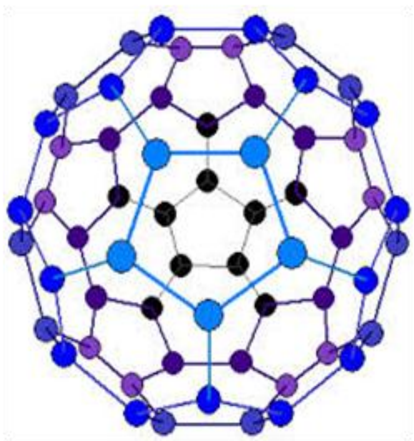
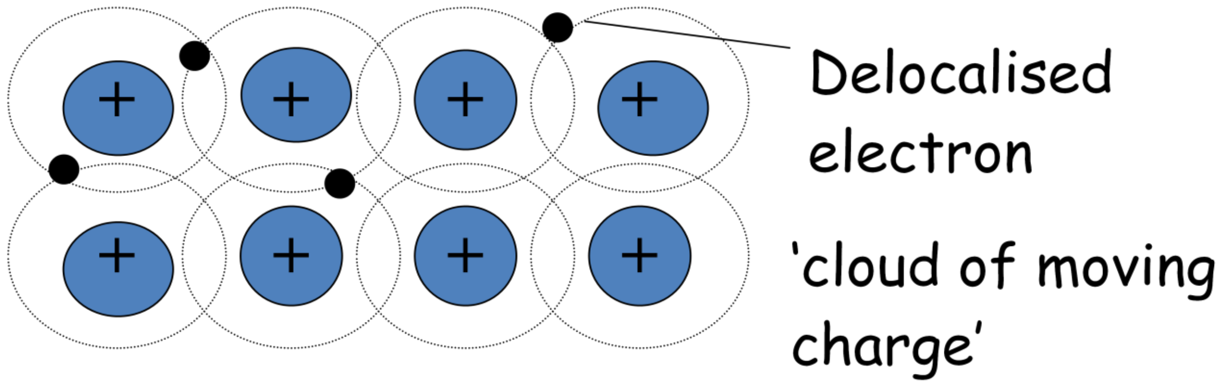
Cathkin High School
CfE Higher Chemistry

Unit 1 : Part 3

Chemical Changes & Structure
Structure and Bonding





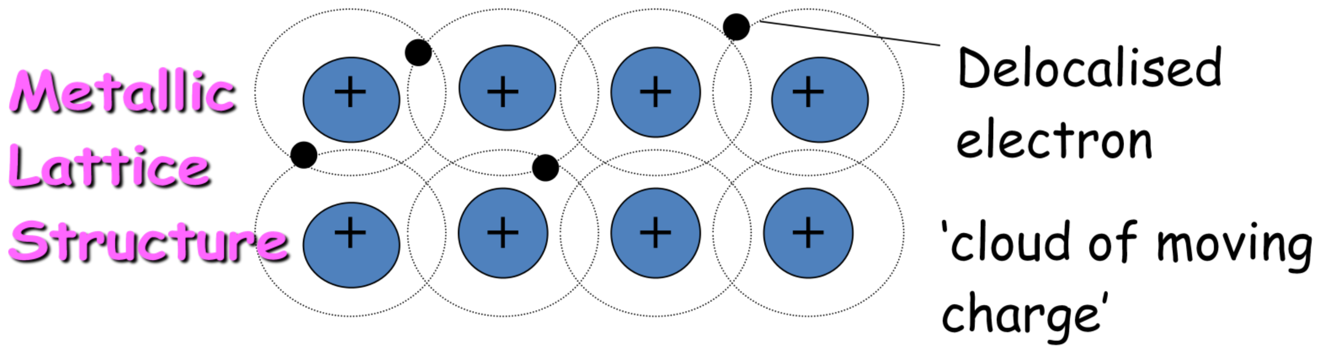
No.	Learning Outcome	Understanding?
1	The bonding types of the first twenty elements; metallic (Li, Be, Na, Mg, Al, K and Ca); covalent molecular (H_2, N_2, O_2, F_2, Cl_2, P_4, S_8 and C_{60} [fullerenes]); covalent network (B, C (diamond, graphite), Si) and monatomic (noble gases)	☺ 😐 ☹
2	Describe the bonding continuum moving from pure non-polar covalent to ionic.	☺ 😐 ☹
3	Explain how polar covalent bonds arise	☺ 😐 ☹
4	Explain how van der Waals forces arise between molecules.	☺ 😐 ☹
5	Describe what causes dispersion forces to exist between gaseous atoms and molecules.	☺ 😐 ☹
6	Explain how the polarity of molecules affects the strength of dispersion forces.	☺ 😐 ☹
7	Explain why certain molecules have a stronger type of van der Waal force called a hydrogen bond	☺ 😐 ☹
8	Explain how the properties of substances are affected by the type of bonding that they exhibit..	☺ 😐 ☹
9	Predict the solubility of a substance from information about solute and solvent polarities.	☺ 😐 ☹

The Formation of Bonds

Bonds are **electrostatic forces** (attractions between positive and negative charges) which hold atoms together.
Atoms form bonds to become more stable - by losing, gaining or sharing electrons.
The **type of bond** formed in a substance depends on the elements involved and their position in the periodic table.

Metallic Bonding

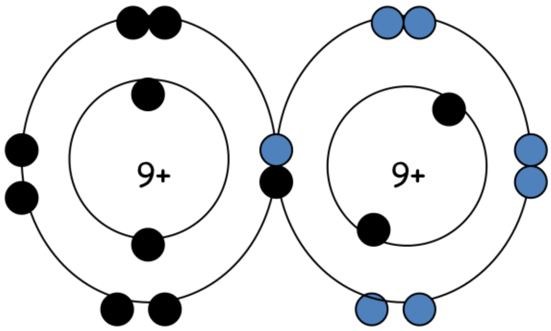
Metallic bonding occurs between the atoms of metal elements.
Metals have little attraction for their outer electrons. These electrons are free to move so are delocalised.



Electrons can move randomly between these partially filled outer shells creating what is called a 'sea' or 'cloud' of electrons around positive metal ions.
The metallic bond is the electrostatic force between positively charged ions and delocalised outer electrons.

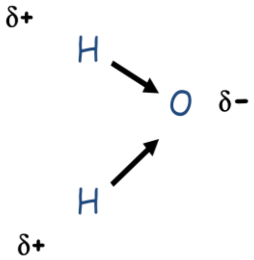
Covalent and Polar Covalent Bonding

Covalent bonding occurs in **non-metal elements.**
A covalent bond is the electrostatic force of attraction between positively charged nuclei and negatively charged outer electrons.
In **non-metal elements**, e.g the diatomic elements, there is **an equal sharing of electrons** between atoms as they have the **same electronegativity**.
e.g. Fluorine

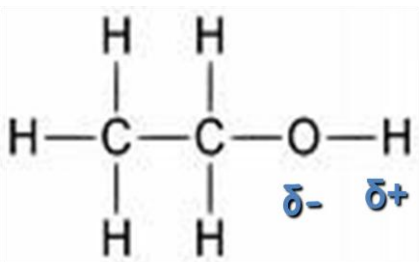


Polar covalent bonding is similar to covalent bonding but is formed in compounds between different non-metal elements as they have **different electronegativities.**
In polar covalent bonds there is an uneven share of electrons.
e.g. **Hydrogen Fluoride** **Water**

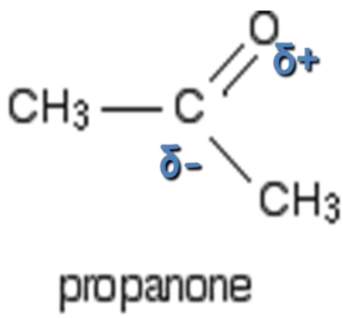
$^{\delta+}$ **H-F**$^{\delta-}$



Ethanol



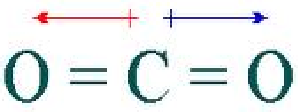
Propanone



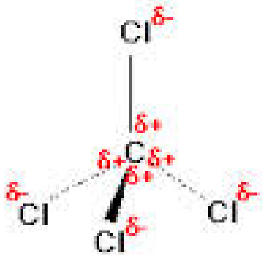
Polar Molecules and Permanent Dipoles

Not all substances with polar covalent bonds will have 'polar molecules'.
If there is a symmetrical arrangement of polar bonds, the polarity cancels out over the molecule as a whole.
e.g.

Carbon dioxide



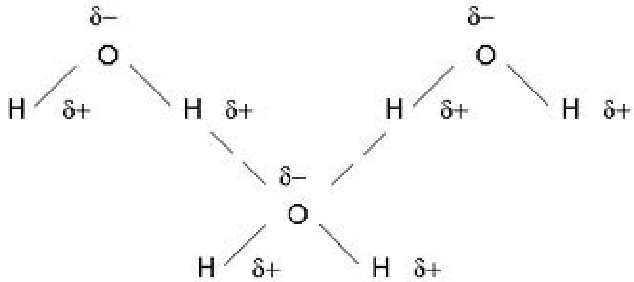
Tetrachloromethane



If the bonds are not symmetrical, the molecule has an overall polarity and is said to have a **permanent dipole**, i.e. each end has a different charge.

e.g. Hydrogen Chloride

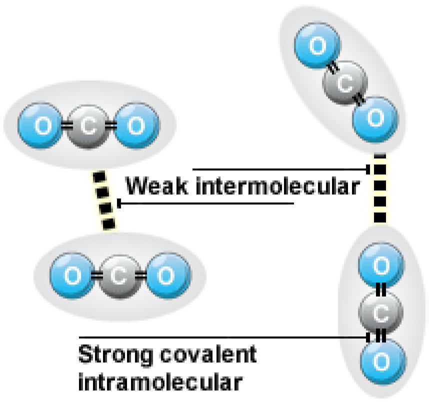
Water



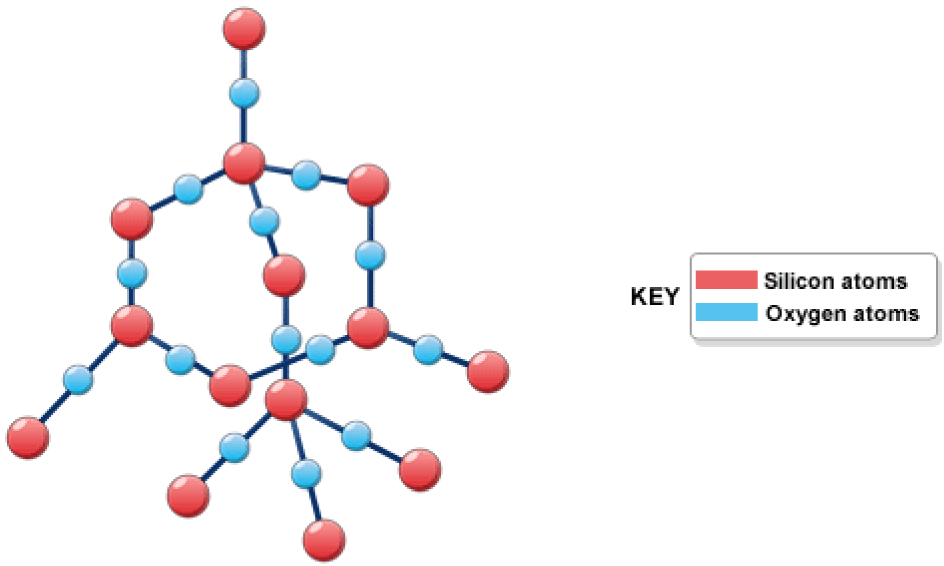
Covalent Structure

Covalent and polar covalent substances are usually made up of discrete molecules, but a few have giant covalent network structures.

e.g. Carbon dioxide – discrete molecules



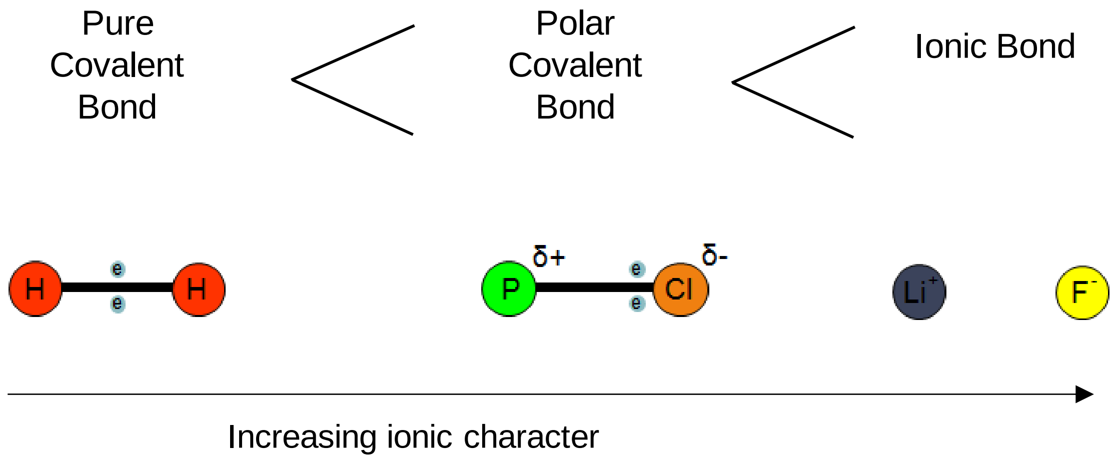
Silicon Dioxide – covalent network structure



(images from BBC Higher Bitesize Chemistry)

The Bonding Continuum

The greater the difference in electronegativity between two elements, the less likely they are to share electrons, i.e. form covalent bonds.



To judge the type of bonding in any particular compound it is more important to look at the properties it exhibits rather than simply the names of the elements involved.

Ionic Bonding

Ionic bonds are formed between **metal and non-metal** elements with a **large difference in electronegativity**.
The non-metal element with the high electronegativity gains the electrons to form a negative ion:

e.g. $Cl + e^- \qquad Cl^-$

The element with the low electronegativity loses electrons to form a positive ion:

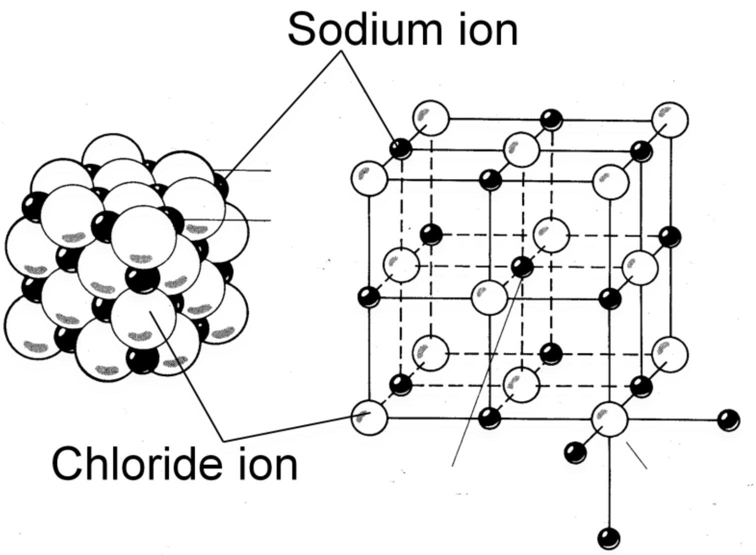
e.g. $Na \qquad Na^+ + e^-$

Both the positive and negative ion will have the same electron arrangement as a noble gas.
Ionic bonding is the electrostatic force of attraction between positively and negatively charged ions.

Structure of Ionic Compounds

The forces of attraction between the oppositely charged ions results in the formation of a regular structure called an **ionic lattice.**

E.g. Sodium chloride



Each Na^+ ion is surrounded by 6 Cl^- ions.

The formula of sodium chloride is NaCl, showing that the ratio of Na^+ to Cl^- ions is 1 to 1

Bonding Between Molecules

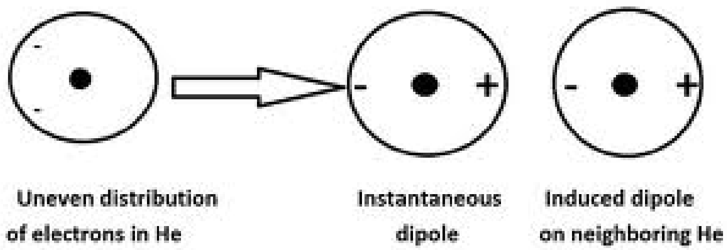
There are attractive forces between covalent and polar covalent molecules which can affect their properties.
These attractions **between** molecules are called Van der Waals or **intermolecular forces** (or bonds).
(intramolecular = within the molecule, e.g. covalent bond)

There are 3 types:

1. London Dispersion Forces
2. Dipole-dipole Attractions (permanent dipole-permanent dipole)
3. Hydrogen Bonds are a special type of dipole-dipole attraction which is particularly strong.

1. London Dispersion Forces

This is the weakest form of intermolecular bonding and it exists between all atoms and molecules.
Dispersion forces are caused by uneven distributions of electrons.

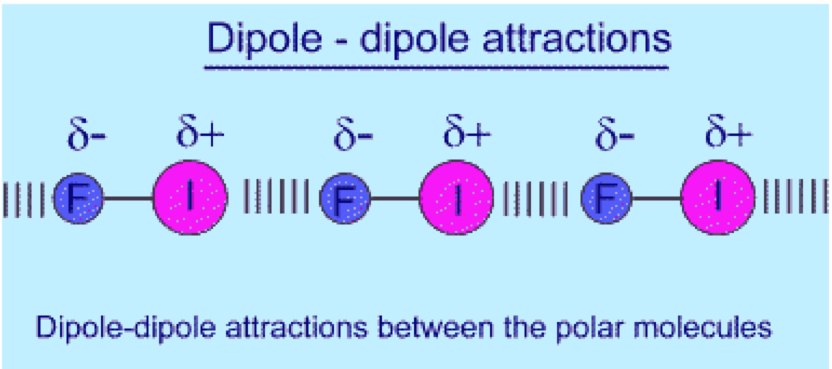


The atom or molecule gets slightly charged ends known as a **temporary dipole**. This charge can then induce an opposite charge in a neighbouring atom or molecule called an **induced dipole.** The **oppositely charged** ends **attract** each other creating the intermolecular force

The relative strength of the force depends on the size of the atoms or molecules.
Dispersion forces increase with increasing atomic and molecular size.
energy

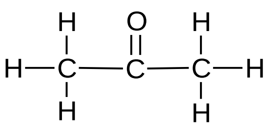
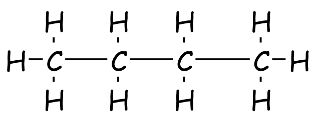
2. Permanent Dipole-Permanent Dipole Attractions

A **polar molecule** is one which has permanently charged ends (**permanent dipole**).
Polar-Polar attractions (permanent dipole-permanent dipole) are the intermolecular force of attraction between the oppositely charged ends of the **polar molecules.**



These forces of attraction between polar molecules are **in addition to London Dispersion Forces.**

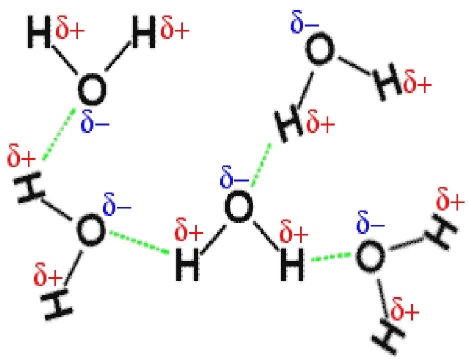
Effect of dipole-dipole attractions

	Propanone	Butane
Formula Mass	58	58
Structure		
Intermolecular forces	London + permanent dipole-permanent dipole	London
Boiling Point	56°C	0°C

Polar molecules have higher boiling points than non-polar molecules of a similar mass due to the permanent dipole-permanent dipole interactions.
Permanent dipole-permanent dipole interactions are stronger than London Dispersion forces.

3. Hydrogen Bonding

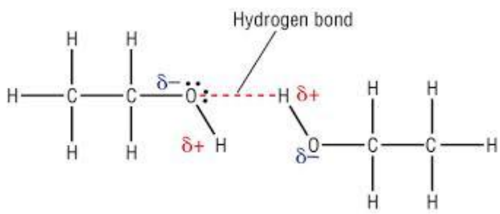
Hydrogen bonds are **permanent dipole-permanent dipole** interactions found between molecules which contain **highly polar bonds.**



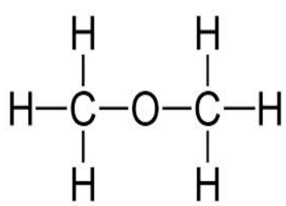
They are usually found in molecules where **hydrogen** is **bonded** to **very electronegative** atoms like **fluorine, oxygen or nitrogen (+ chlorine).**
Other examples include ammonia, alkanoic acids and alkanols.
Hydrogen bonds are **stronger** than **permanent dipole-permanent dipole** attractions and **Van der Waals** but weaker than covalent bonds.

Effects of Hydrogen Bonding

When Hydrogen bonds are present, the compound will have a much **higher melting point** (m.pt) and **boiling point** (b.pt) than other compounds of similar molecular size.
E.g. Ethanol Ether



From this diagram we



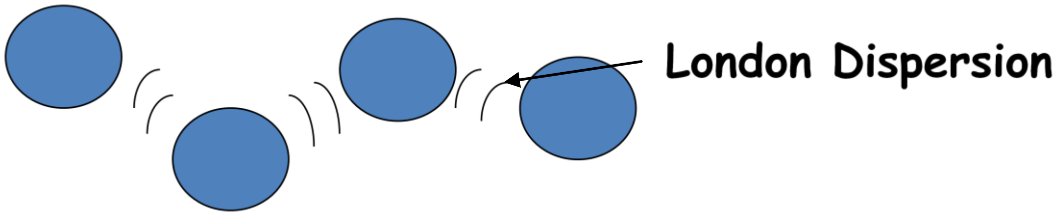
Bonding and Properties of Elements 1-20

Monatomic Elements - Noble Gases

Bonding

All consist of single, unbonded atoms.

Only have London Dispersion forces between the atoms.



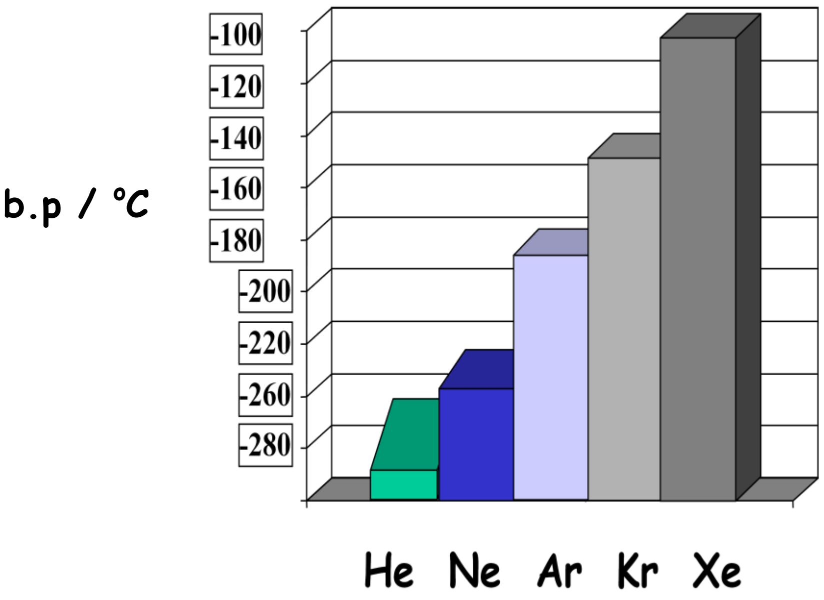
Properties

Low densities, m.pts and **b.pts**

Non conductors of electricity as no freely moving charged particles.

B.pts increase as the size of the atom increases

This happens because the dispersion forces increase



Covalent Molecular Elements (in 1-20)

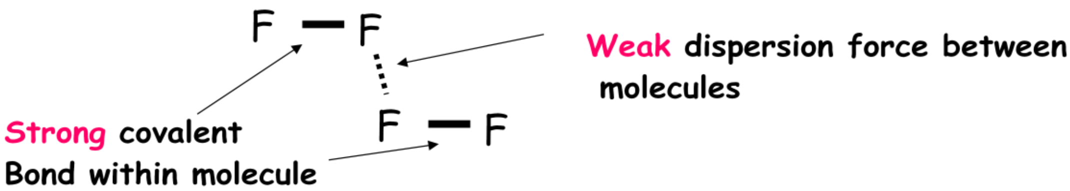
All consist of discrete molecules of varying size.

Fairly low m.pts, b.pts and **densities.**

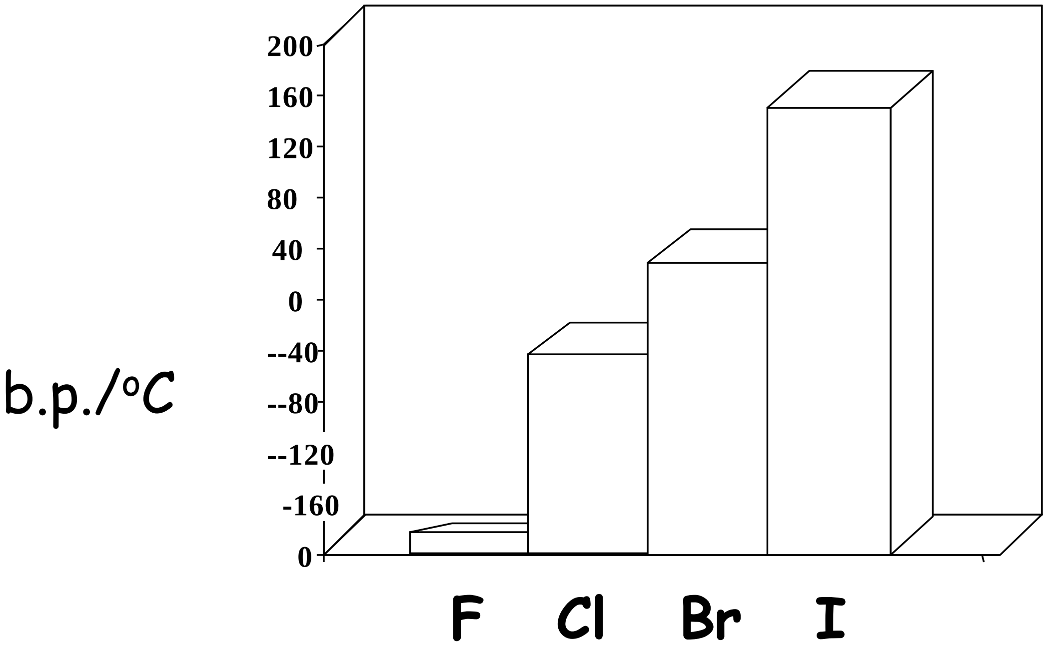
Non-conductors of electricity.

Diatomic elements – H_2, N_2 , O_2 , F_2 , Cl_2

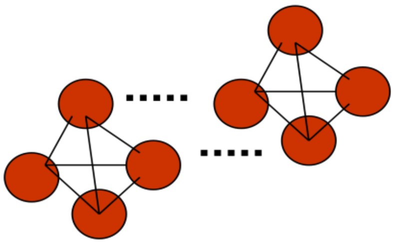
H		
N	O	F
P	S	Cl



As the size of the halogen atom increases, so does the strength of the London dispersion forces.

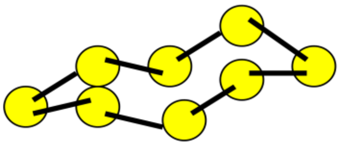


Phosphorus – P_4



m.pt 44°C

Sulphur - S_8



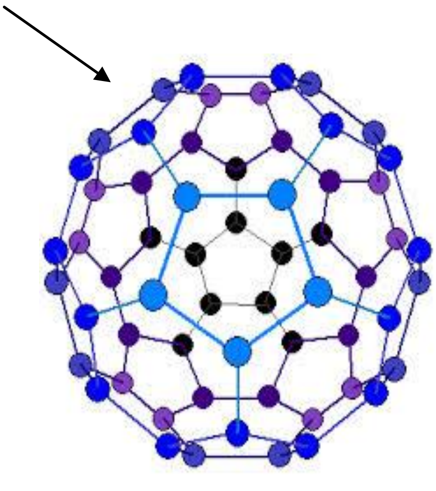
m.pt 113°C

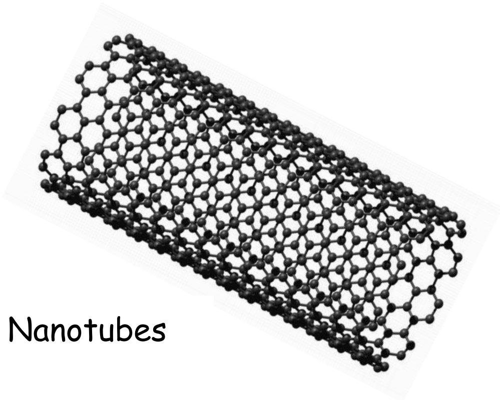
Higher m.pt. because there are stronger London Dispersion forces between larger molecules.

Fullerenes (Carbon)

C_{60} C_{70} C_{240}

Buckminster fullerene C_{60} (Bucky Balls) discovered in the 1980's





Nanotubes

Due to the large molecules , fullerenes have stronger dispersion forces between their molecules than smaller molecules.

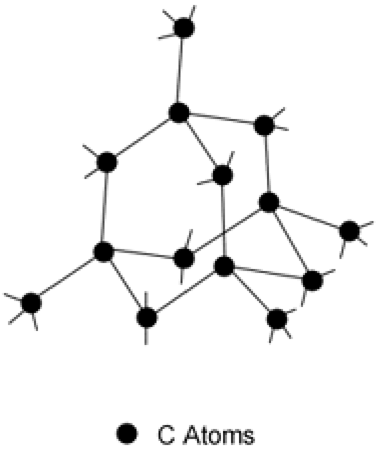
NB – they are molecules not covalent networks

Covalent Network Elements (in 1-20)

Giant network structures containing millions of atoms.
E.g. Carbon exists in 2 main forms…

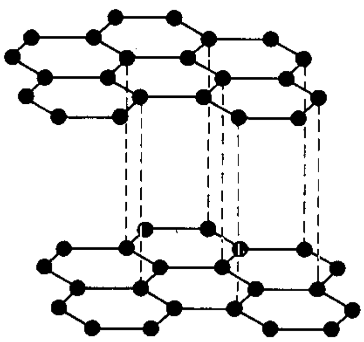
B	C
Si	

Diamond



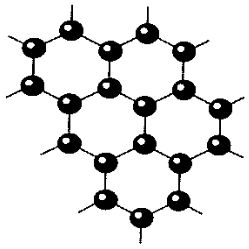
4 bonds per carbon atom – tetrahedral structure
Non-conductor of electricity as no free electrons.
Hardest natural substance as many strong bonds to break so used for drills, cutting tools, etc.

Graphite



3 bonds per carbon atom – layered structure with London dispersion forces between the layers

Conductor of electricity due to delocalised electrons between the layers – used in electrodes.

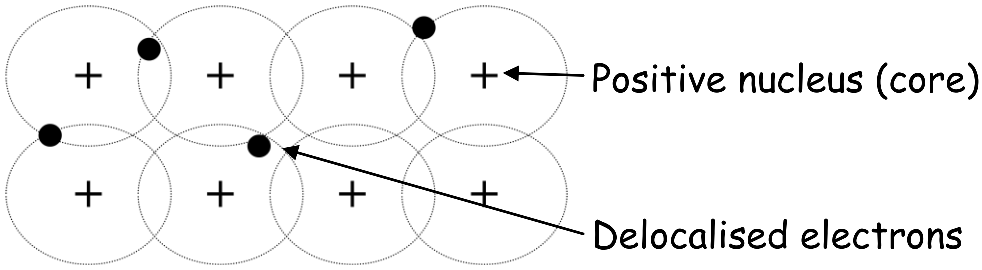


Very soft – the layers break away easily due to weak dispersion forces so good as a lubricant and for drawing (pencils).

Metallic elements (Revision of Nat 5)

Li	Be	
Na	Mg	Al
K	Ca	

All have metallic lattice structure
Solids (except Hg) with high densities, m.pts and b.pts due to the closely packed lattice structure with lots of bonds to break.
M.pts are relatively low compared to the B.Pts as when a metal is molten the metallic bond is still present.
B.pts are much higher as you need to break the metallic bonds throughout the metal lattice.



Metal b.p.'s are dependent on
(i) How many electrons are in the outer shell
(ii) How many electron shells there are.

In a period, the greater the number of electrons in the outer shell the **stronger** the metallic bond.
So the melting point of Al>Mg>Na

Conductors of electricity when solid or liquid due to delocalised outer electrons which are free to move.

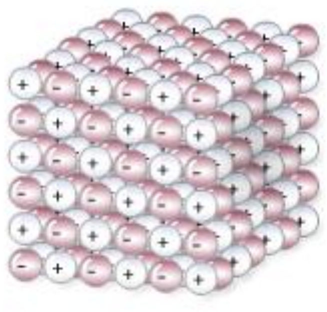
Bonding and Properties of Compounds

Compounds can be split into 3 main groups, depending on their bonding, structure and properties:

1. Ionic Lattice Structures
2. Covalent Network Structures
3. Covalent Molecular Structures

1. Ionic Lattice Structures

All ionic compounds are solids at room temp so have high melting and boiling points.
This is because the ionic bonds holding the lattice together are strong and a lot of energy is required to break them.

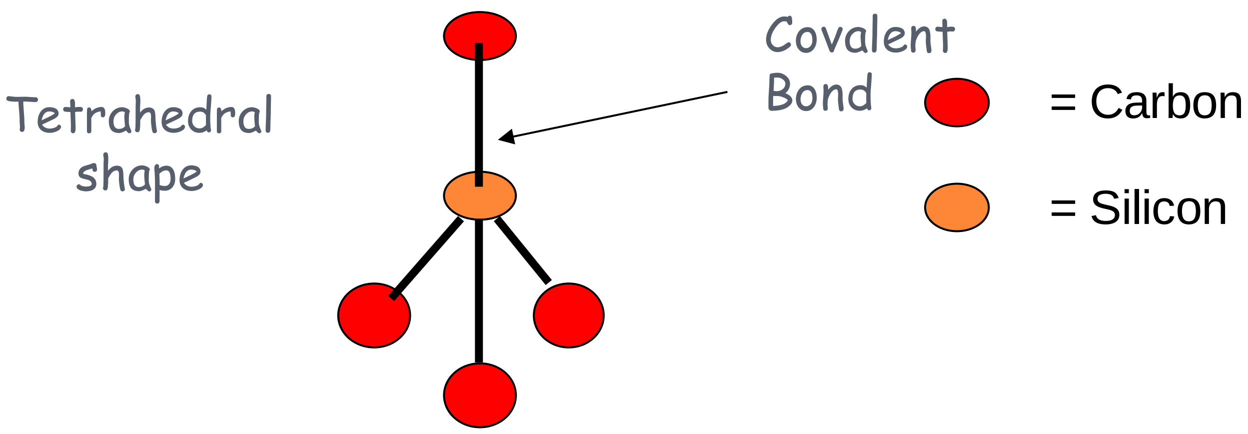


Ionic lattice structure

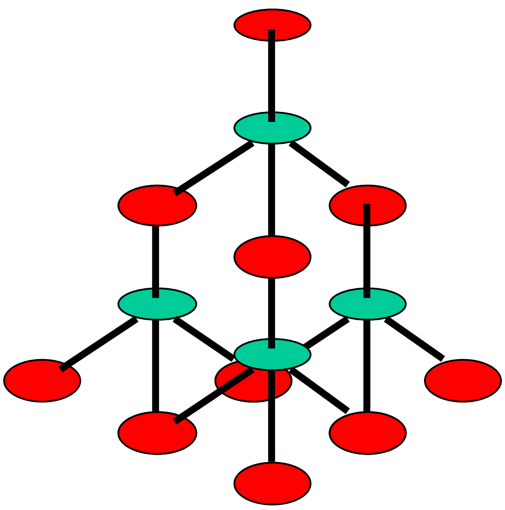
The size of the ions will effect
the strength of the ionic bond and how the ions pack together.
E.g. NaF - m.p 1000°C, NaI - 660°C.
Ionic compounds conduct electricity when dissolved in water or when molten as the ions are free to move.
Electrolysis of an ionic solution or melt causes a chemical change at the electrodes.
They do not conduct when solid as the ions are 'locked in the lattice and cannot move to carry the current.

2. Covalent Network Structures

Covalent networks have **very high melting and boiling points** as many strong covalent bonds need to be broken in order to change state.
They can also be very hard.
E.g. Silicon Carbide (SiC) – carborundum, similar structure to diamond



The 4 carbon atoms are available to bond with another 4 silicon atoms resulting in a covalent network.



It has a high melting point (2700°C)
SiC is used as an abrasive.

Covalent network structures are usually non-conductors of electricity as they have no free moving charged particles.

2. Covalent Molecular Structures

Usually have **low melting and boiling points** as there is little attraction between their molecules.
E.g. Carbon dioxide CO_2: m.pt -57°C (non-polar)

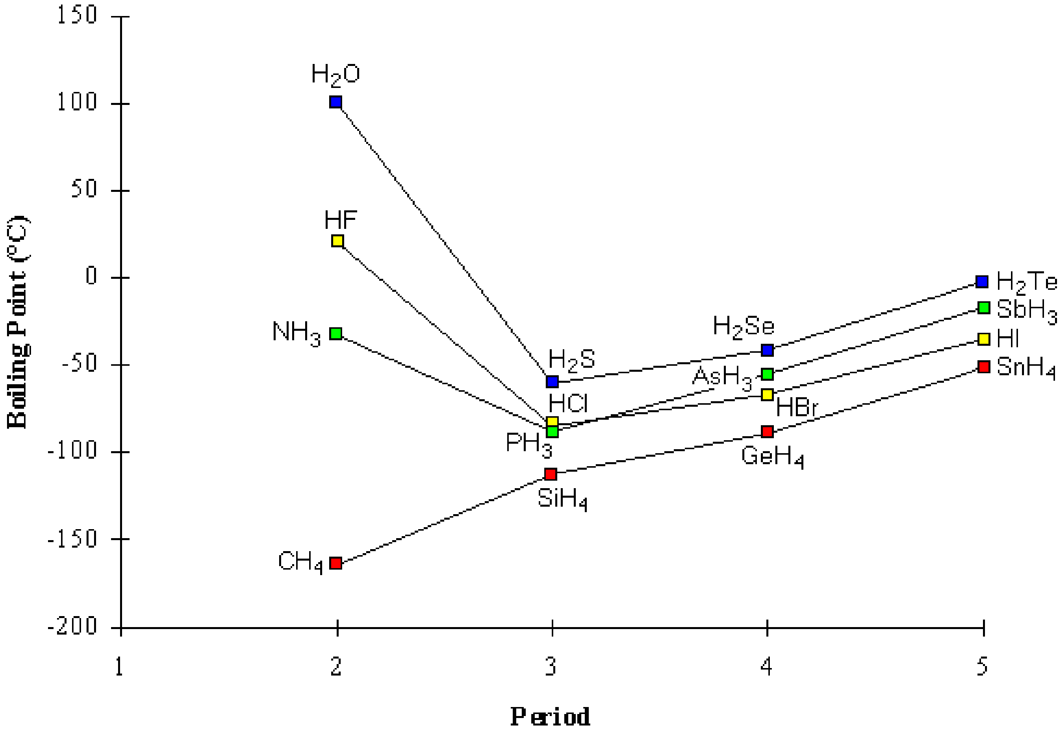
Compounds with **polar molecules** may have **slightly higher m.pts** and **b.pts** than non-polar molecules due to permanent dipole-permanent dipole attractions.

e.g. Iodine chloride	Bromine
I - Cl	Br – Br
b.pt 97°C	b.pt 59°C

When **hydrogen bonds** are present, the compounds will have a much **higher m.pt and b.pt** than other compounds of similar molecular size as more energy is required to separate the molecules. (see earlier note on hydrogen bonding – ethanol and ether)

Physical properties of hydrides

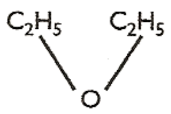
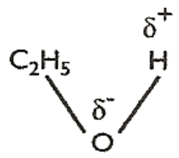
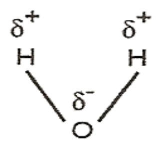
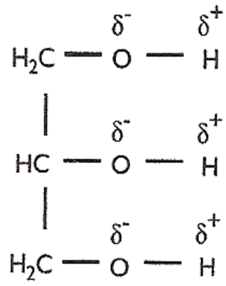
Water has a much higher b.p. than similar compounds containing hydrogen



Hydrogen bonding explains why water, HF and NH_3 have a b.p. higher than expected.
Similarly HF b.p. 19 °C
Whereas: HBr –68 °C and HI –35 °C

Other effects of Hydrogen Bonding

Viscosity

1.Substance	diethyl ether	ethanol	water	glycerol
2.Molecular mass	74	46	18	92
3.Structural Formula				
4. No of –OH groups	0	1	2	3

Increasing Viscosity

Viscosity is not only related to molecular mass but also to Hydrogen bonding.
The –OH groups allow hydrogen bonding between the molecules and this increases the viscosity.

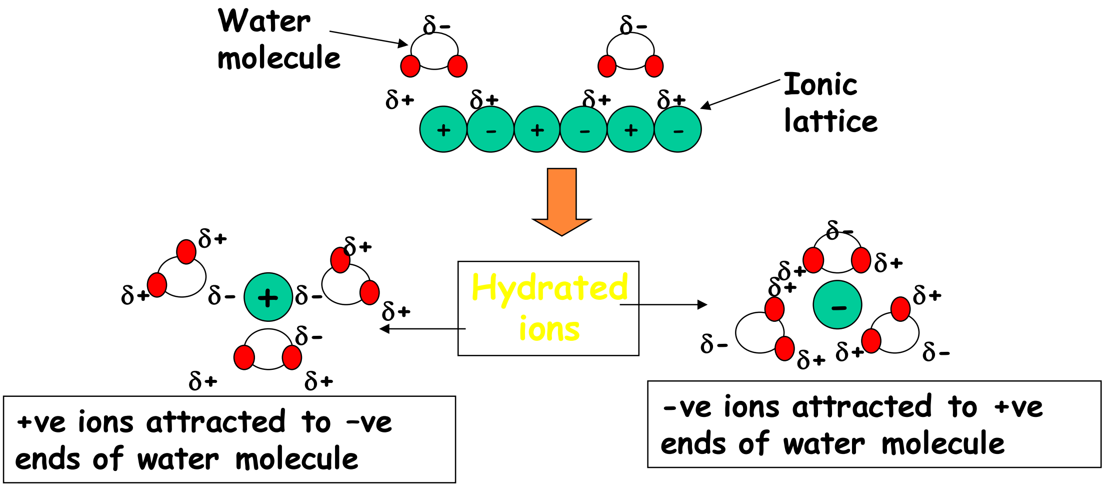
Miscibility

Miscible liquids mix thoroughly without any visible boundary between them, e.g. ethanol and water would be described as miscible but water and oil are **immiscible** as the oil forms a visible layer on water.
Hydrogen bonding aids miscibility (ethanol and water both contain hydrogen bonds).
<u>NB</u> very strongly polar liquids (without Hydrogen bonding) can also be miscible with water.

Bonding, Solubility and Solutions

Ionic lattices and **polar covalent molecular compounds** tend to be:
Soluble in **water** and other **polar solvents**, due to the attraction between the opposite charges.

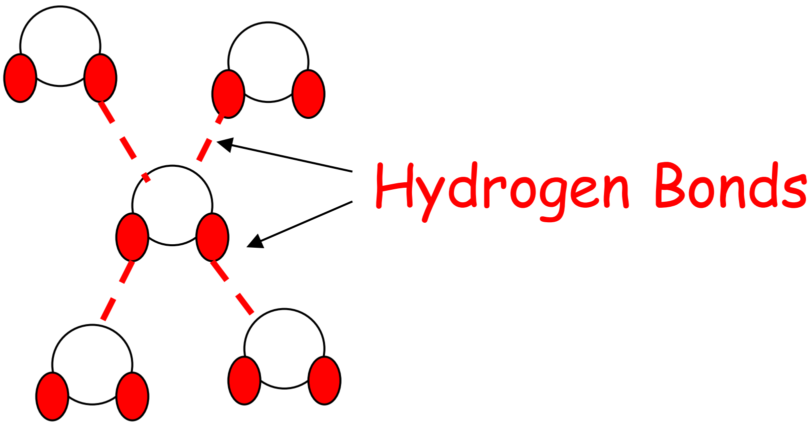
Insoluble in **non-polar solvents**, as there is no attraction between the ions and the solvent molecules.
e.g. when ionic compounds dissolve in water the lattice is broken up and the ions are surrounded by water molecules



Non-polar covalent molecular substances tend to be:
Soluble in **non-polar solvents** like carbon tetrachloride or hexane.
Insoluble in **water** and other **polar solvents** as there are no charged ends to be attracted.

'Like dissolves like'

Hydrogen Bonding and the Properties of Water

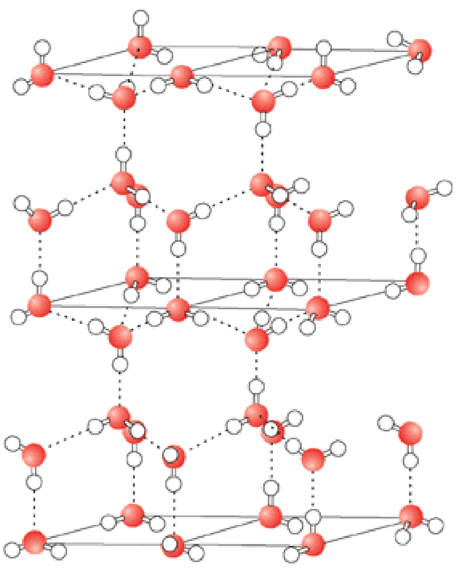


Each water molecule is surrounded by 4 hydrogen bonds
Water has a **high surface tension**. The molecules on the surface
have hydrogen bonds pulling the surface molecules closer together.

Why do pipes burst when water freezes and why does ice float on water?

As matter is cooled, it normally contracts and becomes more dense.

However, as water freezes it expands (at about 4°C) because the strong hydrogen bonds between the molecules force them into an open lattice structure.



This makes the solid ice less dense (takes more space) than the liquid so ice floats on water and pipes burst when water freezes.

Summary of Relative Bond Strengths

Bond Type	Strength (kJ mol $^{-1}$)	
Metallic	80 to 600	
Ionic	100 to 500	Intramolecular
Covalent	100 to 500	
Hydrogen	40	
Dipole-Dipole	30	Intermolecular
Van der Waals	1 to 20	

Bonding and Structure - Glossary

Word	Meaning
Bonding electrons	are shared pairs of electrons from both atoms forming the covalent bond.
Chemical bonding	is the term used to describe the mechanism by which atoms are held together.
Chemical structure	describes the way in which atoms, ions or molecules are arranged.
Covalent bond	a covalent bond is formed when two atoms share electrons in their outer shell to complete the filling of that shell.
Covalent radius	is half the distance between the nuclei of two bonded atoms of an element
Delocalised	Delocalised electrons, in metallic bonding, are free from attachment to any one metal ion and are shared amongst the entire structure.
Dipole	an atom or molecule in which a concentration of positive charges is separated from a concentration of negative charge.
Electronegativity	is a measure of the attraction that an atom involved in a bond has for the electrons of the bond
Fullerenes	are molecules of pure carbon constructed from 5- and 6-membered rings combined into hollow structures. The most stable contains 60 carbon atoms in a shape resembling a football.
Hydrogen bonds	are electrostatic forces of attraction between molecules containing a hydrogen atom bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen, and a highly electronegative atom on a neighbouring molecule.

Word	Meaning
Intermolecular forces	are those which attract molecules together. They are weaker than chemical bonds.
Intramolecular forces	are forces of attraction which exist within a molecule.
Ionisation energy	The first ionisation energy is the energy required to remove one mole of electrons from one mole of gaseous atoms (i.e. one electron from each atom). The second and subsequent ionisation energies refer to the energies required to remove further moles of electrons.
Isoelectronic	means having the same arrangement of electrons. For example, the noble gas neon, a sodium ion (Na^+) and a magnesium ion (Mg^{2+}) are isoelectronic.
Lattice	A lattice is a regular 3D arrangement of particles in space. The term is applied to metal ions in a solid, and to positive and negative ions in an ionic solid.
London Dispersion Forces	are the intermolecular forces of attraction which result from the electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules.
Lone pairs	are pairs of electrons in the outer shell of an atom which take no part in bonding.
Miscible	fluids are fluids which mix with or dissolve in each other in all proportions.
Polar covalent bond	a covalent bond between atoms of different electronegativity, which results in an uneven distribution of electrons and a partial charge along the bond.
Polar covalent bond	a covalent bond between atoms of different electronegativity, which results in an uneven distribution of electrons and a partial charge along the bond.

Word	**Meaning**
Van der Waals' forces	Is the general name given to all intermolecular attractions including London dispersion forces and hydrogen bonding.
Viscosity	is the resistance to flow that is exhibited by all liquids.