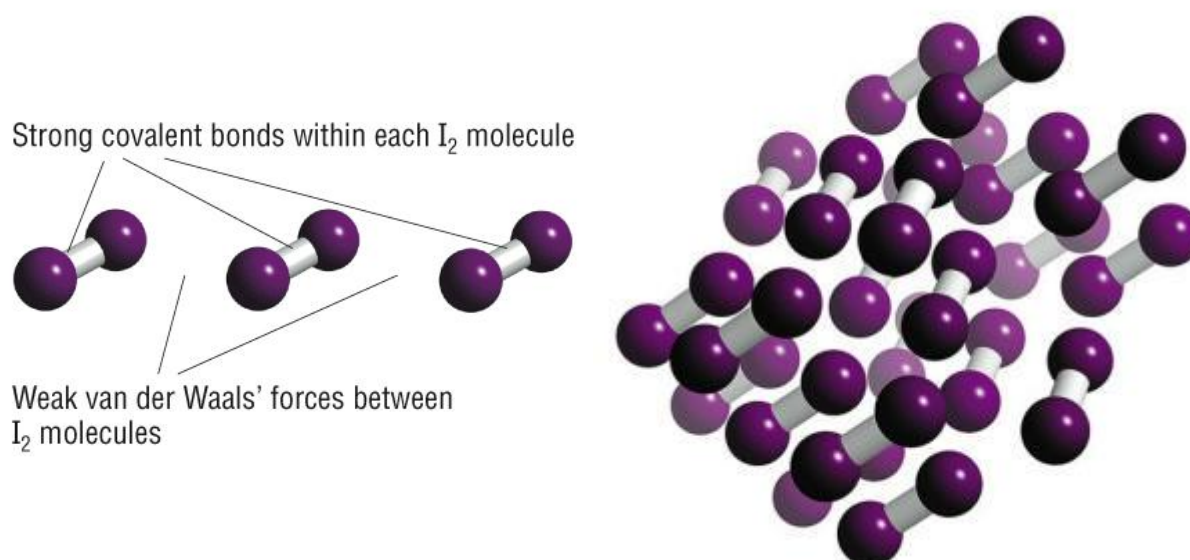
- Covalent compounds fall into 1 of 2 categories:

#### A) Simple molecular lattice

#### B) Giant covalent lattice

#### A) Simple molecular structures:

- These are made up from simple (small) molecules such as:  $\text{CO}_2$ ,  $\text{N}_2$ ,  $\text{O}_2$ ,  $\text{I}_2$  and  $\text{H}_2\text{O}$ .
- 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:

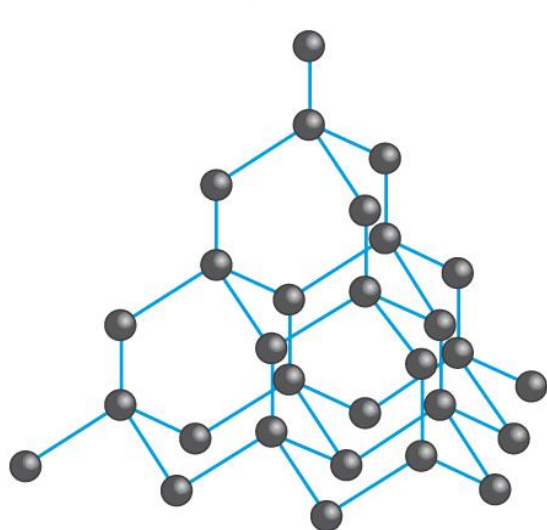
- Simple molecules have no free moving electrons or charges.
- This means they do not conduct electricity.

##### 3) Solubility:

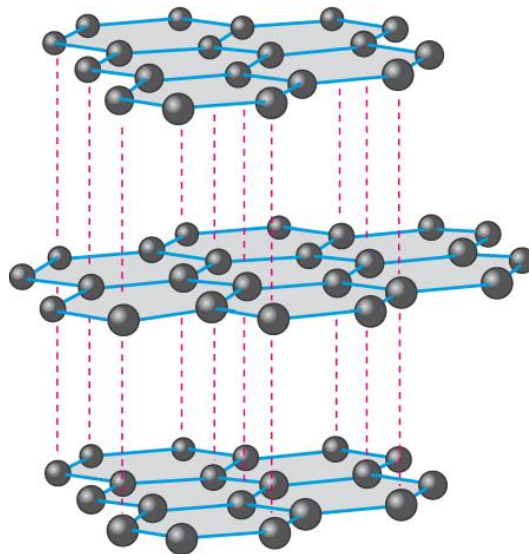
- Simple molecules are only soluble in non - polar solvents like hexane.
- This is because both molecule and solvent have weak forces of attraction between their molecules.

## 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.

### Questions 1-2 P71