

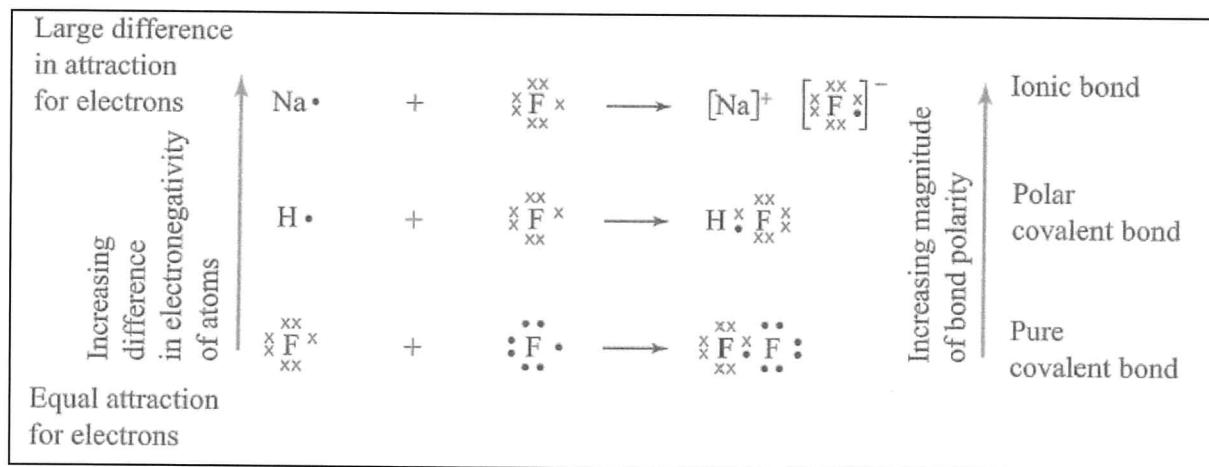
Polar Bonds

The Continuum : Ionic ► Covalent Bonds

- In an attempt to simplify, some books may seem to suggest that covalent and ionic bonds are two separate and completely different types of attachment. A covalent bond is a **shared pair of electrons**. The bond between the two atoms of any diatomic gas, such as chlorine gas, Cl_2 , is certainly equally shared. The two chlorine atoms have exactly the same pull on the pair of electrons, so the bond must be exactly equally shared. In cesium fluoride the cesium atom certainly donates an electron and the fluoride atom certainly craves an electron. Both the cesium ion (Cs^+) and the fluoride (F^-) ion can exist in solution independently of the other. The bond between a cesium and a fluoride ion to make cesium fluoride (CsF) would be clearly ionic because the difference in electronegativities (ΔEN) is so large.
- The amount of pull on an atom has on a shared pair of electrons, called **electronegativity**, is what determines the type of bond between atoms. Considering the Periodic Table without the inert gases, electronegativity is greatest in the upper right of the Periodic Table and lowest at the bottom left. The bond in francium fluoride should be the most ionic. Some texts refer to a bond that is between covalent and ionic called a **polar covalent bond**. There is a range of bonds between **purely ionic** and **purely covalent** that depends upon the electronegativity difference between atoms that are bonded. If there is a large difference in electronegativity, the bond has more ionic character and greater polarity. If the electronegativity of the atoms is more similar, the bond has more covalent character and less polarity.

Electronegativities of the Main Group elements

H						
2.2						
Li	Be	B	C	N	O	F
1.0	1.6	2.0	2.6	3.0	3.4	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.3	1.6	1.9	2.2	2.6	3.2
K	Ca	Ga	Ge	As	Se	Br
0.8	1.0	1.8	2.0	2.2	2.6	3.0
Rb	Sr	In	Sn	Sb	Te	I
0.8	1.0	1.8	2.0	2.0	2.1	2.7
Cs	Ba	Ti	Pb	Bi	Po	At
0.8	0.9	2.0	2.3	2.0	2.0	2.2
Fr	Ra					
0.7	0.9					



Polar Molecules

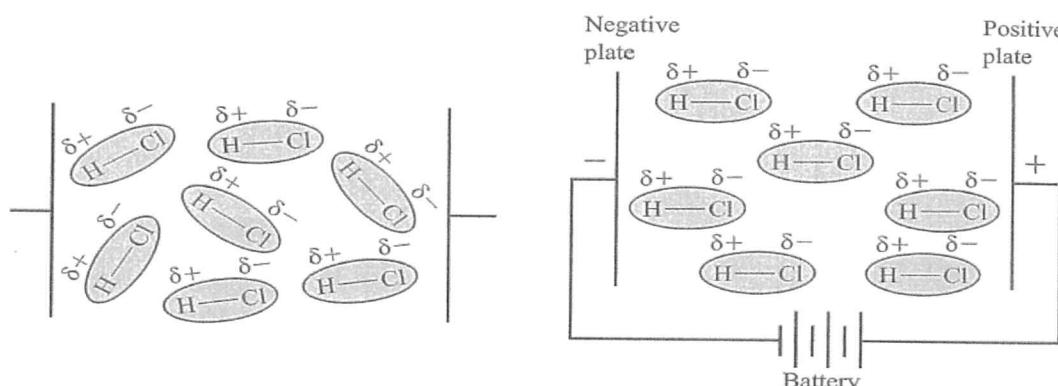
Molecules in which there is an uneven charge distribution are called polar molecules. As a result of the unequal sharing of electrons in any bond, one atom acquires a small positive charge, represented by the Greek letter delta, ($\delta+$) and the other acquires an equal small negative charge ($\delta-$).

In the HCl molecule, for example, the Cl atom is more electronegative (3.2) than the hydrogen atom (2.2). As a result the bonding electrons spend more of their time in the vicinity of the chlorine atom than the hydrogen atom. The result is that the chlorine end of the molecule acquires a small negative charge ($\delta-$) and the hydrogen end acquires a small positive charge ($\delta+$). Thus hydrogen chloride is a polar molecule and may be represented as follows.



The occurrence of a positive charge and an equal negative charge separated by some fixed distance is called an electric dipole.

If a polar molecule such as HCl is placed in an electric field it tends to align itself with the field, as shown in Figure 1. Note that the positive end of the molecule is oriented towards the negative plate and the negative end is oriented towards the positive plate.

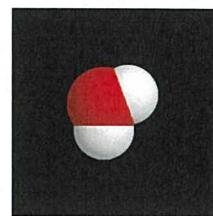


(a) Polar molecules are randomly oriented in the absence of an electric field.

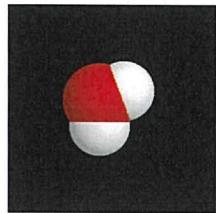
(b) In an electric field polar molecules line up as shown. Since the molecule has a net charge of zero, it does not move in the field.

INTERMOLECULAR FORCES

- As the title suggests, intermolecular forces are **FORCES OF ATTRACTION BETWEEN** like or different MOLECULAR species. They are **not physical bonds** as molecular species form DISCRETE molecules. The species may be monatomic atoms such as the inert gases (eg: Ne) or diatomic elemental molecules (eg: H₂) or even molecular compounds (eg: CO₂, H₂O).
- The extent of these forces depends on the degree of polarity that the molecule possesses as a result of **ELECTRONEGATIVITY DIFFERENCES** (differences in the ability to attract electrons in a bonding situation) between atoms that are COVALENTLY bonded IN the molecule.
- In order to determine the molecules overall polarity the *electronegativity differences* in each bond in the molecule must be considered. The bigger the difference then the bigger will be the DIPOLE in the bond. Dipoles are created as a result of differences in **electron field density** in the vicinity of the atoms involved in the bond. The sharing of electrons is not even due to the differences in the electronegativity of each atom in the bond.
- When the extent of the bond dipoles is determined then the **GEOMETRY** of the molecule must be considered to ensure that there is no VECTOR CANCELLATION of the electric field effects of each dipole. If dipoles are of equal strength and are orientated in opposite directions then they will cancel to produce NO OVERALL polarity. Linear, triangular planar and tetrahedral geometry will frequently produce zero overall polarity despite the presence of 2, 3 or even 4 bond dipoles.
- If a resultant polarity or **OVERALL MOLECULAR DIPOLE** exists then the relative strength can be determined by the degree of electronegativity difference in the various covalent bonds IN the molecule.
- The trend in electronegativity in the periodic table is that it increases towards the right and falls as you go down each group. This means that the metal elements on the left will have the lowest values while FLUORINE, OXYGEN and NITROGEN (in that order) will be the elements with the highest values.
- A bond formed between a metal and a non-metal will have so large a difference that the bond cannot be considered to have **covalent character**; rather the electrons are *taken* by the non-metal and *lost* by the metal in a typical REDOX reaction. The bond formed in this way is an IONIC bond and the very strong ELECTROSTATIC force of attraction between the ions gives the compound its characteristically high melting and boiling points. We are not dealing with intermolecular force in this situation as the compound is not molecular.

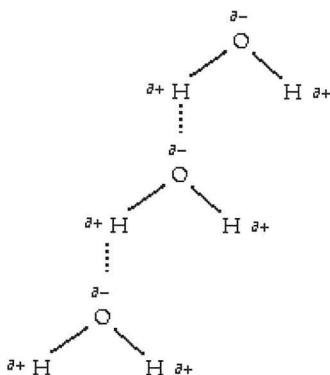


- Intermolecular forces arising from overall polarity do not usually rank in comparison to the extremely strong electrostatic forces in an ionic bond. There will however be greater basis for attraction if the polarity is high. The greatest electronegativity differences in covalent bonds are usually seen when the bond is between a Hydrogen atom and an atom of one of the three most electronegative elements (N,O,F). This is because hydrogen has a relatively low electronegativity for a non-metal.

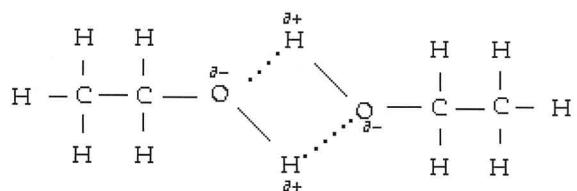


- If a molecule contains at least one of these types of covalent bond (hydrogen bonded to N, O or F) then it will attract other molecules with a relatively large force of attraction known as **HYDROGEN BONDING FORCE**. The other molecules in question may be similar or of a different type, as long as they also contain one such bond somewhere in their structure. Hydrogen bonds are NOT a physical bond, rather they are a **FORCE** of attraction that must be **OVERCOME** (not broken) by heat when melting or boiling the substance. For small molecules this type of **VAN DER WAALS INTERACTION** is considered the strongest type.

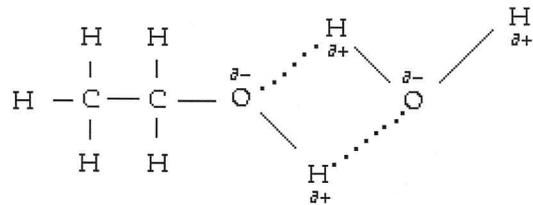
Hydrogen Bonding Between Water Molecules



Hydrogen Bonding Between Ethanol Molecules

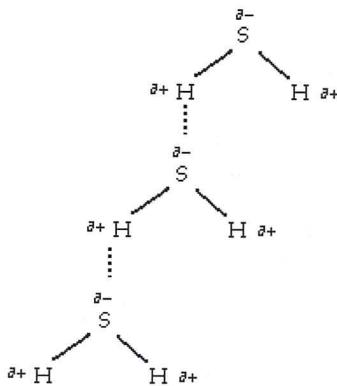
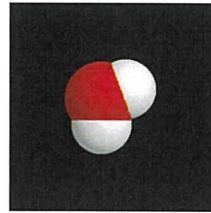


Hydrogen Bonding Between Water Molecules and Ethanol Molecules



- Dipoles may be created in a covalent bond between any non-metal atoms that do not have identical electronegativities. When the electronegativity difference is less than that characteristic of hydrogen bonded molecules the dipoles and hence overall molecular polarities are less. The forces that arise between these molecules are of exactly the **SAME NATURE** but are less extreme and are termed **DIPOLE-DIPOLE FORCES**.

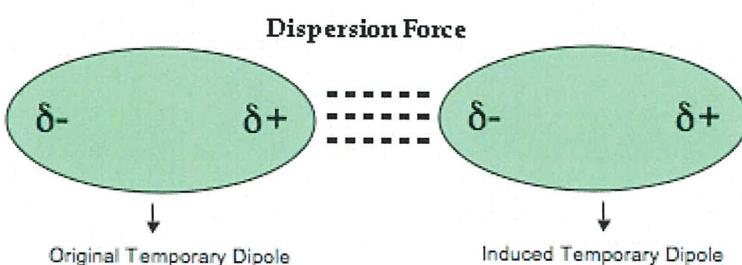
Dipole-Dipole Forces Between Hydrogen Sulphide Molecules



- It is the degree of intermolecular force that determines the melting point and boiling point of molecular compounds and ultimately their phase at room temperature. A molecule that is a solid at room temperature must have larger I.M.F. than one that is a liquid or a gas. The effects of intermolecular force are most pronounced at cold temperatures and are responsible for the regular crystal lattice structures that most compounds form in their solid states. This occurs as molecules freeze and orientate themselves so that the dipoles in their bonds are alternating.
- There are still intermolecular forces operating even when no dipole exists in the molecule as a result of vector cancellation or indeed lack of electronegativity difference in the bonds. Such molecules are said to be **NON-POLAR**. These forces are generally weaker in small molecules and are termed **DISPERSION FORCES** or **London Forces**. The size of this type of force depends **indirectly** on the mass of the molecule. The reality is that it is actually dependant on the *number of electrons* and protons in the atoms of the molecule.
- The term **DISPERSION FORCE** arises as the electrons in any atom, whether bonded or not, can never be exactly symmetrically arranged about their nuclei. Electron orbit has a large degree of randomness which means that at any moment in time the electrons will be asymmetrically arranged to a degree and that small differences in electron field density **must** exist. This leads to small but measurable temporary and ever changing dipoles within the electron clouds of all atoms and even in atoms that are part of bigger molecules. The more electrons an atom has and the more atoms that comprise a molecule then the greater will be the tendency for these dipoles to exist and interact. The greater the dispersion effects the greater will be the intermolecular force between particles.



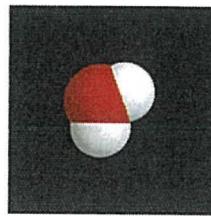
- When one atom or molecule approaches another there may be a mutual influence on each other's electron fields that leads to **INDUCED** dipoles in neighbouring particles. The dipoles interact to lead to a basis of attraction that is dependent on the particles electron amounts – the bigger the atoms or molecules involved the larger the intermolecular force that arise.



- The general order of importance of the Van der Waal's forces when comparing the same molecule if it is relatively small is:

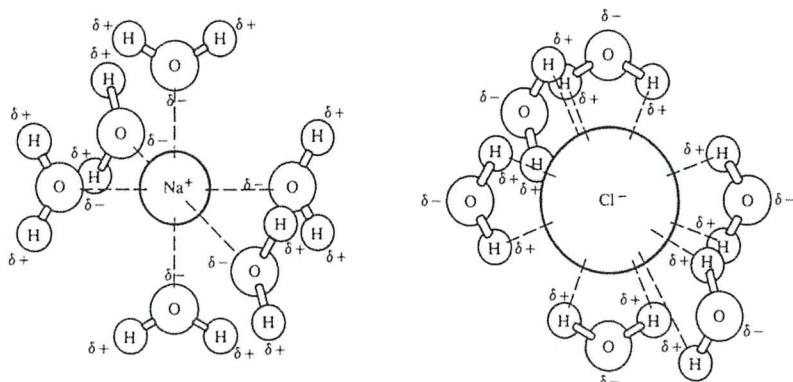
HYDROGEN BONDING >> DIPOLE - DIPOLE FORCES >> DISPERSION FORCES

- When comparing molecules of different sizes this order frequently does not hold. Large molecules that possess no overall polarity frequently have greater I.M.F. than smaller molecules that do, as the dispersion force rises with their mass. Some plastics or polymers may have molecules with 30,000 carbon atoms in a row. These massive molecules have very large dispersion forces and not only are solid at room temperature but may be "bullet -proof".



- The type of intermolecular force that a molecular substance has is a prime determining factor in whether or not it will mix with or **DISSOLVE** in another. Solubility is greatest between POLAR substances and other POLAR substances and NON-POLAR substances with other NON-POLAR substances. NON-POLAR substances will not mix with POLAR substances and vice versa.

RULE : " LIKE DISSOLVES LIKE"



- Many of the properties of molecular liquids may be explained in terms of the relative intermolecular forces in action. VISCOSITY (Rate of Flow) is **lower** in liquids that have higher I.M.F. the liquids will be **more viscous**; will not flow as well, as these forces rise.

Eg: The large Dispersion Forces **between** the molecules in oils prevent them from flowing as well as water even though this small molecule has Hydrogen Bonding Forces operative.

- The surface effect of SURFACE TENSION will be larger for liquids as the intermolecular forces rise leading to greater COHESION between the molecules as they try to form a spherical shape.
- The ease with which a liquid evaporates will also be determined by its' intermolecular forces. Liquids that boil at low temperatures and evaporate easily due to their low I.M.F are said to be VOLATILE. Many volatile liquids may be smelled as a result of evaporating particles interacting with our nasal membranes.

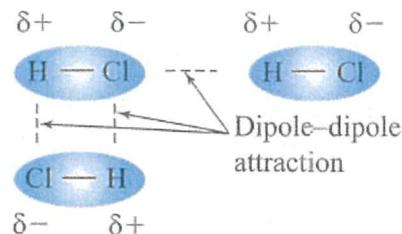
Intermolecular Forces

The melting and boiling points of covalent molecular substances are much lower than those of ionic and metallic substances. This implies that the forces of attraction between the molecules in a molecular crystal are relatively weak. The weak forces between molecules in molecular solids and liquids are called van der Waals forces after the Dutch scientist who first attempted to explain them. Although all intermolecular forces are referred to as van der Waals forces they can be further classified as dipole–dipole forces, dispersion forces and hydrogen bonds.

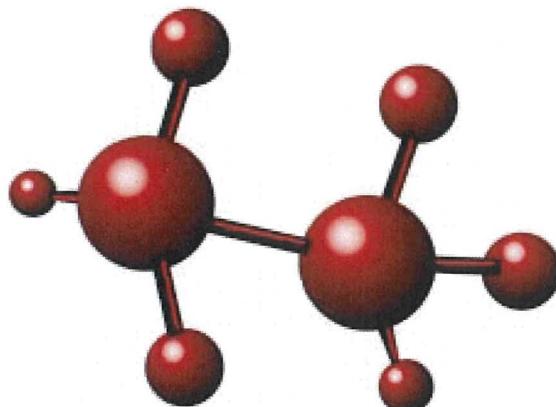
Dipole –Dipole Forces

If two polar molecules such as HCl approach they will tend to orient themselves in such a way that the positive end of one molecule is close to the negative end of another molecule, as shown in Figure . This orientation leads to a lower potential energy.

The force of attraction between the oppositely charged ends of neighbouring polar molecules is called dipole–dipole attraction. The ordered structure illustrated in Figure is most pronounced in the solid state. In the liquid state the molecules are not as ordered due to their higher kinetic energy.



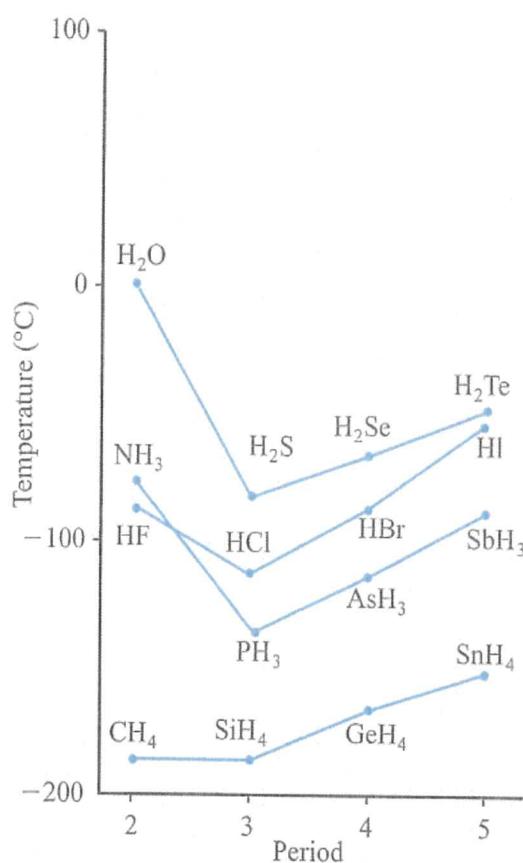
The charged ends of a dipole have a relatively small charge compared with the charges of ions in ionic substances. Consequently, dipole–dipole attractive forces are relatively weak compared with the attractive forces between oppositely charged ions. Thus solid HCl, a polar molecular crystal, melts at $-114\text{ }^{\circ}\text{C}$ whereas NaCl, an ionic crystal, melts at $801\text{ }^{\circ}\text{C}$.



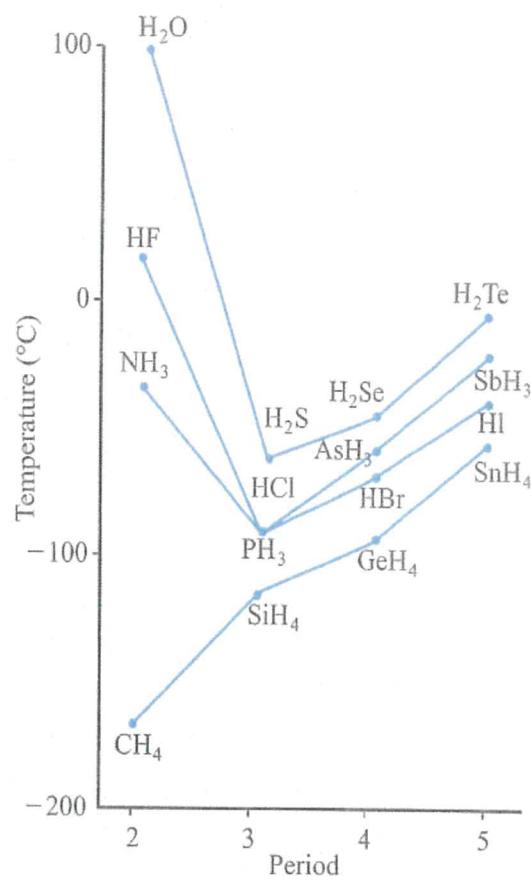
Hydrogen Bonding Forces

The graphs in Figure show the melting points and boiling points of the hydrides of the group IV, V, VI and VII elements. There is a general increase in melting and boiling points with increasing molecular mass for the hydrides of the group IV elements.

The regular increase in melting and boiling points also occurs for the compounds H_2S through H_2Se to H_2Te , but H_2O has a higher melting and boiling point than expected. A similar pattern occurs with the hydrides of groups V and VII, where



(a) Melting points

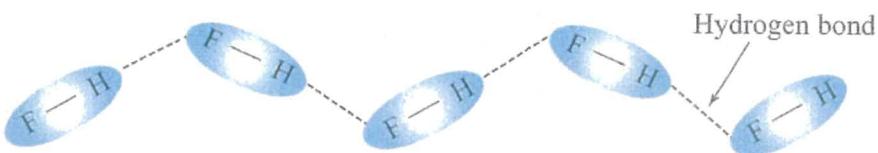


(b) Boiling points

NH_3 and HF have higher melting and boiling points than the trends would suggest. The higher melting and boiling points for NH_3 , H_2O and HF indicate that there must be unusually strong intermolecular forces between the molecules of these substances.

HF , H_2O and NH_3 are all highly polar since they contain the three most electronegative elements, fluorine, oxygen and nitrogen, bonded to hydrogen which has a relatively low electronegativity. This results in very polar molecules and much stronger intermolecular forces than expected for dipole–dipole interactions. These particularly strong intermolecular forces are referred to as hydrogen bonds.

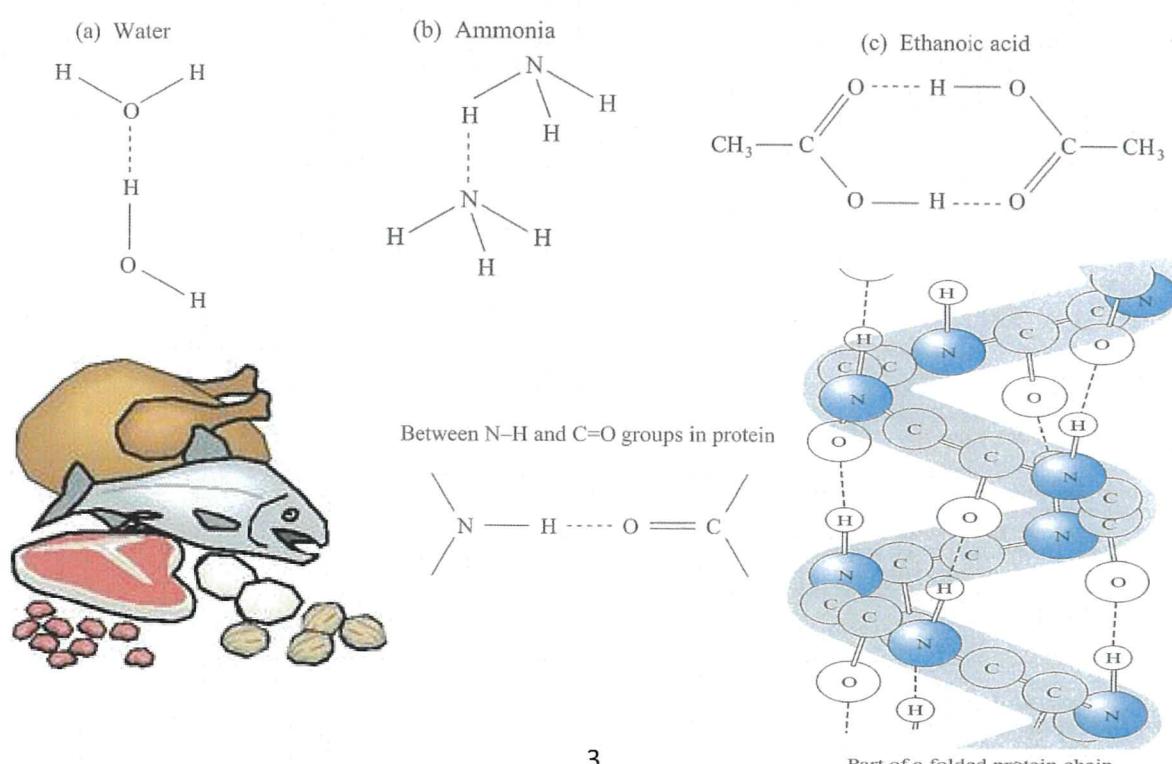
Hydrogen bonding is a special form of dipole–dipole attraction. It is found in systems where a hydrogen atom is bonded to an atom of oxygen, nitrogen or fluorine. When hydrogen is bonded to the highly electronegative fluorine, oxygen or nitrogen atoms, the resulting bond is very polar and the shared electrons are strongly attracted towards the more electronegative atom. Since the hydrogen atom in these circumstances has an appreciable partial positive charge, it experiences a strong attractive force with lone pairs of electrons on oxygen, nitrogen or fluorine atoms of nearby molecules. Hydrogen bonding is particularly strong in solid HF which consists of long chains of HF molecules, shown in Figure 7.16.



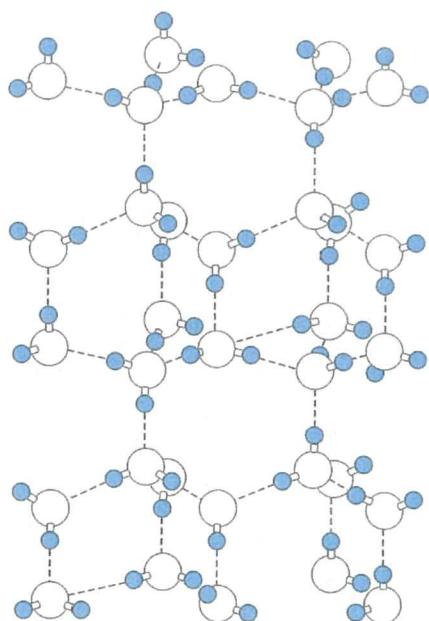
The essential requirements for hydrogen bonding are as follows.

- 1 A hydrogen atom bonded to N, O or F so that the hydrogen atom has an appreciable partial positive charge due to the unequal sharing of the pair of electrons in the covalent bond.
- 2 An unshared pair of electrons on a neighbouring N, O or F atom which can attract the partially positive hydrogen atom.

The strength of hydrogen bonds are in general about ten times those of dipole–dipole forces but about one-tenth those of ionic or covalent bonds. Hydrogen bonding is important in many chemical systems. Some examples are included in Figure 7.17. These include the formation of hydrogen bonds between simple molecules such as H_2O , NH_3 and CH_3COOH . Figure 7.17 also illustrates the formation of hydrogen bonds between N–H and C=O groups. Hydrogen bonds such as these are extremely important in biological systems and play a critical role in determining the structure of proteins and the double helix of DNA.



The presence of hydrogen bonding accounts for many of the unique properties of water. For example, the arrangement of water molecules in ice creates a very open structure which causes the density of ice to be less than that of liquid water. Figure 7.18 illustrates the arrangement of water molecules in ice. When ice melts, the regular lattice breaks up and the water molecules can pack more closely to form a liquid of somewhat higher density. Without hydrogen bonding, ice would sink to the bottom of oceans and lakes, a process that in cold climates would cause the death of fish and other aquatic life.



Hydrogen bonding



London dispersion forces



Ion-dipole forces



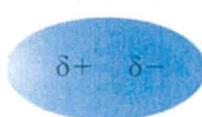
Dipole-dipole forces



Dispersion Forces

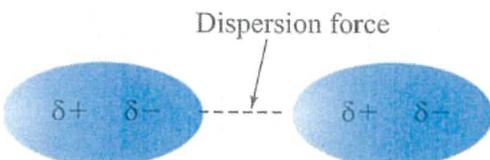
If there were no intermolecular forces between the molecules of some substances then these substances would exist only as gases, because there would be no forces to attract the molecules together. However, every substance can be condensed to a liquid and finally to a solid if it is cooled to a low enough temperature. Thus some form of intermolecular force must exist in every substance, even when the molecules consist of identical atoms as in O_2 or N_2 and do not possess a permanent dipole.

Intermolecular forces between non-polar molecules arise from the fact that at any instant a molecule may have a temporary dipole. The temporary dipole occurs at any moment when the electrons in an atom or molecule are not symmetrically distributed. This means that the electron charge cloud will not be symmetrical, as indicated in Figure



Temporary Dipole due to unsymmetrical electron charge distribution

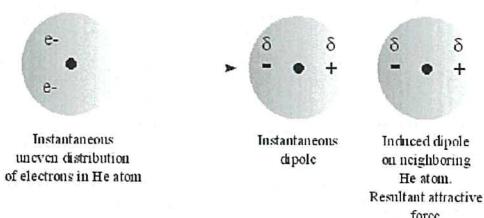
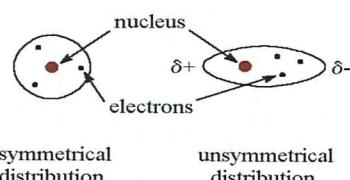
When an atom or molecule with a temporary dipole is near another atom it influences the charge distribution in the second atom. The temporary dipole is said to induce a dipole in the neighbouring atom. The two species then attract each other as shown in Figure



Attraction between two species with temporary dipoles.

The weak forces of attraction between atoms or molecules arising from such temporary dipoles are known as dispersion forces. Dispersion forces are the weakest kind of intermolecular forces. It should be noted that any particular temporary dipole exists only for an instant of time because the distribution of electrons changes constantly. Elements that form molecular crystals held together by dispersion forces include hydrogen and the non-metals on the right-hand side of the periodic table

Dispersion forces between molecules increase in strength with the number of electrons present in the molecules. The number of electrons is greater in larger molecules so that dispersion forces are also related to molecular size and mass. This is illustrated by examining the melting point trends of the noble gases and halogens shown in



Melting Points of the Noble Gases and Halogens

Noble gas	Atomic mass	Number of electrons	Melting point (°C)	Halogen	Molecular mass	Number of electrons	Melting point (°C)
He	4	2	-272	F_2	38	18	-220
Ne	20	10	-249	Cl_2	71	34	-101
Ar	40	18	-189	Br_2	160	70	-7
Kr	84	36	-157	I_2	254	106	114
Xe	131	54	-112				

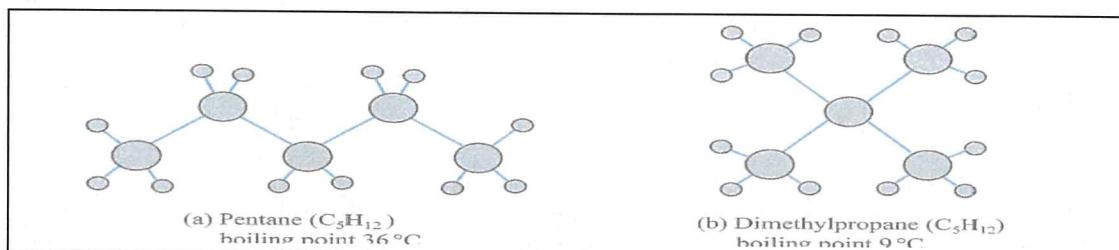
It should be noted that dispersion forces contribute to the attractive forces between all molecules. However, in the case of highly polar molecules, the dipole-dipole forces are much stronger and the effect of the dispersion forces is not as significant. Between molecules that are non-polar or only slightly polar, dispersion forces are the main intermolecular forces acting. Consider the melting points of the compounds HCl, HBr, and HI shown

Melting points of HCl, HBr and HI

Substance	Molecular mass	Melting point (°C)
HCl	36.5	-114
HBr	81	-87
HI	128	-51

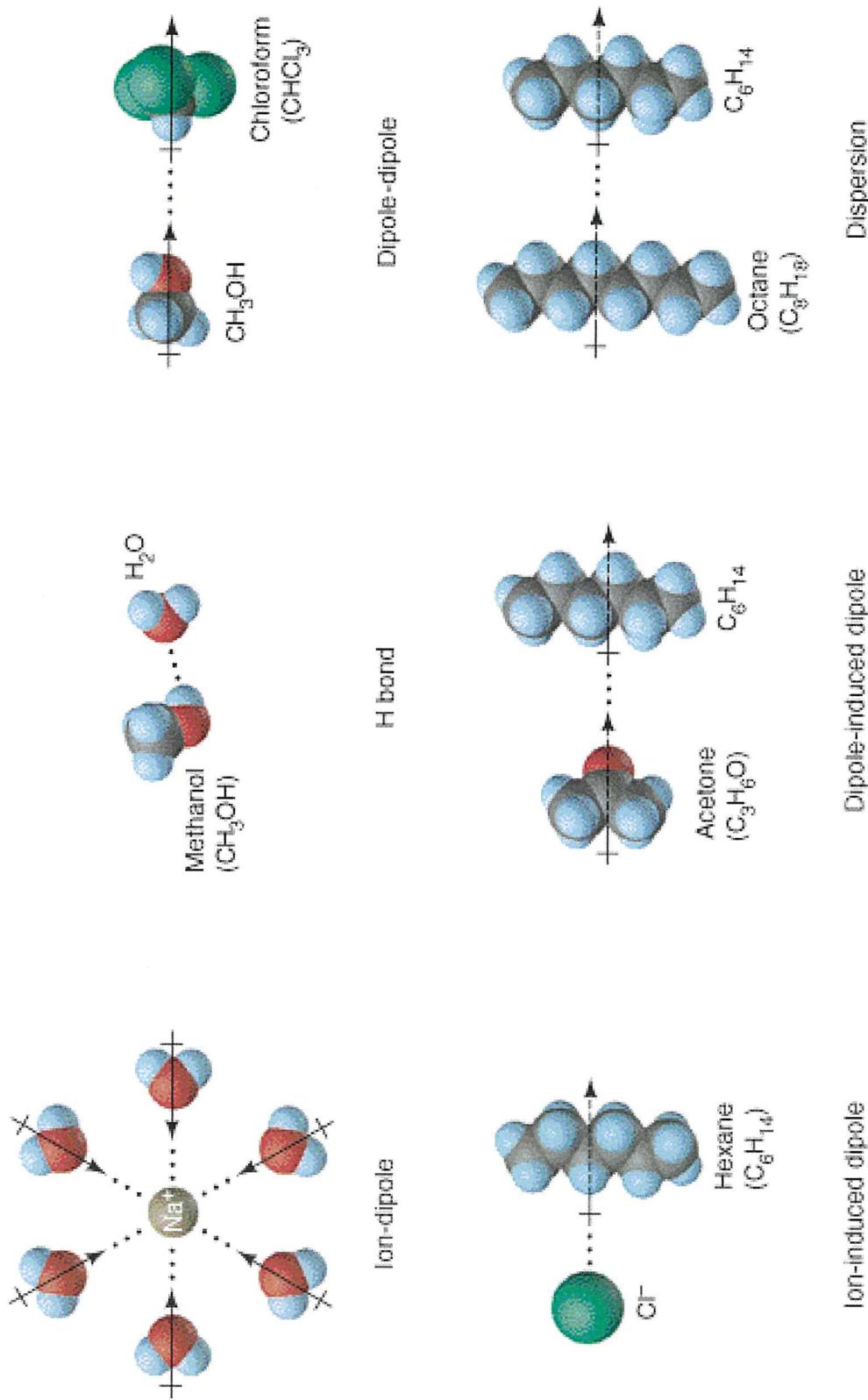
Although the HCl molecule is the most polar of the series, HI has the highest melting point. In this series of compounds the increase in dispersion forces more than compensates for the decrease in dipole-dipole forces as the electronegativity of the halogen element decreases.

The shapes of molecules also influence the strength of dispersion forces. To determine the effect of shape consider the two substances pentane and dimethylpropane that exist as non-polar molecules with the same molecular formula, C_5H_{12} , but differ in shape. The shapes of the two molecules are shown

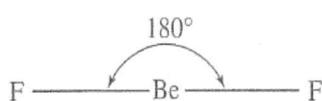
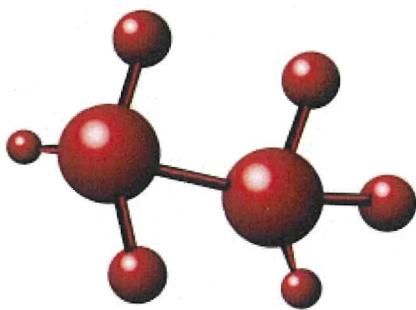


The overall attraction between molecules is greater in the case of pentane, a linear chain-like molecule with a zig-zag structure, because there can be relatively large surface contact over the entire length of the molecule. In the case of dimethylpropane, less contact is possible between molecules, thus the dispersion forces are weaker and its boiling point is lower.

Overview



Shape of Simple Molecules



The shape of a molecule depends on the arrangement of the electron pairs surrounding the central atom in the molecule. To predict the shapes of molecules the Valence Shell Electron Pair Repulsion (VSEPR) model is used. The main idea of the electron pair repulsion model is that the electron pairs in the valence energy level of an atom repel each other and are therefore arranged as far apart as possible to minimise the repulsion between them. The electron pairs which influence molecular shape include both bonding pairs and non-bonding or lone pairs in the valence energy level.

For molecules such as BeF_2 , the central Be atom has two electron pairs in its valence energy level after forming a covalent bond with each F atom. The two electron pairs achieve maximum separation by taking up positions on opposite sides of the central atom. The shape of BeF_2 is therefore linear.

PHE7

Repulsion between these electron pairs is minimised if about the central atom such that the electron pairs are at an angle of 180° . The shape of BF_3 , which has three electron pairs in its valence energy level, is described as trigonal planar.

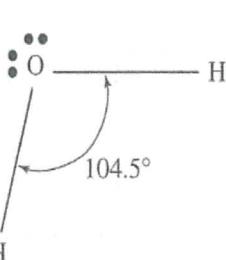
a triangular arrangement in the same plane with has three electron pairs planar.

The geometric arrangement of four electron pairs is tetrahedral. An example of a tetrahedral molecule is methane (CH_4). In CH_4 the four hydrogen atoms lie at the corners of a tetrahedron with the carbon atom at its centre. For a tetrahedral molecule, the bond angles are

that minimises repulsion CH_4 . In CH_4 the four the carbon atom at its 109.5° .

The NH_3 molecule has four electron pairs around the central atom, but one pair is a lone pair. The arrangement of electron pairs will therefore be tetrahedral, but the arrangement of the atoms is pyramidal, with the nitrogen atom at the apex of the pyramid and the three hydrogen atoms forming its base. The lone pair of electrons exerts a slightly stronger repulsive force on the bonding pairs and so the H–N–H bond angle is slightly less than the tetrahedral angle, being about 107° .

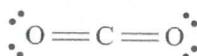
In the H_2O molecule there are two lone pairs of electrons. The arrangement of electron pairs will again be tetrahedral, but the arrangement of atoms is bent or V-shaped. The H–O–H bond angle is smaller than the tetrahedral angle because the repulsion due to the two lone pairs is slightly greater than the repulsion due to the two bonding pairs.



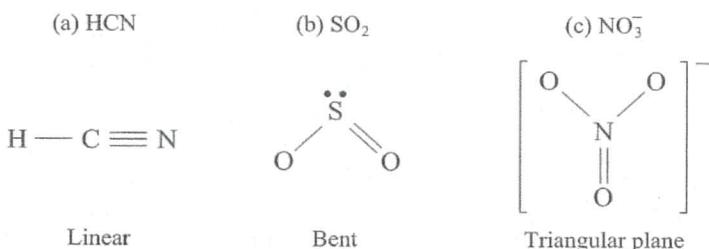
Number of electron pairs in valence level of central atom	Number of bonding pairs attached to central atom	Number of lone pairs attached to central atom	Shape (with respect to atoms)	Representation of shape	Examples
1	1	0	linear		H ₂ , HCl
2	2	0	linear		BeF ₂
3	3	0	triangular planar		BF ₃
4	4	0	tetrahedral		CH ₄ , CF ₄
4	3	1	pyramidal		NH ₃ , PCl ₃
4	2	2	V-shaped (bent)		H ₂ O, F ₂ O
4	1	3	linear		Cl ₂ , F ₂

So far only the shapes of molecules containing single bonds have been considered.

If a molecule contains a double or triple bond, its shape can still be predicted by considering the double or triple bond as one region of negative charge. To illustrate this point, consider the CO₂ molecule. The bonding in CO₂ can be represented as follows.

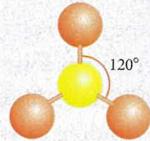
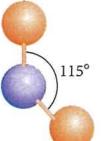
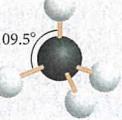
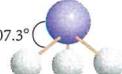
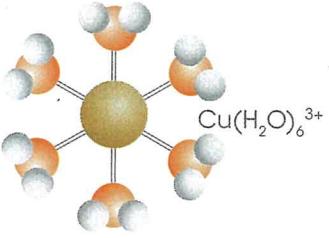


The carbon atom forms two double bonds with oxygen atoms. So far as molecular geometry is concerned there are only two regions of negative charge about the carbon atom. The bonds are directed as far apart as possible, giving a 180° bond angle. The CO₂ molecule is therefore linear. Figure 7.7 illustrates the shapes of other species containing multiple bonds.



The shapes of molecules can be determined by experimental methods such as infrared spectroscopy. Experimental evidence confirms the molecular shapes predicted using the electron pair repulsion model.

Table C6.1 Shapes of molecules

Generic formula (M, X = atoms; E = lone pair of electrons)	Number of electron pairs	Lone pairs on M	Molecular shape
MX	1	0	Linear  HCl
MX_2	2	0	Linear  CO ₂
MX_3	3	0	Trigonal planar  SO ₃
MX_2E	3	1	V-shaped  NO ₂
MX_4	4	0	Tetrahedral  CH ₄
MX_3E	4	1	Trigonal pyramidal  NH ₃
MX_2E_2	4	2	V-shaped  H ₂ O
MX_6	6	0	Octahedral  Cu(H ₂ O) ₆ ³⁺



VSEPR AND THE 3D STRUCTURES

Visit this website to see the VSEPR structures of some common molecular geometries.

Shape and Polarity

Where a molecule contains more than one polar bond, the polarity of the molecule is found by adding vectorially each of the individual bond dipoles. The bond dipoles are represented by vectors with the arrow heads representing the negative end of the dipoles. From the known molecular shapes the molecular dipole can be determined by adding the vectors representing each of the individual bond dipoles.

For a molecule to be polar the following conditions must apply:

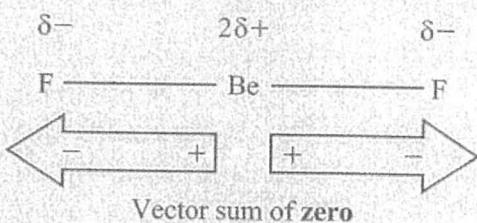
- 1 The molecule must contain one or more polar bonds.
- 2 The molecule must not be completely symmetrical.

The following examples illustrate the procedure for determining polarity.

Example 1.



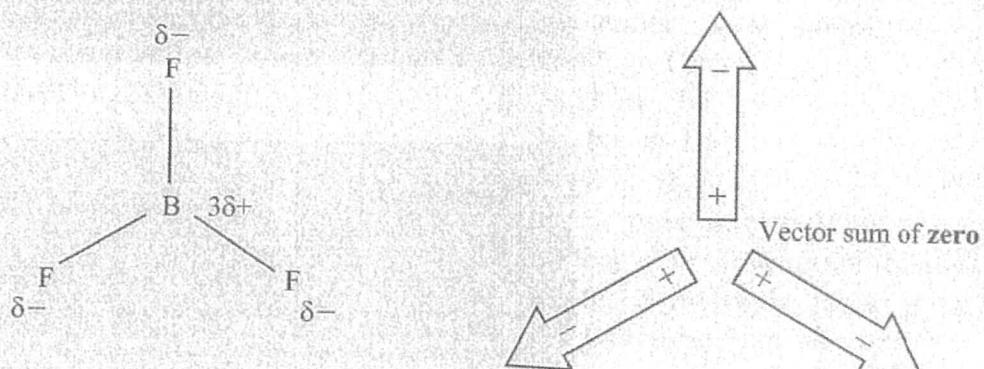
According to the electron pair repulsion model BeF_2 is linear. The fluorine atoms, being more electronegative, attract the bonding electrons more strongly than the beryllium atom so that the Be–F bonds are polar. The dipoles of the two bonds have the same magnitude but opposite directions. The vector sum of the dipoles is therefore zero and the molecule is non-polar.



Example 2.



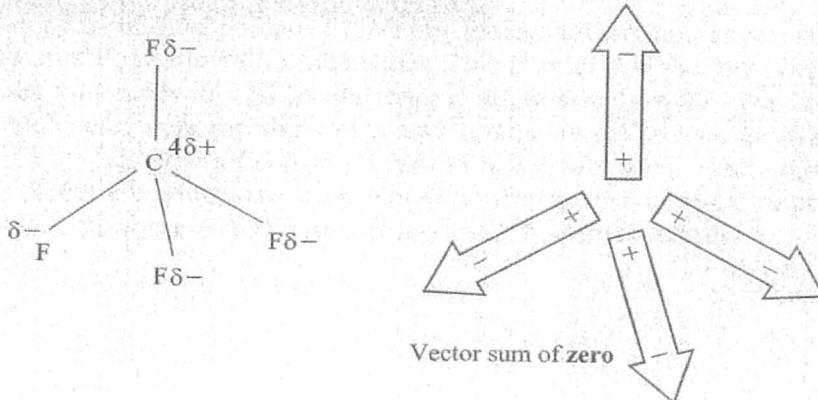
Since the electronegativity of fluorine is greater than boron, boron trifluoride has three polar bonds. The molecule is triangular planar and there is no net molecular dipole because the sum of the vectors is zero. The molecule is therefore non-polar.



Example 3.



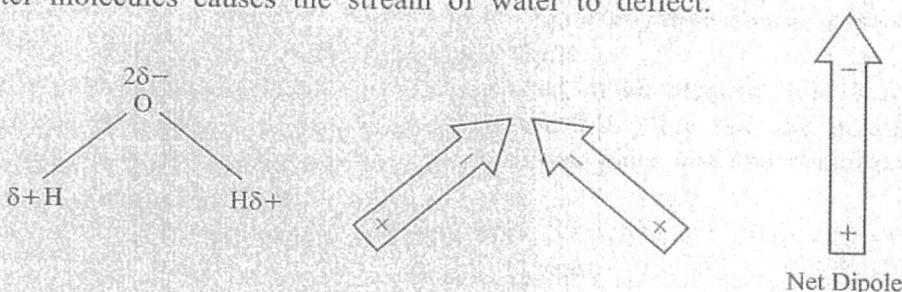
A fluorine atom attracts electrons more strongly than a carbon atom. CF_4 therefore has four polar bonds. The molecule is tetrahedral and because the sum of the vectors is zero there is no net dipole and the molecule is non-polar.



Example 4.



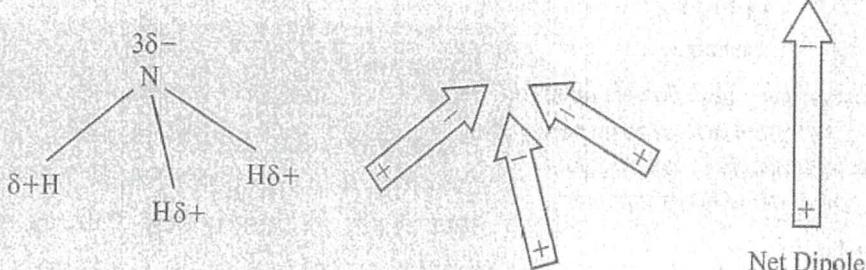
An oxygen atom attracts electrons more strongly than a hydrogen atom and both O–H bonds are polar. The vector sum of these two dipoles produces a net molecular dipole as shown. Thus the water molecule is a polar molecule. This can be shown to be true by holding a charged rod close to a stream of water running from a tap. The attraction between the rod and the polar water molecules causes the stream of water to deflect.



Example 5.



As the electronegativity of nitrogen is greater than hydrogen, the three N–H bonds in ammonia are all polar. Since the molecule is not completely symmetrical the vector sum of these dipoles is non-zero and an ammonia molecule is therefore polar.



Example 5.



The electronegativity of chlorine is greater than that of carbon and the electronegativity of carbon is slightly greater than that of hydrogen. Although the shape of a molecule of CH_3Cl is tetrahedral, the polarity of the C–Cl bond is different from the C–H bonds so there is an unsymmetrical distribution of charge. The vector sum of the dipoles is therefore not zero and the molecule is polar.

