

CHAPTER 4: Chemical Bonding

- 4.1 Ionic Bonding
- 4.2 Covalent Bonding
- 4.3 Shapes of Molecules
- 4.4 Electronegativity, Bond Polarity, Bond Length and Bond Energy
- 4.5 Intermolecular Forces
- 4.6 Metallic Bonding
- 4.7 Bonding and Physical Properties of Substances

Learning outcomes:

- (a) describe ionic (electrovalent) bonding, as in sodium chloride and magnesium oxide, including the use of 'dot-and-cross' diagrams.
- (b) describe, including the use of 'dot-and-cross' diagrams,
 - (i) covalent bonding, as in hydrogen, oxygen, chlorine, hydrogen chloride, carbon dioxide, methane, ethene.
 - (ii) co-ordinate (dative covalent) bonding, as in the formation of the ammonium ion and in the Al_2Cl_6 molecule.
- (c) explain the shapes of, and bond angles in, molecules by using the qualitative model of electron-pair repulsion (including lone pairs), using as simple examples: BF_3 (trigonal), CO_2 (linear), CH_4 (tetrahedral), NH_3 (pyramidal), H_2O (non-linear), SF_6 (octahedral), PF_5 (trigonal bipyramid).
- (d) describe covalent bonding in terms of orbital overlap, giving σ and π bonds, including the concept of hybridisation to form sp , sp^2 and sp^3 orbitals.
- (e) explain the shape of, and bond angles in, ethane and ethene in terms of σ and π bonds.
- (f) predict the shapes of, and bond angles in, molecules analogous to those specified in (c) and (e).
- (g) describe hydrogen bonding, using ammonia and water as simple examples of molecules containing N-H and O-H groups.
- (h) understand, in simple terms, the concept of electronegativity and apply it to explain the properties of molecules such as bond polarity and the dipole moments of molecules.
- (i) explain the terms bond energy, bond length and bond polarity and use them to compare the reactivities of covalent bonds.
- (j) describe intermolecular forces (van der Waals' forces), based on permanent and induced dipoles, as in $CHCl_3(l)$; $Br_2(l)$ and the liquid noble gases.
- (k) describe metallic bonding in terms of a lattice of positive ions surrounded by mobile electrons.

- (l) describe, interpret and/or predict the effect of different types of bonding (ionic bonding, covalent bonding, hydrogen bonding, other intermolecular interactions, metallic bonding) on the physical properties of substances.*
- (m) deduce the type of bonding present from given information.*
- (n) show understanding of chemical reactions in terms of energy transfers associated with the breaking and making of chemical bonds.*

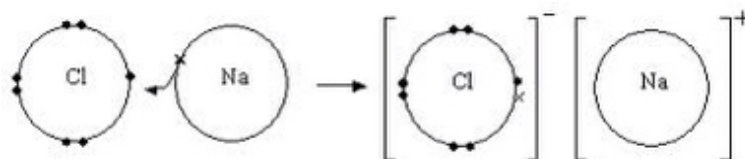
4.1 Ionic Bonding

Formation of ionic bond

- 1) *Ionic bond* is the electrostatic force of attraction between oppositely-charged ions formed by the complete transfer of electrons from an atom to another atom.
- 2) Ionic bond is also called electrovalent bond.
- 3)
 - i. An atom(usually a metal) that loses electron(s) will form a positive ion(cation).
 - ii. The electron is then **transferred** to another atom.
 - iii. The atom that gains the electron(usually a non-metal) will form a negative ion(anion).
 - iv. The cations and anions are then attracted by strong electrostatic force of attraction. The force of attraction constitutes the ionic bond.
- 4) The force of attraction between cation and anion is very strong, therefore ionic bond is a very strong bond.
- 5) Ionic bonds are non-directional, each cation will attract any neighbouring anion and vice versa to form a huge ionic lattice.
- 6) The compound formed as a result of ionic bond is called ionic compound. An example is sodium chloride, NaCl.

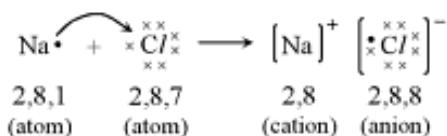
Dot-and-cross diagram

- 1) A dot-and-cross diagram shows:
 - i. the outer electron shell only.
 - ii. that the charge of the ion is spread evenly using a square bracket.
 - iii. the charge of each ion.
- 2) It is also called the Lewis diagram.



3) Using dot-and-cross diagram to represent the formation of ionic bond:

Formation of sodium chloride (NaCl)



[No electrons are shown in the 'dot-and-cross' diagram of cations – since the valence electrons are transferred completely to the anions.]

- Na atom loses an electron to form Na^+ , which has a stable neon structure (2,8).
- Cl atom gains an electron to form Cl^- , which has a stable argon structure (2,8,8).
- **Electrostatic attraction** between Na^+ and Cl^- ions constitutes the ionic bond.

Strength of ionic bonds

1) The strength of ionic bond is a measure of the electrostatic force of attraction between the ions.

2) The force of attraction between the oppositely-charged ions is proportional to the charge on the ions and inversely proportional to the square of distance between the ions.

$$E \propto \frac{(Q^+) (Q^-)}{d^2}$$

3) The strength of ionic bond is manifested in the melting point of the ionic compound.

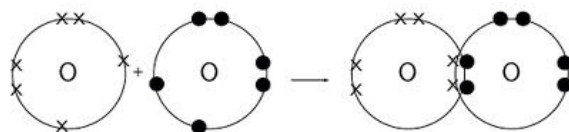
- 4) i. For instance, the melting point of NaCl is higher than NaBr.
 ii. This is because the size of Br^- ion is larger than Cl^- ion. Therefore the distance between Br^- and Na^+ is larger than that of between Cl^- and Na^+ .
 iii. As a result, the electrostatic force of attraction between Na^+ and Cl^- is stronger than that of between Na^+ and Br^- ion.

- 5) i. The melting point of NaCl is lower than MgCl_2 .
 ii. This is because Mg^{2+} ion has a higher charge than Na^+ ion. Besides that, the size of Mg^{2+} ion is smaller than Na^+ ion.
 iii. The above two factors causes the electrostatic force of attraction between Mg^{2+} and Cl^- to be stronger than that of between Na^+ and Cl^- .

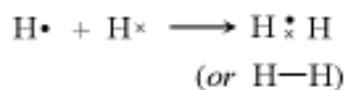
4.2 Covalent Bonding

Formation of covalent bond

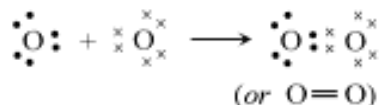
- 1) *Covalent bond* is the electrostatic force of attraction that two neighbouring nuclei have for a localised pair of electrons shared between them.
- 2) Covalent bond is formed without transferring electrons, instead, the atoms **share** their valence electron(s) to achieve duplet/octet electronic configuration.



- 3) The shared pair of electrons constitutes the covalent bond.
- 4) Using dot-and-cross diagram to represent the formation of covalent bond:



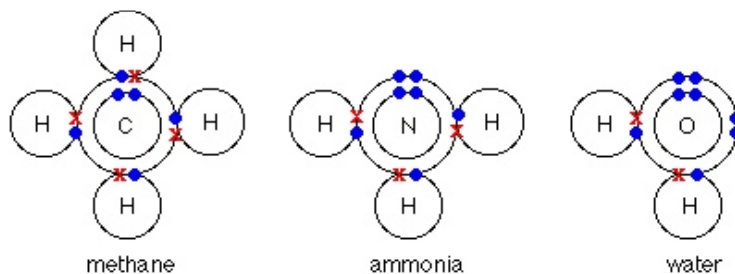
Formation of hydrogen
gas, H₂



Formation of oxygen gas,
O₂

Single bond

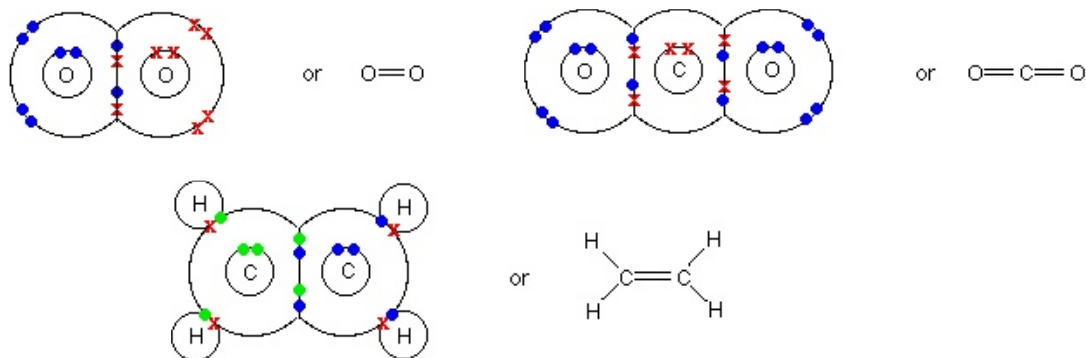
- 1) Single bond is formed when **one pair** of electrons is shared between two atoms.
- 2) Examples of compounds with single bonds:



Double bond

1) Double bond is formed when **two pairs** of electrons are shared between two atoms.

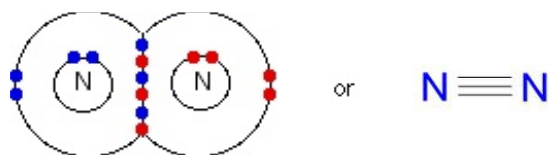
2) Examples of compounds with double bonds:



Triple bond

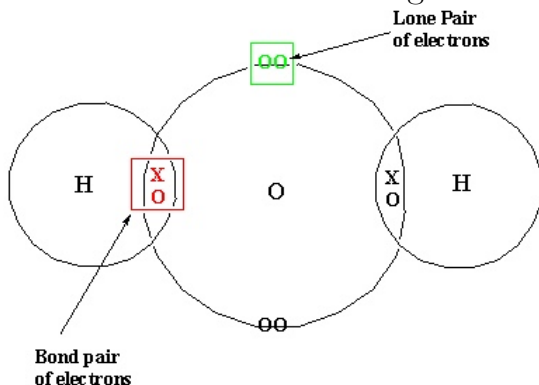
1) Triple bond is formed when **three pairs** of electrons are shared between two atoms.

2) Examples of compounds with triple bonds:



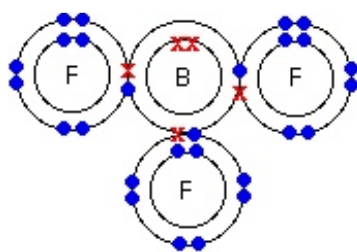
Lone pair and bond pair of electrons

1) The pair of electrons used in covalent bonding is called the bond pair while the pair of electrons not used in covalent bonding is called the lone pair.

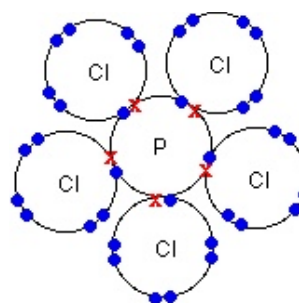


Octet-deficient and expanded octet species

- 1) In general, atoms tend to share their electrons to achieve a duplet/octet electronic configuration - the octet rule.
- 2) i. In octet-deficient species, the central atom has **less** than eight electrons.
ii. Some examples are boron trifluoride, BF_3 and nitrogen monoxide, NO .
- 4) i. In expanded octet species, the central atom has **more** than eight electrons.
ii. An example is phosphorus(V) chloride, PCl_5 .
iii. This is possible only for Period 3 elements and beyond, this is because starting from Period 3, the atoms have **empty d orbitals** in the third energy level to accommodate more than eight electrons.



Octet deficient

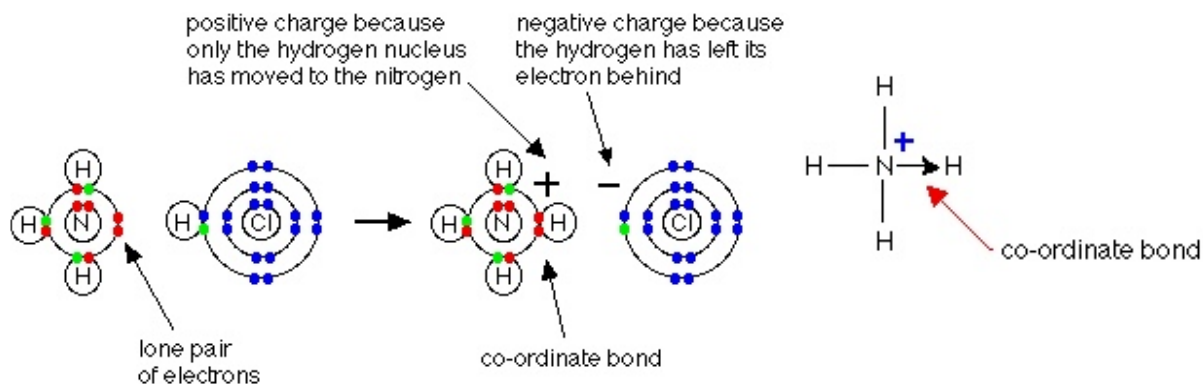


Expanded octet

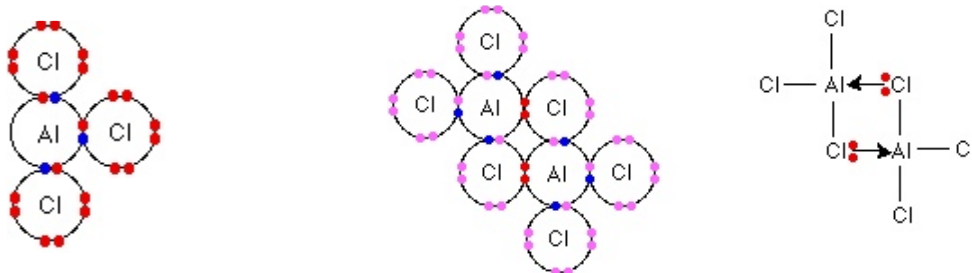
Co-ordinate bond (dative covalent bond)

- 1) A co-ordinate bond is formed when one atom provides **both** the electrons needed for a covalent bond.
- 2) Conditions of forming a co-ordinate bond:
 - i. one atom has a **lone pair** of electrons.
 - ii. another atom has an **unfilled orbital** to accept the lone pair, in other words, an electron-deficient species.
- 3) Once the bond is formed, it is identical to the other covalent bonds. It does not matter where the electrons come from.
- 4) In a displayed formula, a co-ordinate bond is represented by an arrow, the head of the arrow points away from the lone pair which forms the bond.

- 5) An example is the reaction between ammonia and hydrogen chloride. In this reaction, ammonium ion is formed by the transfer of hydrogen ion (an octet deficient species) from hydrogen chloride to the lone pair of electrons in the ammonia molecule.



- 6) Another example is aluminium chloride. At high temperature, it exists as AlCl_3 . At low temperature (around $180\text{-}190^\circ\text{C}$), it exists as Al_2Cl_6 , a dimer (two molecules joined together). This is possible because lone pairs of electrons from the chlorine atom form co-ordinate bonds with the aluminium atom.



Tips to draw dot-and-cross diagram for covalent molecules

- 1) Identify the central atom and terminal atom(s). For example, in ammonia, the nitrogen is the central atom while the hydrogens are the terminal atoms.
- 2) During the sharing of electrons, the terminal atoms **must** attain octet configuration (or duplet for hydrogen) but not necessarily for the central atom.
- 3)
 - i. If the central atom is from Period 2 of the Periodic Table, the total number of electrons surrounding it **cannot exceed eight** (but can be less than eight).
 - ii. If the central atom is from Period 3 and beyond, the total number of electrons surrounding it **can exceed eight**.

- 4) i. For polyatomic anions, the negative charge will be distributed among the **most electronegative** atom(s). This is to decrease the charge density on a particular atom and to stabilise the ion.
ii. For polyatomic cation, the positive charge will be distributed among the **less electronegative** atom(s). The reason is same as above.
- 5) If the terminal atom already has octet configuration (for example, Cl⁻), it will contribute two electrons to the central atom to form a co-ordinate bond.

4.3 Shapes of Molecules

Valence shell electron pair repulsion (VSEPR) theory

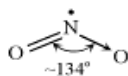
- 1) All electrons are negatively-charged, so they will repel each other when they are close together.
- 2) So, a pair of electrons in the bonds surrounding the central atom in a molecule will repel the other electron pairs. This repulsion forces the pairs of electrons apart until the repulsive forces are minimised.

- 3) The amount of repulsion is as follow:

lone pair–lone pair repulsion > lone pair–bond pair repulsion > bond pair–bond pair repulsion

and, bond pair–bond pair repulsion > bond pair–lone electron repulsion

e.g.



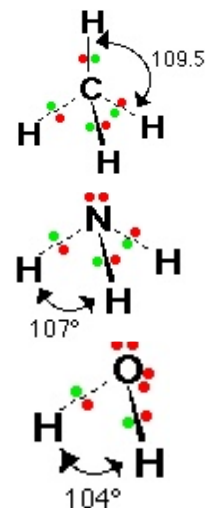
The *one lone* electron in NO₂ exerts *less repulsion* than bond pairs of electrons and so, the bond pairs spread out more from the ideal bond angle of 120°.

- 4) General steps to determine the shape of a molecule:
 - i. determine the number of valence electrons in the central atom.
 - ii. find the total number of electrons surrounding the central atom by adding the number of shared electrons to it. (Dot-and-cross diagram might be necessary)
 - iii. find the number of electron pairs by dividing the total number of electrons by two.
 - iv. determine how many pairs is/are bond pairs and lone pairs. (A double bond or triple bond is counted as one bond pair)
 - v. refer to the table to obtain the shape of the molecule

Total Domains	Generic Formula	Picture	Bonded Atoms	Lone Pairs	Molecular Shape	Electron Geometry	Example	Hybridization	Bond Angles
1	AX		1	0	Linear	Linear	H ₂	s	180
2	AX ₂		2	0	Linear	Linear	CO ₂	sp	180
	AXE		1	1	Linear	Linear	CN ⁻		
3	AX ₃		3	0	Trigonal planar	Trigonal planar	AlBr ₃	sp ²	120
	AX ₂ E		2	1	Bent	Trigonal planar	SnCl ₂		
	AXE ₂		1	2	Linear	Trigonal planar	O ₂		
4	AX ₄		4	0	Tetrahedral	Tetrahedral	SiCl ₄	sp ³	109.5
	AX ₃ E		3	1	Trigonal pyramid	Tetrahedral	PH ₃		
	AX ₂ E ₂		2	2	Bent	Tetrahedral	SeBr ₂		
	AXE ₃		1	3	Linear	Tetrahedral	Cl ₂		
5	AX ₅		5	0	Trigonal bipyramid	Trigonal bipyramid	AsF ₅	sp ³ d	90 and 120
	AX ₄ E		4	1	See Saw	Trigonal bipyramid	SeH ₄		
	AX ₃ E ₂		3	2	T shape	Trigonal bipyramid	ICl ₃		
	AX ₂ E ₃		2	3	Linear	Trigonal bipyramid	BrF ₂ ⁺		
6	AX ₆		6	0	Octahedral	Octahedral	SeCl ₆	sp ³ d ²	90
	AX ₅ E		5	1	Square pyramid	Octahedral	IF ₅		
	AX ₄ E ₂		4	2	Square planar	Octahedral	XeF ₄		

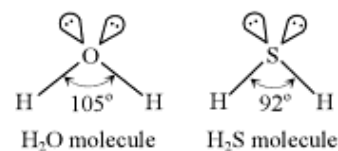
Effect of lone pair on bond angle

- 1) For methane, ammonia and water, the electron pair geometries are tetrahedral. However, the molecular geometries are different.
- 2) In methane, all the bonds are identical, repulsion between the bonds is the same. Thus, methane has a perfect tetrahedral structure with bond angle **109.5°**.
- 3) In ammonia, the repulsion between the lone pair and the bond pairs is stronger than in methane. This forces the bond angle to decrease slightly to **107°**.
- 4) In water, there are two lone pairs and thus the repulsion is the greatest, the two bond pairs are pushed closer to one another and the bond angle is reduced to **104.5°**.



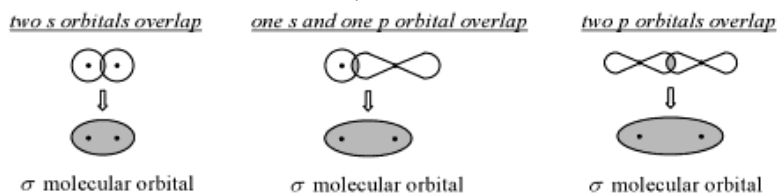
Effect of electronegativity on bond angle

- 1) Water and hydrogen sulfide have the same general shape with the same number of bond pairs and lone pairs. However, their bond angles are different.
- 2) This is because oxygen has a higher electronegativity than sulphur. The bond pairs of electrons are closer to the oxygen atom compared to the sulfur atom.
- 3) This results in **greater repulsion** in the O-H bonds than in the S-H bonds. Therefore, the bond angle increases from 92.5° to 104.5°.

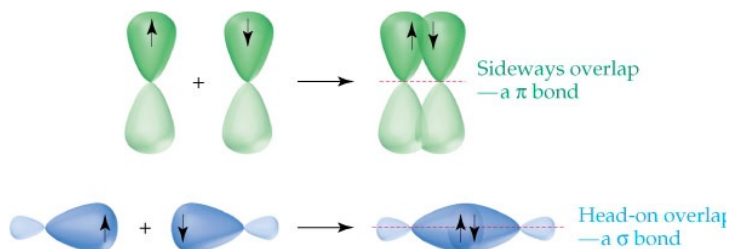
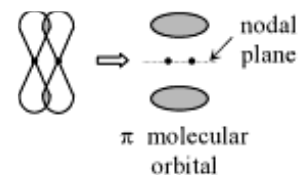


Sigma(σ) bond and pi(π) bond

- 1) A sigma bond is formed by orbitals from two atoms overlapping **end-to-end**.
- 2) In a sigma bond, the electron density is concentrated between the two nuclei.



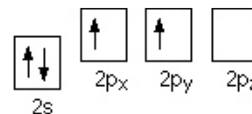
- 3) A pi bond is formed by the p orbitals from two atoms overlapping **sideways**.
- 4) In a pi bond, there are two regions of high electron density alongside the nuclei.
- 5) A pi bond is weaker than a sigma bond because the overlapping of charge clouds is less than in a sigma bond.
- 6) In covalent molecules, **single bonds** are sigma bonds(σ), a **double bond** consists of one sigma bond and one pi bond($1\sigma, 1\pi$), and a **triple bond** consists of one sigma bond and two pi bonds($1\sigma, 2\pi$).



Hybridisation

- 1) *Hybridisation* is the mixing of atomic orbitals to produce a new set of hybrid orbitals of equivalent energies. This is a theory used to account for the discrepancies in explaining the shapes of covalent molecules.

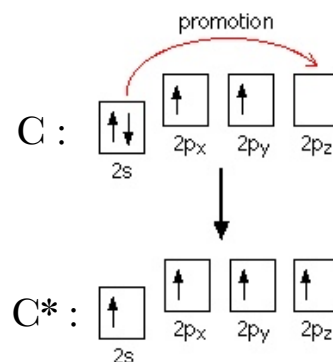
- 2) There is a problem with simple view on methane, CH_4 . Methane has two unpaired electrons only in its outer shell to share with the hydrogen atoms, but the formula of methane is not CH_2 .



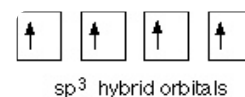
- 3) The concept of hybridisation is used to account for this discrepancy.
- 4) General steps in hybridisation:
 - i. promotion of electron.
 - ii. mixing of orbitals to produce a new set of hybrid orbitals of equivalent energies(sp, sp^2 or sp^3 hybrid orbitals)
 - iii. forming of a new molecular orbital.

sp³ hybridisation

- 1) An example of compound which undergoes sp³ hybridisation is methane, CH₄.
- 2) The carbon atom uses some energy to **promote** one of its electron from 2s to empty 2p orbital so that there are **four unpaired electrons** for covalent bonding.
- 3) The carbon now is said to be in an **excited state(C*)**.



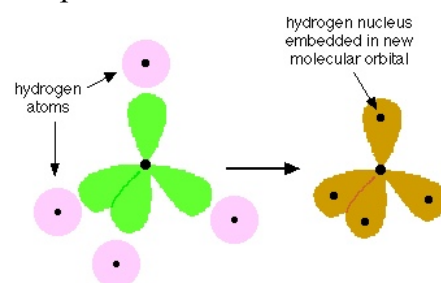
- 4) The orbitals then 'mix' or **hybridise** to produce **four hybrid orbitals of equivalent energies**. The new orbitals are called sp³ hybrid orbitals because they are made from one s orbital and three p orbitals.



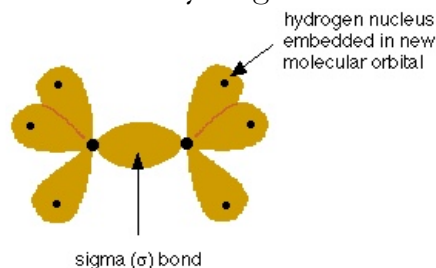
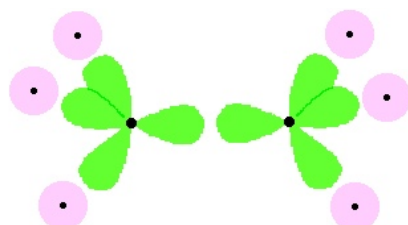
- 5) Each hybrid orbital has one big lobe and one small lobe. They rearrange themselves so that they are as far as possible to form a tetrahedral geometry. The hybrid orbitals are 109.5° apart.



- 6) The s orbitals from the hydrogen atoms then overlap with the four hybrid orbitals to form **four sigma bonds** because the overlapping is end-to-end. All the bonds are identical.



- 7) Another example is ethane, C₂H₆. The two carbon atoms undergo sp³ hybridisation to form four hybrid orbitals. The two carbon atoms are bonded by overlapping one of their hybrid orbitals. The remaining ones then overlap with the s orbitals of the hydrogen atoms.

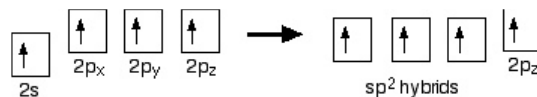


- 8) The bond angle is approximately 109.5°. This is an approximation because all the bonds are not identical.

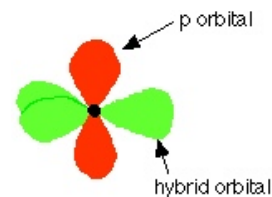
sp² hybridisation

1) An example of compound which undergoes sp² hybridisation is ethene, C₂H₄.

2) The same thing happens as in sp³ hybridisation, except that this time the carbon atoms 'mix' or hybridise **three of the four orbitals only** because the carbon atom is bonding with three other atoms only.

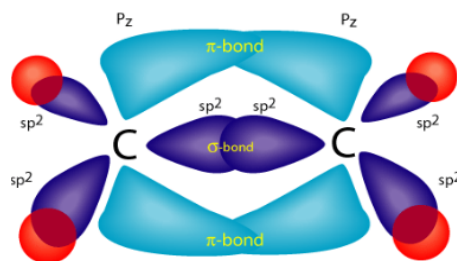


3) This produces **three sp² hybrid orbitals** because they are made from one s orbital and two p orbitals. Another p orbital **remains unchanged** and it is perpendicular to the plane containing the hybrid orbitals.

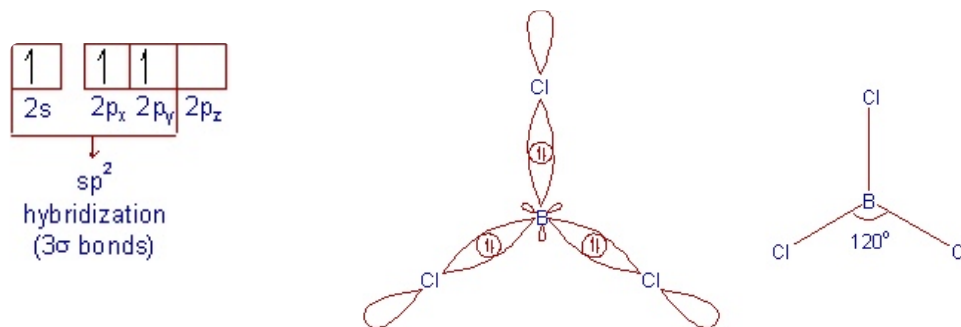


4) The hybrid orbitals rearrange themselves so that they are as far as possible, that is, a trigonal planar arrangement, the hybrid orbitals are 120° apart.

5) The hybrid orbitals then overlap with s orbitals from the hydrogen atoms and another hybrid orbital from the other carbon atom to form **five sigma bonds**. The remaining p orbitals overlap sideways to form **a pi bond**. A **double bond** is formed between the two carbon atoms.



6) Another example is boron trichloride, BCl₃. The boron atom undergoes sp² hybridisation to produce three sp² hybrid orbitals. The hybrid orbitals rearrange themselves to form a trigonal planar geometry. The p orbitals from chlorine atoms then overlap with the hybrid orbitals to form three sigma bonds.



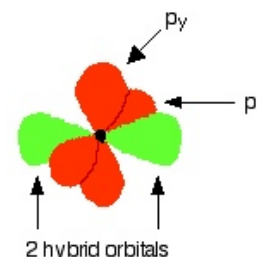
sp hybridisation

1) An example of compound which undergoes sp hybridisation is ethyne, C_2H_2 .

2) The same thing happens as in sp^3 and sp^2 hybridisation, except that this time the carbon atoms 'mix' or hybridise **two of the four orbitals** only because the carbon atom is bonding with two other atoms only.

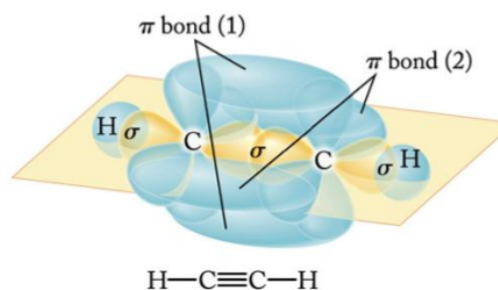
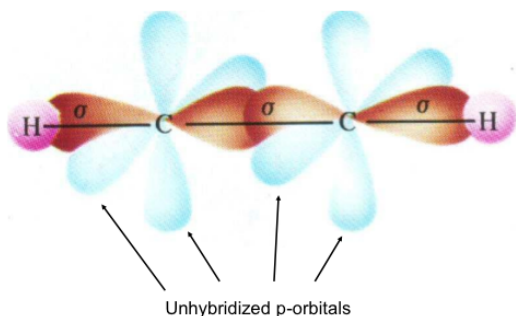


3) This produces **two sp hybrid orbitals** because they are made from one s orbital and one p orbital. The other two p orbitals **remain unchanged** and they are perpendicular to each other and to the two hybrid orbitals.



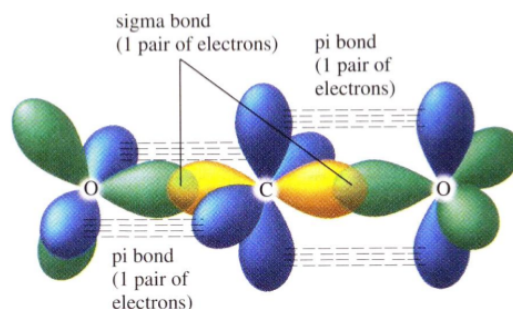
4) The hybrid orbitals rearrange themselves so that they are as far as possible, that is, a linear arrangement, the hybrid orbitals are 180° apart.

5) The hybrid orbitals overlap with the s orbitals from the hydrogen atoms and to the hybrid orbital from the other carbon atom to form **three sigma bonds**. The remaining p orbitals overlap sideways to form **two pi bonds**. A **triple bond** is formed between the two carbon atoms.



Example of covalent molecule with multiple hybridisations

1) In carbon dioxide, CO_2 , the carbon atom undergoes sp hybridisation while the oxygen atoms undergo sp^2 hybridisation. The overlapping of the hybrid and p orbitals are shown in the diagram.



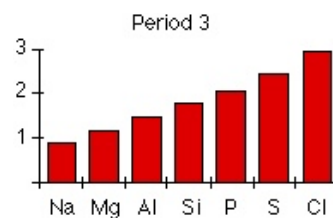
4.4 Electronegativity, Bond Polarity, Bond Length and Bond Energy

Electronegativity

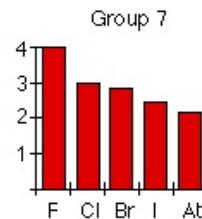
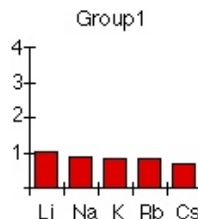
- 1) *Electronegativity* is the ability of an atom which is covalently bonded to the other atom to attract the bond pair of electrons towards itself.
- 2) The more electronegative an atom is, the higher the tendency of that atom to attract the bond pair of electrons towards itself.
- 3) The Pauling scale is commonly used to quantify the value of electronegativity of a particular element.
- 4) **Fluorine** is the most electronegative element because of its small size, followed by **oxygen** and **nitrogen**.

Trends of electronegativity values in the Periodic Table

- 1) i. The value of electronegativity **increases** across a Period (from left to right).
ii. This is because the number of proton increases across a Period. Therefore the amount of positive charge in the nucleus also increases.
iii. The shielding effect by inner electrons remains constant.
iv. Therefore the attraction towards the bond pair of electrons increases, making it more electronegative.



- 2) i. The value of electronegativity **decreases** down a Group.
ii. This is because the size of the atoms increases down a Group. Therefore the distance between the nucleus and the bond pair of electrons also increases.
iii. The shielding effect by inner electrons is also greater.
iv. Therefore the attraction towards the bond pair of electrons decreases, making it less electronegative.



Bond Polarity

1) i. When two covalently-bonded atoms have the same electronegativity, the electron cloud is evenly distributed between the two atoms.



ii. The bond is described as a 'pure' covalent bond or non-polar bond.

iii. Some examples are H₂, Cl₂ and Br₂.

2) i. However, when an atom is more electronegative than the other, the more electronegative atom will attract the bond pair of electrons more towards itself. The electron cloud is not evenly distributed or **distorted**.



ii. The more electronegative end acquires a partial negative charge while the less electronegative end acquires a partial positive charge.

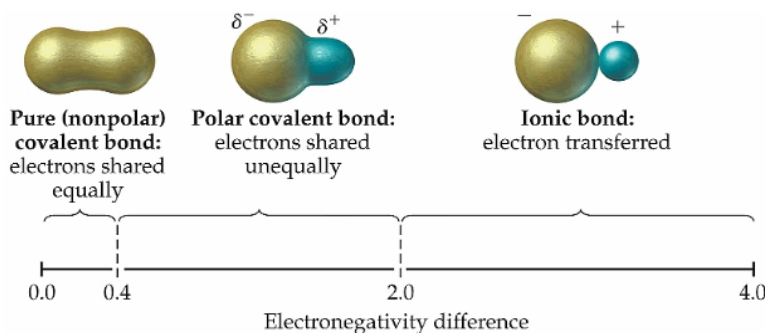
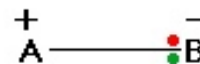
iii. The bond is said to be **polarised**, or, a polar bond.

iv. Some examples of compound which contain polar bond(s) are HCl and CH₄.

v. The molecule is described as 'covalent with some ionic character'.

3) i. When the electronegativity difference between the two atoms is very great, the less electronegative atom will lose its electrons and the more electronegative atom will gain the electrons.

ii. An ionic bond will be formed.



Polar and non-polar molecules

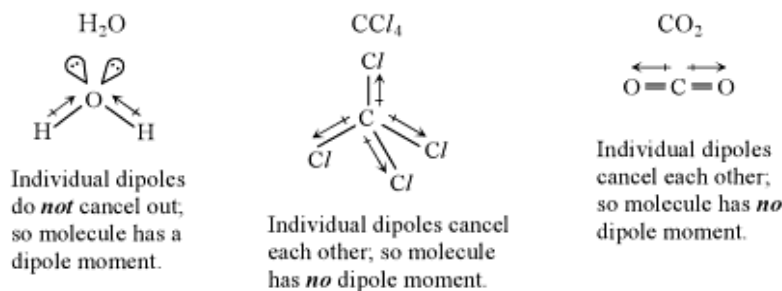
1) A molecule is polar, and thus, has a dipole moment ($\mu \neq 0$) if:

i. the bonds are polarised

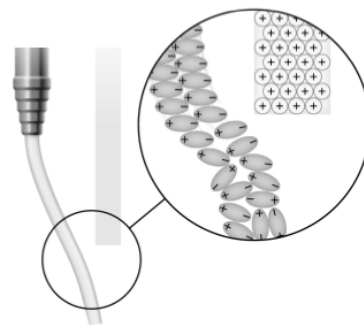
ii. the dipole of the bonds do not cancel out each other (in other words, it is asymmetrical)

2) The *dipole moment*, μ is the product of charges and the distance between the centre of the charges.

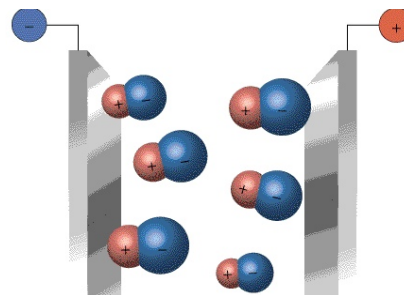
3) Examples:



- 4) A liquid containing polar molecules **can be deflected** by a charged rod brought near to it. This is because there is a positive end and a negative end in polar molecules. So, irregardless of the charge on the rod, one end of the molecule will always be attracted towards the charged rod.



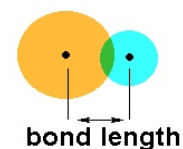
- 5) When polar molecules are placed in an electric field, the positive end of the molecule will face towards the negative terminal while the negative end of the molecule will face towards the positive terminal.



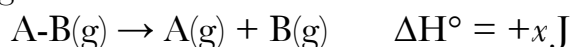
- 6) Generally, polar molecules are more reactive than non-polar molecules because many chemical reactions are started by a reagent attacking an electrically-charged end of the polar molecule. An example is CO is more reactive than N₂ although both of them have triple bonds because CO is polar while N₂ is not.

Bond length and bond energy

- 1) *Bond length* is the distance between two nuclei of two atoms joined by a covalent bond.



- 2) *Bond energy* is the energy needed to break one mole of covalent bonds between two atoms in the gaseous state.

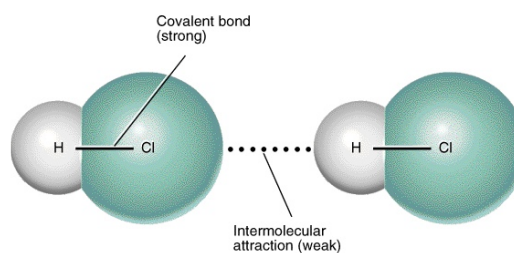


- 4) i. The **shorter** the bond length, the **stronger** is the bond.
 ii. The **greater** the bond energy, the **stronger** is the bond.
 iii. Hence, the **shorter** and bond length, the **greater** the bond energy.

4.5 Intermolecular Forces

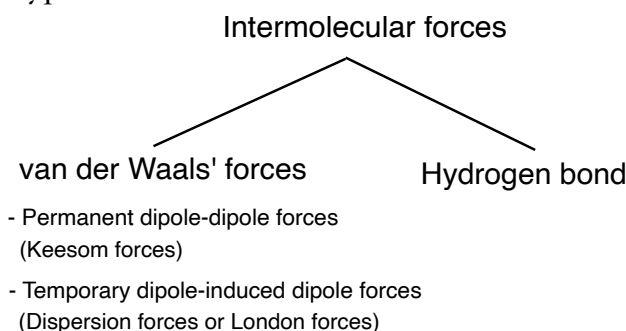
Intermolecular and intramolecular forces

- i. Intramolecular force is the force of attraction that hold individual atoms together **in a molecule**.
ii. Commonly referred to as covalent bonds.



- Intermolecular force is the force of attraction **between one molecule and the neighbouring molecule**.

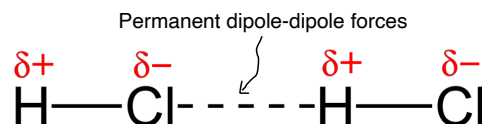
- There are several types of intermolecular forces:



- Intermolecular forces are **much weaker** than intramolecular forces.
- Intermolecular forces are responsible for the melting and boiling points of substances, as well as their physical states.

Permanent dipole-dipole forces (Keesom forces)

- Polar molecules have a negative end and a positive end. The oppositely-charged ends will attract one another. The forces of attraction is called permanent dipole-dipole forces.

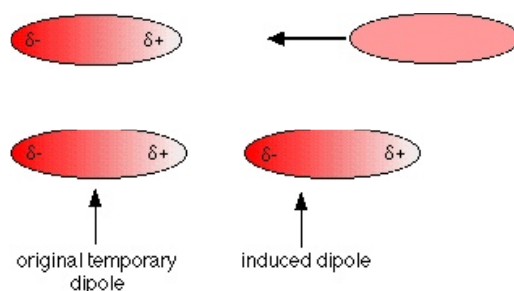


- Only polar molecules** will experience permanent dipole-dipole forces.

Temporary dipole-induced dipole forces (Dispersion forces or London forces)

- 1) Electrons are mobile, and constantly revolving around the nucleus. Even in non-polar molecules, there is a high possibility that at any given instant, the electron density is higher on one side than the other.
- 2) This will result in the formation of **temporary dipole** (or instantaneous dipole) because it only lasts for just a tiny fraction of time.
- 3) In the next instance, the distribution of electron density will change and the molecule has a new temporary dipole.

- 4) The temporary dipole set up can distort the electron charge clouds of the neighbouring molecules, giving rise to **induced dipoles**.



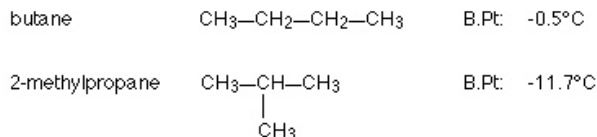
- 5) The forces of attraction between temporary dipoles and induced dipoles give rise to dispersion forces.

- 6) **All molecules (polar and non-polar)** will experience dispersion forces. For polar molecules, permanent dipole-dipole forces exist **in addition** with dispersion forces.

Factors affecting strength of dispersion forces

- 1) **Number of electrons in the molecule (Size of molecule)**
 - As the number of electrons increases, the size of the molecule also increases.
 - This causes the attraction between the nucleus and the electrons to get weaker, the electrons become progressively easier to be distorted.
 - This causes more temporary dipoles to be set up and the dispersion forces get stronger.
 - The boiling point of halogens increases going down the Group. The size of the molecules increases and thus the van der Waals' forces become stronger.
- 2) **Number of contact points between the molecules (Surface area)**
 - For two molecules of the same number of electrons, the one which has a larger surface area has more contact points between the molecules.
 - The more contact points between molecules, more temporary dipoles can be set up and the dispersion forces become stronger.

- Butane is a **long chain** molecule while 2-methylpropane is more **spherical** (more compact). This causes the number of contact point between butane molecules to be more than in 2-methylpropane.
- Therefore the van der Waals' forces of attraction in butane is stronger than in 2-methylpropane. This causes the boiling point of butane to be higher.

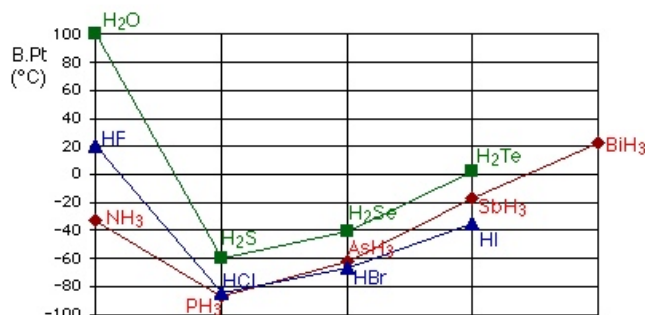


Hydrogen bonding

- 1) *Hydrogen bond* is the electrostatic force of attraction between a hydrogen atom (which is covalently bonded to a small and highly electronegative atom) and the lone pair of electrons of another small and highly electronegative atom.
- 2) The 'small and highly electronegative atoms' are fluorine, oxygen and nitrogen.
- 3) The conditions necessary for forming hydrogen bonds:
 - i. a **hydrogen atom** bonded **directly** to a **small and electronegative atom**, causing it to acquire a significant amount of partial positive charge.
 - ii. a **lone pair** of electrons from the small and electronegative atom.
- 4) The attraction between the hydrogen atom and the lone pair of electrons constitutes the hydrogen bond.

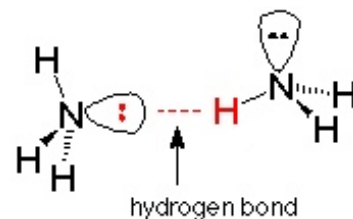
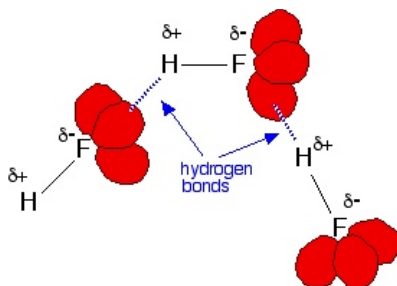
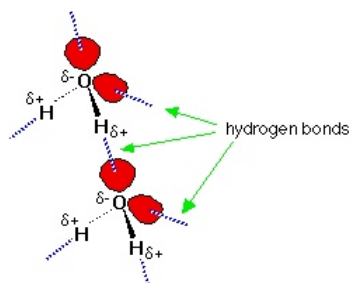
5) The evidence of hydrogen bonding:

- H_2O , HF and NH_3 have exceptionally high boiling points compared to other compounds with greater number of electrons.
- This suggests that there must be an additional intermolecular force other than van der Waals' forces exists between them.



- This additional intermolecular force is called the hydrogen bond.

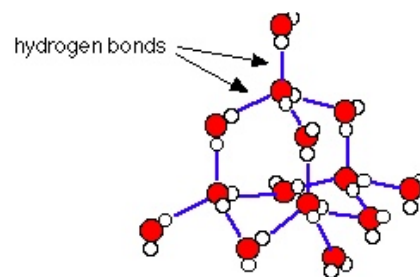
6) H_2O has a higher boiling point compared to NH_3 and HF. This is because a H_2O molecule can form, on average four hydrogen bonds. In NH_3 , the number of hydrogen bonds is restricted by the one lone pair of electrons in the NH_3 molecule. In HF, it is restricted by the number of hydrogen atoms.



The peculiar properties of water

1) **Solid ice is less dense than liquid water**

- Most solids are denser than their liquids, this is because in the solid state, the molecules are packed closer and occupies smaller volume.
- However, solid ice is less dense than liquid water.
- This is because in the solid state, the water molecules are bonded to each other through hydrogen bonding to form a **giant 3-dimensional tetrahedral structure**. This structure has a lot of **empty spaces** in between the water molecules.
- When heat is supplied, some of the hydrogen bond breaks down and the rigid structure collapses, filling up the spaces in between them, thus decreasing the volume occupied and increasing its density.



2) **Solubility of substances in water**

- Most covalent compounds are not soluble water while ionic compounds are.
- However, some covalent compounds are soluble in water because they can form hydrogen bonds with water, some examples are ammonia and alcohols

3) **High boiling point of water**

- The boiling point of water is exceptionally high compared to other Group 16 hydrides(see graph at page 21).
- This is because water molecules are capable of forming hydrogen bonds between themselves. This additional type of intermolecular forces increases its boiling point significantly.

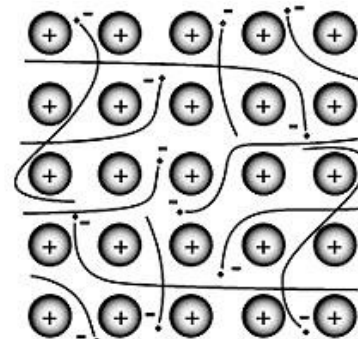
4) **High surface tension and viscosity**

- Water has a high viscosity. This is because hydrogen bonding reduces the ability of water to slide over each other, making it more viscous.
- Water has a high surface tension. This is because hydrogen bonds also exert a significant downward force at the surface of liquid.

4.6 Metallic Bonding

Metallic bonds

- 1) *Metallic bond* is the electrostatic force of attraction between the delocalised electrons and the positive metal ions.
- 2) Take sodium as an example, in the solid state, sodium atoms are packed so closely together that the 3s orbitals (containing an electron) overlap with one another to form a giant molecular orbital.
- 3) The 3s electrons are then free to move throughout the lattice structure, it is no longer bound to any sodium atom. The electrons are said to be **delocalised**. The delocalised electrons are sometimes referred to as '**sea of delocalised electrons**'.
- 4) The attraction between the delocalised electrons and the positive metal ions constitutes the metallic bond.



Factors affecting the strength of metallic bonds

- 1) **Number of electrons in the sea of delocalised electrons**
 - The **more** the number of electrons delocalised, the **stronger** the metallic bond.
 - For example, the melting point of magnesium is higher than sodium.
 - This is because one magnesium atom can donate two electrons from its outer shell into the sea of delocalised electrons while sodium can only donate one.
 - There is twice amount of delocalised electrons in magnesium than in sodium. So, the metallic bond in magnesium is stronger due to the stronger attraction.
- 2) **Size of the metal**
 - The **smaller** the size of the metal, the **stronger** the metallic bond.
 - For example, the melting point of magnesium is lower than beryllium although both of them have two electrons delocalised per atom.
 - This is because the size of magnesium is bigger, causing the attraction between the delocalised electrons and the positive nucleus to be weaker. The shielding effect in magnesium is also more than in beryllium.
 - So, the metallic bond in magnesium is weaker than in beryllium.

4.7 Bonding and Physical Properties of Substances

Physical state at room temperature

1) **Ionic compounds**

- Ionic compounds are solids at room temperature, this is because the ionic bond holding the oppositely-charged ions is very strong, a lot of energy is required to overcome the strong forces of attraction.

2) **Covalent compounds**

- Most simple covalent molecules (like water and ammonia) are liquids or gases at room temperature. This is because the intermolecular force between the molecules is weak, little energy is required to overcome it.
- Some simple covalent molecules (like iodine) are solids because the intermolecular force between them is strong enough.
- For giant covalent structures (like diamond and silicon dioxide), they are solids at room temperature. This is because the covalent bonds holding the atoms are very strong, a lot of energy is required to overcome it.

3) **Metals**

- Metals (except mercury) are solids at room temperature. This is because the metallic bond holding the metal ions is very strong, a lot of energy is required to overcome it.

Electrical conductivity

1) **Ionic compounds**

- Ionic compounds do not conduct electricity in the solid state, this is because the ions are not free to move.
- In the molten or aqueous state, it conducts electricity because the ions are free to move (mobile ions are present).

2) **Covalent compounds**

- Simple covalent molecules do not conduct electricity because the ions are **not free to move**.
- Some giant covalent structures (like graphite) are able to conduct electricity because the electrons are free to move (delocalised). (See also Chapter 5)

3) **Metals**

- Metals conduct electricity because the delocalised electrons are free to move.

Solubility

1) **Ionic compounds**

- Most ionic compounds are soluble in water (polar solvents) and insoluble in non-polar solvents.
- This is because in non-polar solvents, the molecules are held together by weak intermolecular forces. The ionic bonds in ionic compounds are much stronger and the energy needed to break them is high. If ion-solvent bonds are formed, the energy released is **not enough to compensate** the energy absorbed, making the entire structure **unstable**.
[Note: To form ion-solvent bonds, the ion-ion (ionic bond) and solvent-solvent (van der Waal's forces) bonds must be broken first.]
- While in polar molecules, there is an attraction between polar molecules and the ions. So, the energy released when ion-solvent bonds are formed is **enough to compensate** the energy needed to break the strong ionic bonds.

2) **Covalent compounds**

- This is because when a covalent molecule is dissolved in a non-polar solvent, the molecule-solvent attraction is strong enough to compensate the energy needed to break the weak van der Waal's forces between covalent molecules.
- While when dissolved in polar solvents, the energy needed to break the attractions in polar solvents is too high. The energy released when molecule-solvent attraction set up is not enough to compensate it. This makes the structure to gain energy overall, making it less stable.
- Some covalent compounds react with water rather than dissolving in it. For example, hydrogen chloride reacts with water to form hydrogen ions and chloride ions, and the ions are soluble.

3) **Metals**

- Metals do not dissolve in polar and non-polar solvents.
- However, some metals like sodium and calcium can react with water.