Compounds and Bonding

COVALENT NETWORK COMPOUNDS **2.1**
At the end of this unit you should be able to

1. State what is meant by a covalent network compound
2. Describe the bonding and structure in compounds like silicon carbide and silicon dioxide
3. Describe and explain the melting point and hardness of covalent network structures
4. Describe and explain some uses of covalent network substances which depend upon their structures
5. Explain how the formula of a covalent compound is related to its structure

Silicon carbide SiC

Similar to diamond with each carbon surrounded tetrahedrally with 4 silicon atoms and vice versa. All the bonds are covalent.

Stick picture of silicon carbide in here

* Silicon carbide can be described as a covalent network substance as the covalent bonds extend throughout the whole structure.
* The covalent bond between each carbon and silicon atom consists of a pair of shared electrons.
* Each atom achieves a stable octet of outer electrons.
* The structure can be described as a lattice as there is a regular repeating pattern of carbon and silicon atoms.
* SiC is the empirical formula foe silicon carbide, it gives the simplest ration of atoms present.
* SiC will have a high melting and boiling point as strong covalent bonds must be broken, this takes a lot of energy which is only available at high temperatures.

Silicon dioxide SiO_2

Stick a picture of silicon dioxide here.

* Silicon dioxide can be described as a covalent network structure as the covalent bonds extend throughout the whole structure.
* Each silicon atom is bonded tertrahedrally by four oxygen atoms whilst each oxygen atom is bonded to two silicon atoms.
* SiO_2 is the empirical formula giving the ratio of silicon to oxygen atoms.
* SiO_2 will have a high melting and boiling point as strong covalent bonds must be broken, this takes a lot of energy, which is only available at high temperatures.

Properties and uses of silicon carbide and silicon dioxide.

Properties	Carborumdum SiC	Silica SiO_2
Melting point	decomposes at 2700^{o}C	~ 1610^{o}C
Hardness	extremely hard	very hard
Reactivity	very unreactive	fairly unreactive

All these properties arise because of the strong bonding present, it takes a lot of energy to break the bonds.

Uses	Carborundum SiC	Silica SiO_2
	Furnace linings. Crucibles. Abrasives	Abrasives Building materials Glass

Both these materials‘ uses are a direct results of their properties i.e. hardness and unreactivity. Silica can be used in making glass, as it is transparent. i.e. allows light to pass through it.

IONIC COMPOUNDS **2.2**

At the end of this unit you should be able to

1. Explain the type and size of the charge on a simple ion

2. Explain why two element ionic compounds contain both a metal and a non-metal ion.

3. Describe what holds the ions together in an ionic bond

4. Describe the bonding in a simple ionic compound such as sodium chloride explain the use of network and lattice

5. Explain how the formula of an ionic lattice is related to its structure.

6. Describe how changing the size of ions affect the strength of the ionic bond and state that changing melting points can show this difference.

7. Name some molecular ions and explain their arrangements in terms of electron arrangements.

8. Describe the different types of bonding present in compounds containing molecular ions and explain how their solutions can conduct electricity.

Atomic ions

The simplest ions are atomic ions i.e. atoms that have gained or lost electrons. Atoms gain or lose electrons in order to obtain more stable electron arrangements usually ones the same as a noble gas. Noble gases have more stable electron arrangements because they have lower energy content than other electron arrangements.

Atoms are neutral since they have equal numbers of positive protons and negative electrons. Metals become positive ions when they lose electrons whilst non-metal atoms become negative by gaining electrons.

Element	Atom electron arrangement	Ion electron arrangement	Ion-electron equation
magnesium	2,8,2	2,8	$Mg \Rightarrow Mg^{2+} + 2e$ ***($Mg - 2e \Rightarrow Mg^{2+}$)***
potassium	2,8,8,1	2,8,8	$Na \Rightarrow Na^{+} + e$
aluminium	2,8,3	2,8	$Al \Rightarrow Al^{3+} + 3e$
chlorine	2,8,7	2,8,8	$^{1}/_{2}Cl_2 + e \Rightarrow Cl^-$
oxygen	2.6	2,8	$^{1}/_{2}O_2 + 2e \Rightarrow O^{2-}$
phosphorus	2,8,5	2,8,8	$P = 3e \Rightarrow P^{3-}$

An **ionic bond** is the electrostatic attraction between a positive and a negative ion

Picture of sodium chloride here

Each sodium ion Na^+ is surrounded by six (6) chloride ions and each chloride ion Cl^- by six (6) sodium ions.

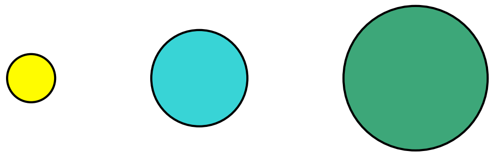
This gives a coordination of 6:6 and the ratio of 1:1

The formula is the simplest ratio of the ions present i.e. NaCl

Charge density

Density is the mass / volume, similarly charge density is the charge / volume.

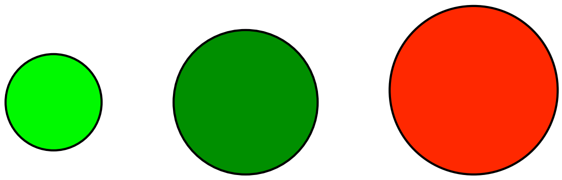
Singly charged ions eg Na^+ K^+ Cs^+



Charge density

$Na^+ > K^+ > Cs^+$

F^- Cl^- Br^-



Charge density

$F^- > Cl^- > Br^-$

The greater the charge density the greater the attraction the ion will have for ions of the opposite charge and the stronger the bond.

The strength of ionic bonds in a compound can be determined by looking at their melting points.

Compound	Melting point (^oC)	Compound	Melting point (^oC)
sodium fluoride	995	potassium fluoride	857
sodium chloride	801	potassium chloride	770
sodium bromide	750	potassium bromide	735

Doubly charged ions

Doubly charged ions would have higher charged densities and their ionic bonding should be stronger , hence they should have higher melting points

e.g. magnesium oxide M.Pt. 2852 ^oC

Molecular ions

Ionic compounds can be formed from more than two elements. Groups of atoms joined together by covalent bonds can also form ions e.g. sulphate ions

Picture of molecular ions here

Note Each atom gains a stable outer shell of electrons by sharing electrons, as would happen in 'normal' covalent bonds

The diagram below represents copper(II) sulphate

Picture of copper(II) sulphate here

The two main types of bonding present are covalent and ionic bonding.

Note *Only the ionic bonding is broken when this compound melts or dissolves in water.
*Copper(II) sulphate would conduct when molten or in solution as the ions would be free to move.
*Copper(II) sulphate is considered an ionic compound as it has the properties associated with ionic compounds rather than covalent ones.
i.e. It has a high melting point.
It is soluble in water.
It conducts when molten.

Molecular Compounds **2.3**

At the end of this unit you should be able to

1. Explain what is meant by a covalent molecular compound

2. Describe what holds the atoms and what holds the molecules together in covalent molecular compounds

3. Draw and use outer electron pictures to show how stable electron arrangements are achieved in covalent molecular compounds

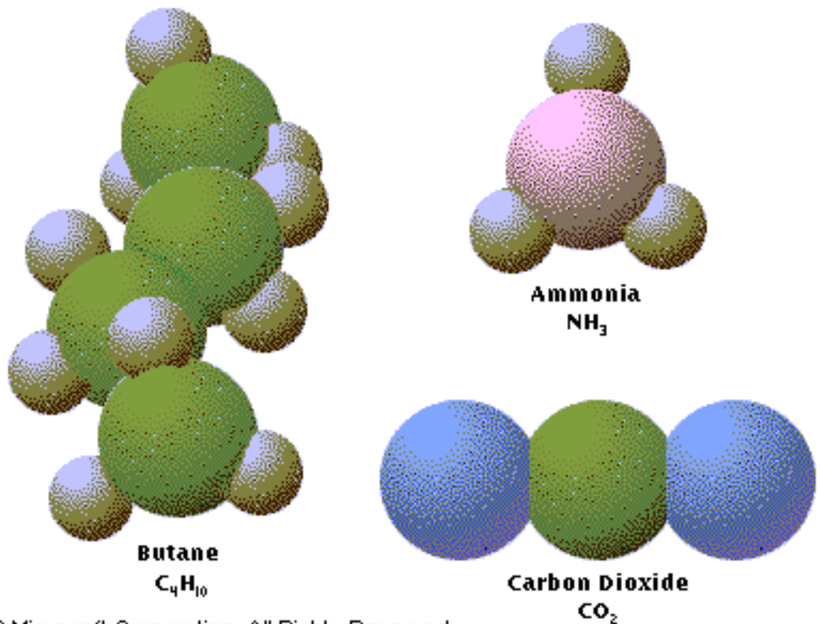
4. Explain in terms of Van der Waals forces why covalent molecular compounds have low melting points

5. Explain how the size of atoms in molecules affect the Van der Waals forces and hence the boiling points of compounds.

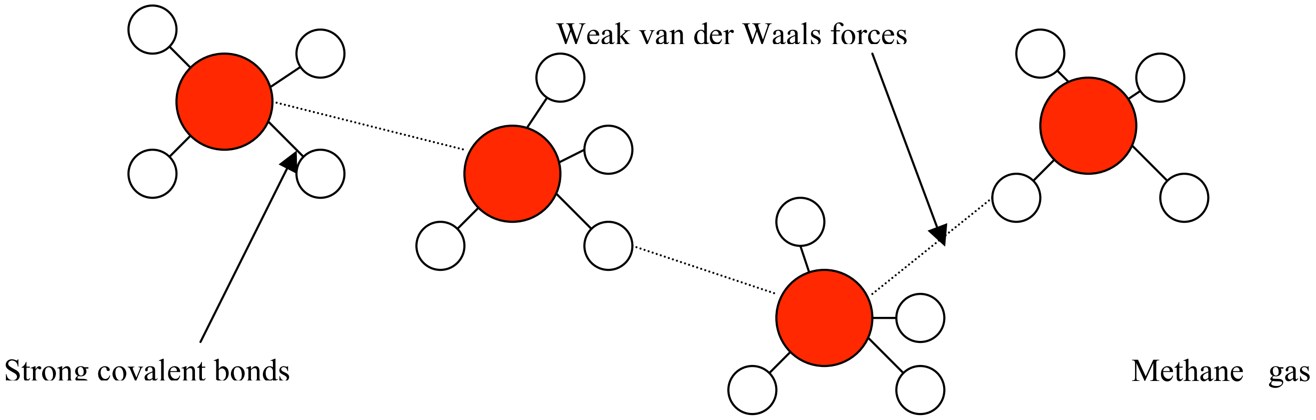
6. Explain why the number of atoms in a molecule affect the Van der Waals forces and hence the boiling points of compounds

Common Molecules

Molecules are compounds made up of specific combinations of atoms. Familiar substances may theoretically be divided into single molecules, as modeled here, but no further. Like a strict recipe in which atoms are the ingredients, each molecule has a chemical formula. If any ingredients are subtracted or changed, the molecule becomes something completely different.



Any compound which is a liquid or a gas at room temperature is always an example of a covalent molecular compound. All the compounds above are examples of covalent molecular compounds.



In the water molecules above there are strong covalent bonds formed by sharing two electrons, the covalent bonds hold the atoms together within each molecule, it takes a lot of energy to break these bonds, the type of energy needed is only available at high temperatures
e.g. > 1500^{o}C.
Note Covalent bonds are intramolecular forces i.e. within molecules

There are also Van der Waals forces formed by induced temporary dipoles holding the molecules together, the Van der waals forces are weak; it takes only a little energy to break them. The energy required is available at lower temperatures e.g. < 300^{o}C.
Note Van der Waals forces are intermolecular forces i.e. between molecules

Molecules are held together in a group by forces called van der Waals forces. These forces are usually weaker than those that hold the molecule itself together. The force between molecules depends on how far apart they are. When two molecules are widely separated, they attract each other. When they come very close together, they repel each other.

In a solid, the molecules are so arranged that the forces that attract and repel are balanced. The molecules vibrate about these positions of balance, but they do not move to different parts of the solid. As the temperature of a solid is raised, the molecules vibrate more strongly. When the Van der Waals forces can no longer hold the molecules in place, the solid melts and becomes a liquid.

In a liquid, the molecules move about easily, but they still have some force on one another. These forces are strong enough to form a film like surface on a liquid and prevent it from flying apart.

Note *A covalent molecular substance is a compound made from distinct molecules
*A single covalent bond consists of two shared electrons, the attractions between the shared electrons and the nuclei of each atom are what hold the atoms together.
*A double covalent bond as exists in carbon dioxide consists of four shared electrons.

Either dot/cross diagrams or structural formulae can represent covalent molecules

Name and molecular Formula	Dot/cross diagram	Structural formulae
Tetrafluoromethane/ carbon tetrafluoride **CCl_4**		
Carbon dioxide		
Phosphorus trichloride		
Hydrogen iodide		

Effect of increasing atomic size.

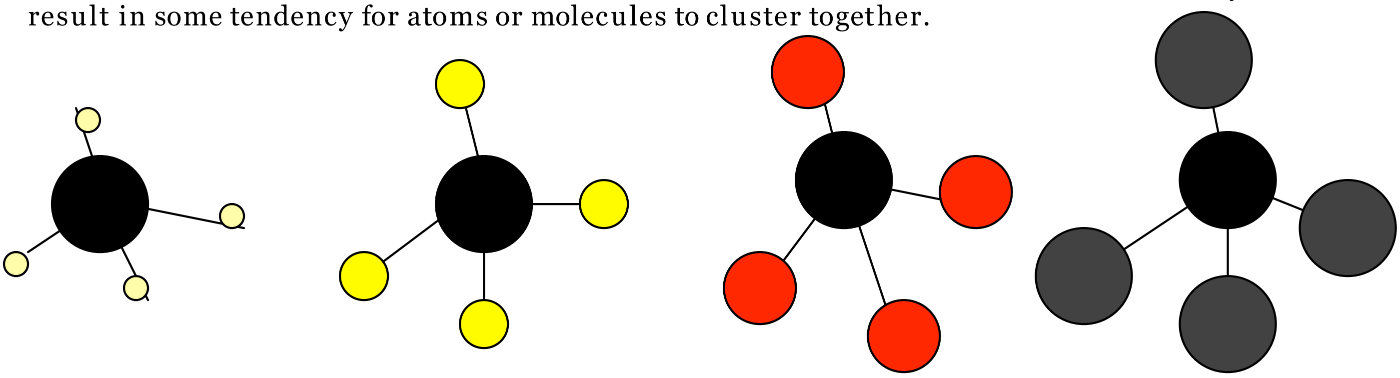
The B.Pt. and M.Pts. of covalent molecules depend upon the strength of the Van der Waals forces between the molecules.

Van der Waals Forces

In liquids and solids, a number of weakly attractive electrostatic forces, called van der Waals forces, exist between atoms and molecules. These intermolecular forces act over short distances, where molecules are packed closely together, and they are important in governing the physical properties of liquids and solids. Van der Waals forces consist of a number of different interactions, all of which rely on the existence or creation of dipoles.

A dipole in one molecule may also induce or accentuate a dipole in a neighbouring molecule, resulting in a force of attraction between them. A temporary fluctuation in the electron density of an atom or molecule creates a temporary dipole, which may then induce dipoles in other molecules.

Van der Waals forces are much weaker than ionic and covalent bonds. Nevertheless, they do result in some tendency for atoms or molecules to cluster together.



CF_4 M.Pt. = -184°C | CCl_4 M.Pt. = -23°C | CBr_4 M.Pt. = 90°C | CI_4 M.Pt. = 171°C

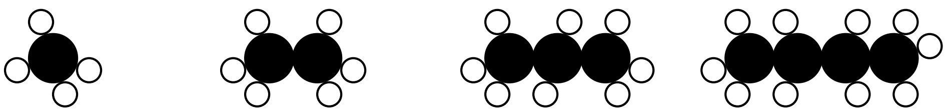
As the size of the molecule increases the amount of wobble increases in a similar way in, which a large jelly wobbles, more than a small one. As the wobble increases the strength of the Van der Waals forces increase.

The strength of the Van der Waals forces increase with increasing numbers of electrons.

The number of electrons in each molecule is the factor, which determines the strength of the Van der Waals forces. The number of electrons usually increases with

- Increasing size of molecule
- Increasing atomic and/or mass number
- Increasing number of atoms in molecule

These factors will increase van der Waals forces and give higher melting and boiling points.



methane B.Pt. = -164°C | ethane B.Pt. = -89°C | propane B.Pt. = -42°C | butane B.Pt. = -1°C

The boiling point increases from methane through butane because the number of electrons per molecule and hence the Van der Waals forces increase.

Polar Covalent Bonds **2.4**

At the end of this topic you should be able to

1. Explain why in molecules the electrons are not always equally shared

2. Explain the term's a) Pure covalent bond
b) Polar covalent bond
c) Permanent dipole

3. Draw a diagram to illustrate the charges resulting from a polar covalent bond

4. State what is meant by electronegativity and explain why it increases across a period and decreases down a group in the Periodic table

5. Determine the electronegativity of a covalent bond and explain how this can be used to work out the relative polarity of any covalent bond

6. Draw a diagram of a water molecule showing its shape and the polarity of each of its bonds

7. Determine whether molecules containing polar bonds are polar molecules.

8. Carry out and describe a test to show whether a liquid is polar or not.

Covalent bonding involves the sharing of pairs of electrons. When both atoms are the same the electrons will be shared equally as both atoms have the same amount of attraction for the shared pair(s) of electrons

This attraction for the shared pair of electrons is called electronegativity.

When different atoms form a covalent bond the electrons are not usually shared equally as the different atoms will have varying attractions for the shared pair(s) of electrons. When this occurs a polar covalent bond is formed.

Electronegativity

Electronegativity is the attraction that an atoms has for the bonded pair of electrons in a covalent bond.

The electronegativity of any atom (element) depends upon two factors

1. The number of protons in the nucleus of its atoms.
2. The size of the atom.

Number of protons

The greater the number of protons the greater the attraction for the bonded pair of electrons.

Hence the greater the electronegativity.

Size of atom

The smaller the atom the less distance between the nucleus and the bonded pair of electrons and the greater the attraction between the nucleus and the bonded electrons

Hence the greater the electronegativity.

How does this work out in relation to the elements?

H						
2.2						
Li	Be	B	C	N	O	F
1.0	*1.5*	*2.0*	*2.5*	*3.0*	*3.5*	*4.0*
Na	Ca	Al	Si	P	S	Cl
0.9	*1.2*	*1.5*	*1.9*	*2.2*	*2.5*	*3.0*

The attraction for bonded electrons increases across a period

The attraction for the bonded electrons dereases down a group.

Note The noble gases do not appear, as they do not usually form covalent bonds.

Attraction across a period

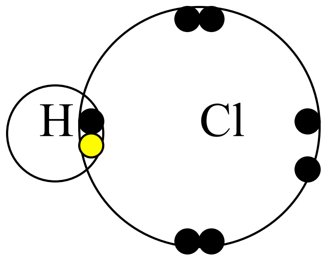
Going across a period the number of protons in the nucleus increases, the size of the atoms depend upon the number of electron shells, which is constant across any period. The electronegativity increases across any period in the Periodic Table.

The size of atom decreases across any period due to increased attraction from the nucleus.

Attraction down a group

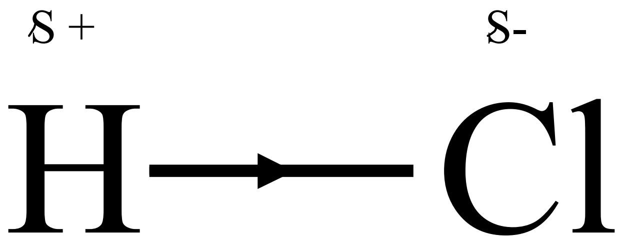
As we descend any group in the periodic table the number of electron shells increase, this makes the atoms larger, as it is the outermost electrons that are involved in the bonding, the distance between the bonded electrons and the nucleus increases, this decreases the attraction from the nucleus for the bonded pair of electrons. The elecronegativity decreases down any group.

Hydrogen chloride



H-Cl

The hydrogen has a slightly positive charge whilst the chorine has a slight negative charge



An arrow shown on the covalent bond shows the direction of polarity, pointing toward the negative end

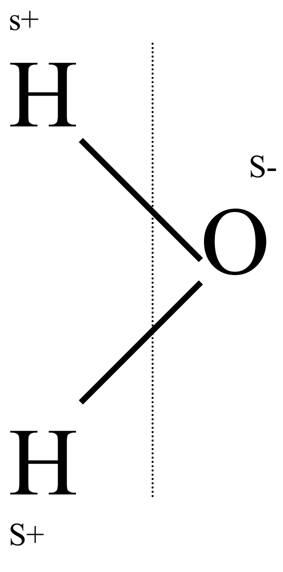
This results in a permanent dipole.

The strength of this permanent dipole can be measured by finding the electronegativity difference. The greater the difference the stronger the dipole.

Polar Molecules

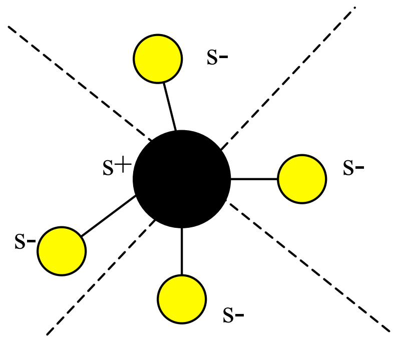
A polar molecule is one, which has a permanent slightly positive side and a permanent slightly negative side such as hydrogen chloride above.

Water



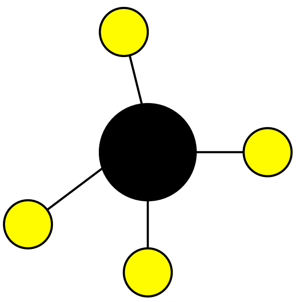
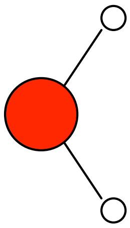
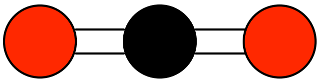
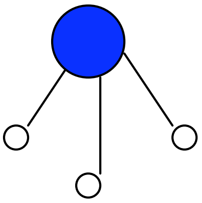
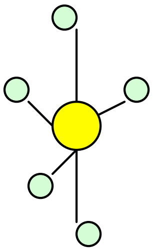
As can be seen a line can be drawn through the molecule to leave one side positive and one side negative.
Hence water has both polar bonds and is a polar molecule.

Tetrachloromethane



Note A line cannot be drawn through this molecule to leave a positive and a negative side.
This molecule has polar bonds.
However is not a polar molecule.

As you can see it is impossible to draw a straight line through a molecule of tetrachloromethane and have a negative side and a positive side. The molecule itself is non-polar although the bonds are polar.

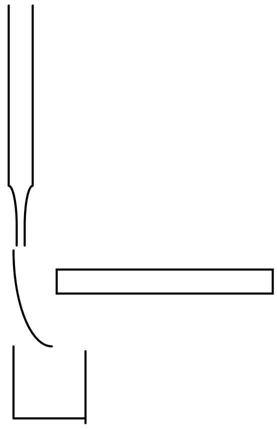
Molecule	polar bonds	polar molecule
 Tetrachloromethane	Yes	No
 Water	Yes	Yes
 Carbon dioxide		
 Ammonia		
 Pentafluorosulphur		

Detecting Polar Molecules

Liquids, which are polar, can be deflected using a charged rod. A fine jet of liquid is passed near a charged rod; if the jet is deflected then the liquid is polar.

The following liquids were tested using a fine jet from a burette

Liquid	Polar/ Non- polar
paraffin	
methanol	
propanone	
trichloromethane	



For each of the polar compounds get molecular models and construct the compounds to obtain their shapes, draw a line through each to show a positive and negative side.

Methanol

Propanone

Trichloromethane

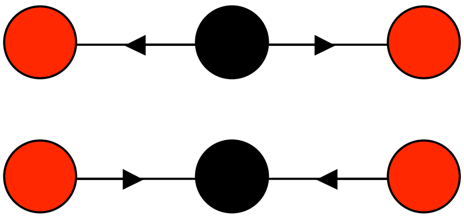
Polar – Polar Attractions 2.5

At the end of this topic you should be able to

1. Draw molecular diagrams for any molecule showing directions of the polar bonds
2. Decide whether a molecule with polar bonds will itself be polar or not
3. Explain what polar-polar attractions are, and compare their strength with Van der Waals forces
4. Explain the solubilities of polar and non-polar compounds in terms of energy changes
5. Explain the differences in boiling points of polar and non-polar molecules explaining the significance of comparing molecules with similar sizes

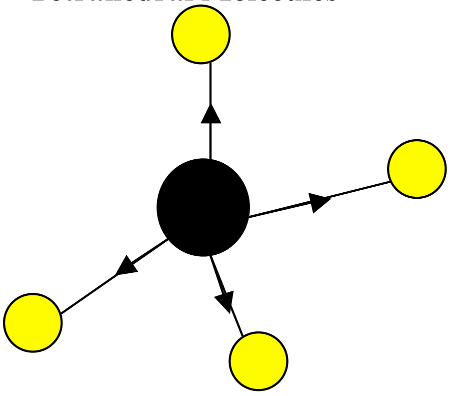
When polar bonds result in polar molecules, polar – polar attractions result. These are intermolecular forces, which affect the physical properties of compounds. The forces between molecules, which are polar, are stronger than the forces between molecules, which are not. Polar molecules have permanent dipoles whilst Van der waals forces are between temporary dipoles.

In completely symmetrical molecules the dipoles cancel each other out. The molecules are not polar and only Van der Waals forces exist between molecules

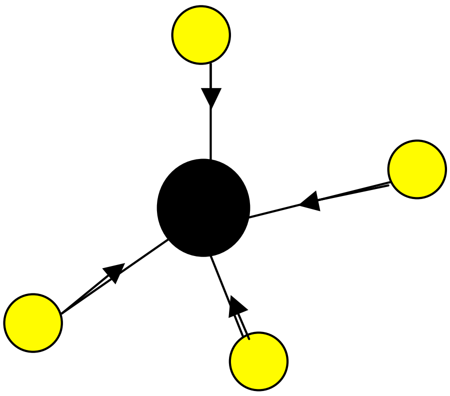


linear opposition

Tetrahedral Molecules



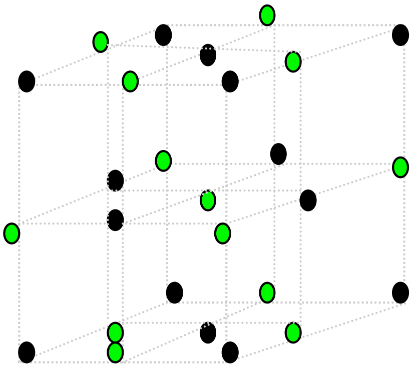
tetrahedral opposition



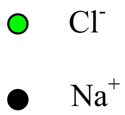
When dipoles are symmetrically opposed the molecule is non-polar.

Intermolecular Attractions and Solubility's

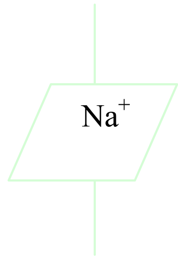
Water is a polar molecule and ionic substances tend to dissolve in polar substances



Cubic sodium chloride



When placed in water the water molecules pull the ions apart.



Solubility of Covalent Compounds

Intermolecular forces are important in determining whether a compound will dissolve in a solvent or not. If the intermolecular attractions are fairly similar then the two molecules will mix and move amongst each other easily. Liquids which dissolve in each other are said to be miscible. However if one substance has much stronger intermolecular forces then its molecules will tend to cluster together and form a separate layer. Liquids like this are said to be immiscible.

Experiment on Solubilities

Polar –polar attractions are weak forces and have little effect on solubilities between organic compounds.
The polar –polar attractions in water are much stronger than most and only the most polar molecules are miscible with water.

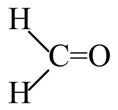
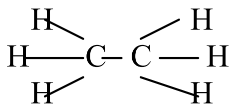
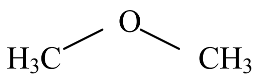
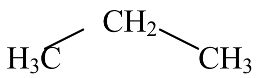
Energy changes when dissolving

The forces between strongly covalent molecules like water are fairly strong. If water were to dissolve non -polar substance energy would need to be added to break the polar - polar attractions between the water molecules. When water dissolves a strongly polar substance, more polar - polar attractions are formed and energy is released. As all substances try to be low in energy, water will not generally dissolve non - polar substances but will dissolve polar ones.

Melting and Boiling points of Molecular Compounds

When comparing molecules to determine whether the polar-polar attractions have any affect on the M.Pt and B.Pt. it is only fair to compare molecules of similar size, shape and number of electrons otherwise the differences could be due to the Van der Waals forces.

Compare the following

Name	Shape of molecule	Polar or non-polar	B.Pt.
Iodine chloride Bromine	I-Cl Br-Br		
Methanal Ethane	 		
Ether Propane	 		

Hydrogen Bonding 2.6

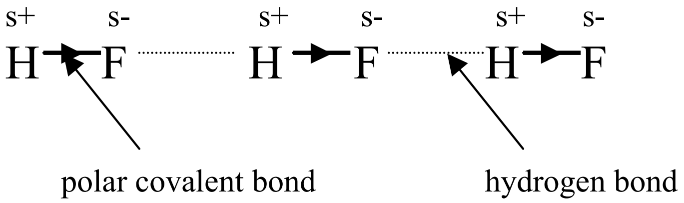
At the end of this topic you should be able to

1. *State that for hydrogen bonding to take place then a hydrogen atom must be joined to a nitrogen, oxygen or a fluorine atom*
2. *Compare the strengths of van der Waals forces with polar - polar interactions and also hydrogen bonding, being aware that all are weaker than metallic, ionic or covalent bonding*
3. *Draw molecular diagrams to represent hydrogen bonding*
4. *State that hydrogen bonding increases both melting and boiling points as the hydrogen bonding requires extra energy to break the permanent forces of attraction.*
5. *Describe an experiment to compare viscosities of liquids*
6. *Explain why hydrogen bonding increases the viscosity of substances*

All bonding forces are electrostatic in nature. Some are extremely weak like Van der Waals and polar-polar attractions; others are strong like ionic bonds, pure and polar covalent bonds.
An intermediate type of bonding exists this is HYDROGEN BONDING.

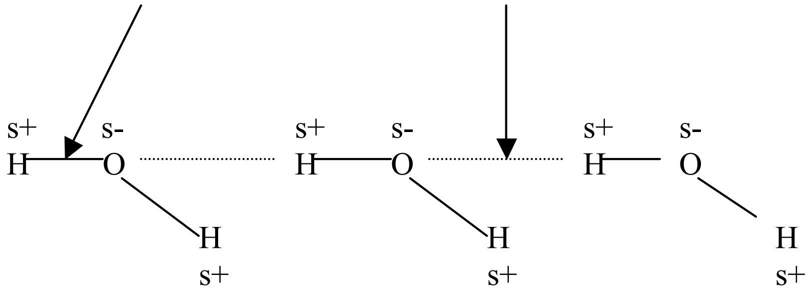
Hydrogen bonding only occurs when the atoms involved have a small highly electronegative atom joined to hydrogen. Only three elements have atoms that fulfil this criterion, these are fluorine, oxygen and nitrogen.

Hydrogen fluoride

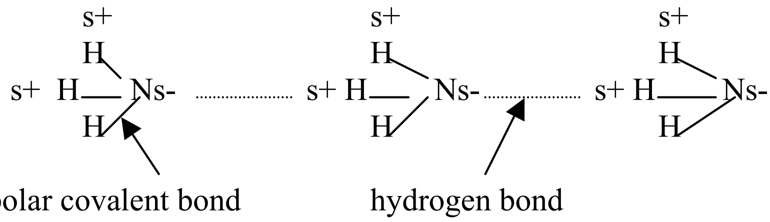


Fluorine , oxygen and nitrogen are small,l highly electronegative atoms. Only these atoms will form hydrogen bonds Other atoms with high electronegativity such as chlorine although forming polar bonds do not form hydrogen bonds as their atoms are too large.

Water



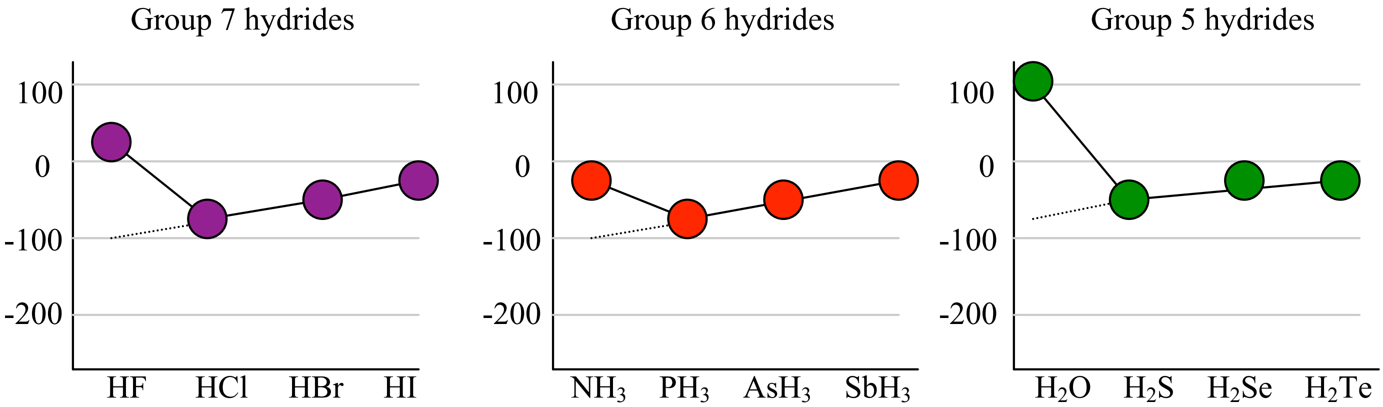
Ammonia



Effects of Hydrogen Bonding

Hydrogen bonding raises the M.Pt and B.Pt of any compound containing H-F , -O-H or –N-H bonds.

Boiling points of Group 7, 6 and 5 hydrides(^oC)

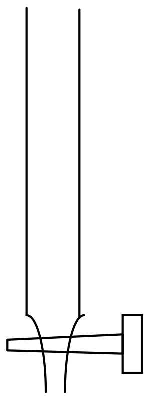


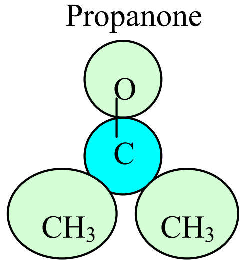
Compound whose molecules experience hydrogen bonding have higher M.Pts. and B.Pts. as more energy is need to break the stronger attractions between the molecules.

Note The broken lines show what the B.Pts would be if no hydrogen bonds were present.

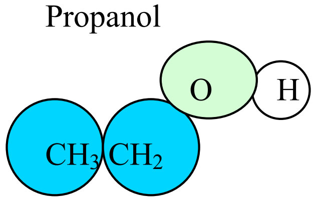
The viscosity of a liquid is a measure of how the molecules stick together. The more viscous the liquid the less runny it is.

The viscosity can be compared by running equal volumes through a burette.
We are comparing propanone and propanol as both have similar number of electrons and will have similar Van der Waals forces





No hydrogen bonding present
Time to pass through burette =



Hydrogen bonding present
Time to pass through burette =

Water 2.7

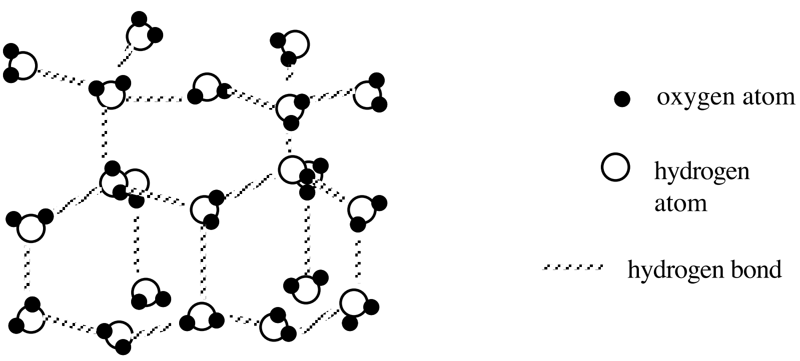
At the end of this unit you should be able to
1. Describe and explain why water has a higher boiling point than the other Group 6 elements
2. Describe and explain the arrangement of water molecules in ice
3. Explain in terms of its structure why ice is less dense than water
4. Explain the types of compounds likely to soluble in water.
5. Describe an experiment to compare viscosities of molecular liquids.

Water is such a commonplace substance that everyboby tends to take it for granted. In chemistry, however, the properties of water are more the exception than the rule. In many ways water is a very unusual compound.

Boiling point and melting point.

As we have already seen the boiling point and melting point is high owing to hydrogen bonding.
Structure of ice

Ice floats on water. This is unusual as most solids are more dense that their liquids and would sink.



If you look carefully you will see that each water molecule is surrounded by four hydrogen bonds
This arrangement not only makes the structure strong but also spaces out the water molecules and so prevents them from packing closely together

Simplified view of water molecule in ice.

Water is made up from small polar molecules, the main intermolecular forces are hydrogen bonds. When a substance dissolves in water it must produce energy to break any forces between the particles of the substance and also the hydrogen bonds between the water molecules.

e.g. Dissolving sodium chloride

Bonds / Forces broken	Bonds/Forces formed
Ionic bonds in sodium chloride	Hydration energy for Na^+
Hydrogen bonds in water	Hydration energy for Cl^-