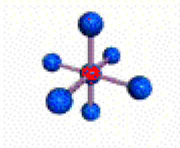
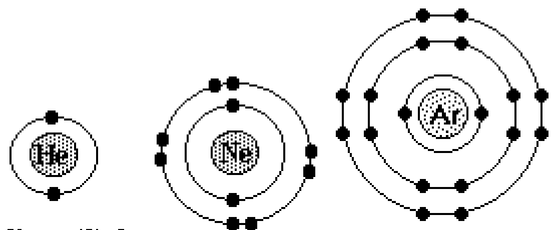
Structure, Bonding and properties



1. Why do atoms bond together?

Some atoms are very reluctant to combine with other atoms and exist in the air around us as single atoms. These are the **Noble Gases** and have **very stable electron arrangements** eg **2**, **2.8** and **2.8.8 because their outer shells are full**. The first three are shown in the diagrams below and explains why Noble Gases are so reluctant to form compounds with other elements.



helium (2) 2 neon (10) 2.8 Ar (18) 2.8.8 **(atomic number) and electron arrangement.**

All other atoms therefore, bond together to become electronically more stable, that is to become like Noble Gases in electron arrangement. Bonding produces new substances and usually involves only the 'outer shell' or 'valency' electrons

The phrase CHEMICAL BOND refers to the strong electrical force of attraction between the atoms or ions in the structure. The combining power of an atom is sometimes referred to as its **valency** and its value is linked to the number of outer electrons of the original uncombined atom

COVALENT BONDING - **sharing electrons** to form **molecules** with **covalent bonds,** the bond is usually formed between **two non-metallic elements in a molecule.**

IONIC BONDING - By one atom **transferring electrons** to another atom.

COORDINATE COVALENT BONDING A coordinate covalent bond is special because it involves a shared pair of electrons that came from a single atom.

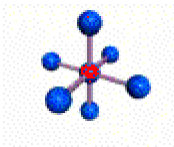
METALLIC BONDING The **crystal lattice of metals consists of ions** NOT atoms surrounded by a '**sea of electrons**' forming **giant lattice**. These free or 'delocalised' electrons are the 'electronic glue' holding the particles together. There is a **strong electrical force of attraction between** these **mobile electrons (-)** and the 'immobile' **positive metal ions (+)** and this is the **metallic bond**.

- **An ion is an atom or group of atoms carrying an overall positive or negative charge**
  - **eg Na^+, Cl^-, $[Cu(H_2O)]^{2+}$, SO_4^{2-}** etc.
- If a particle, as in a neutral atom, has equal numbers of protons (+) and electrons (-) the particle charge is zero ie no overall electric charge.
- The proton number in an atom does not change BUT the number of associated electrons can!
- If negative electrons are removed the excess charge from the protons produces an overall positive ion.
- If negative electrons are gained, there is an excess of negative charge, so a negative ion is formed.
- The charge on the ion is numerically related to the number of electrons transferred.
- For any atom or group of atoms, for every electron gained you get a one unit increase in negative charge, for every electron lost you get a one unit increase in the positive

The atom **losing electrons** forms a **positive ion (cation)** and is usually a **metal**.

The atom **gaining electrons** forms a **negative ion (anion)** and is usually a **non-metallic element**.

NOBLE GASES are very reluctant to share, gain or lose electrons to form a chemical bond. For most other elements the types of bonding and the resulting properties of the elements or compounds are described in detail below. **In all the electronic diagrams ONLY the outer electrons are shown.**

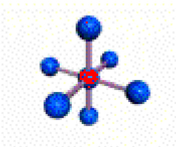


2. Covalent Bonding

Covalent bonds are formed by atoms **sharing electrons** to form **molecules**. This type of bond usually formed between **two non-metallic elements.** The molecules might be that of an **element ie one type of atom** only OR from **different elements** chemically combined to form a **compound**.

The covalent bonding is caused by the mutual electrical attraction between the two positive nuclei of the two atoms of the bond, and the electrons between them.

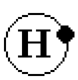
One **single covalent bond** is a sharing of **1 pair of electrons**, two pairs of shared electrons between the same two atoms gives a double bond and it is possible for two atoms to share 3 pairs of electrons and give a triple bond.



The bonding in Small Covalent Molecules

The simplest molecules are formed from two atoms and examples of their formation are shown below. The electrons are shown as dots and crosses to indicate which atom the electrons come from, though all electrons are the same. The diagrams may only show the outer electron arrangements for atoms that use two or more electron shells. Examples of simple covalent molecules are …

Example 1: 2 hydrogen atoms (1) form the molecule of **the element hydrogen H_2**

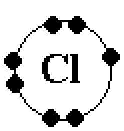


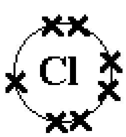
and

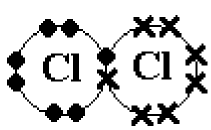

combine to form


where both atoms have a pseudo helium structure of 2 outer electrons around each atom's nucleus. **Any covalent bond is formed from the mutual attraction of two positive nuclei and negative electrons between them**.

Example 2: 2 chlorine atoms (2.8.7) form the molecule of **the element chlorine Cl_2**

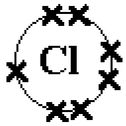
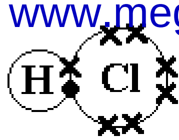


and


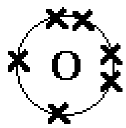
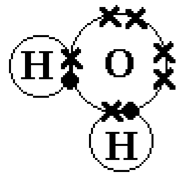
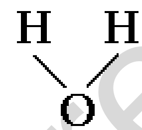
combine to form


where both atoms have a pseudo argon structure of 8 outer electrons around each atom. All the other halogens would be similar eg F_2, Br_2 and I_2 etc. Valency of halogens is 1 here.

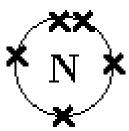
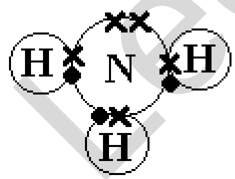
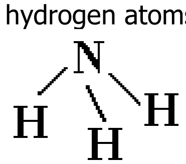
Example 3: 1 atom of hydrogen (1) combines with 1 atom of chlorine (2.8.7) to form the molecule of the **compound hydrogen chloride HCl**

 and  combine to form  where hydrogen is electronically like helium and chlorine like argon. All the other hydrogen halides will be similar eg HF, HBr and HI etc.

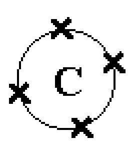
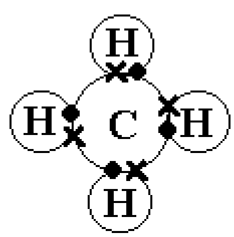
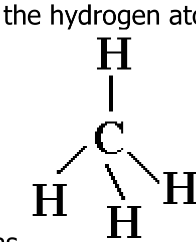
Example 4: 2 atoms of hydrogen (1) combine with 1 atom of oxygen (2.6) to form the molecule of the **compound** we call **water H_2O**

 and  and  combine to form  so that the hydrogen atoms are electronically like helium and the oxygen atom becomes like neon. The molecule can be shown as  with two hydrogen - oxygen single covalent bonds. Hydrogen sulphide will be similar, since sulphur (2.8.6) is in the same Group 6 as oxygen. Valency of oxygen and sulphur is 2 here.

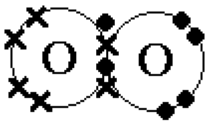
Example 5: 3 atoms of hydrogen (1) combine with 1 atom of nitrogen (2.5) to form the molecule of the **compound** we call **ammonia NH_3**

three of  and one  combine to form  so that the hydrogen atoms are electronically like helium and the nitrogen atom becomes like neon. The molecule can be shown as  with three nitrogen - hydrogen single covalent bonds (

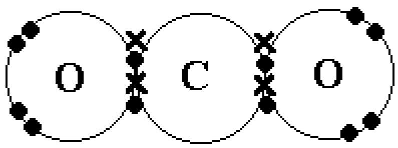
Example 6: 4 atoms of hydrogen (1) combine with 1 atom of carbon (2.4) to form the molecule of the **compound** we call **methane CH_4**

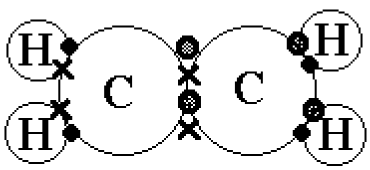
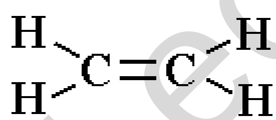
four of  and one of  combine to form  so that the hydrogen atoms are electronically like helium and the nitrogen atom becomes like neon. The molecule can be shown as  with four carbon - hydrogen single covalent bonds. SiH_4 will be similar because silicon (2.8.4) is in the same group as carbon.

All the bonds in the **above examples are single covalent bonds**. Below are three examples 7-9, where there is a **double bond** in the molecule, in order that the atoms have stable Noble Gas outer electron arrangements around each atom. Carbon has a valency of 4.

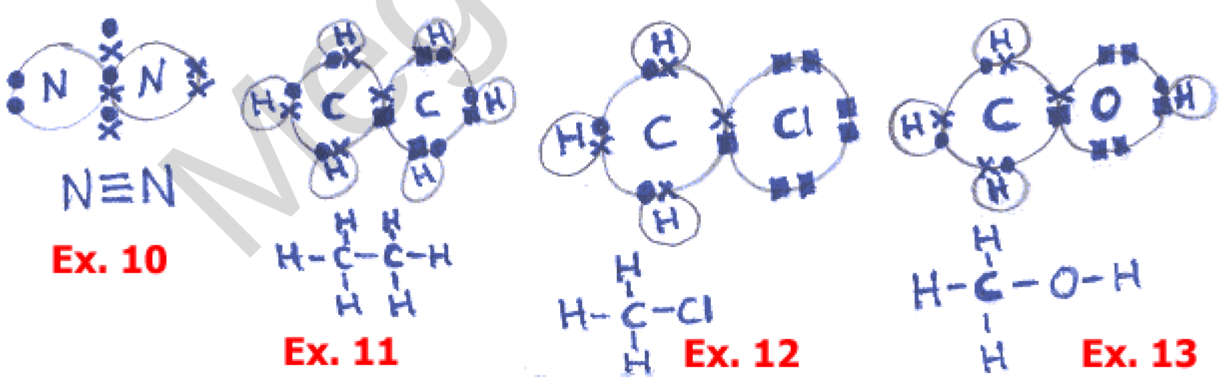
Example 7:  Two atoms of oxygen (2.6) combine to form the molecules of the element **oxygen O_2.**

The molecule has one O=O double covalent bond **O=O**. Oxygen valency 2.

Example 8:  One atom of carbon (2.4) combines with two atoms of oxygen (2.6) to form **carbon dioxide CO_2.** The molecule can be shown as **O=C=O** with two carbon = oxygen double covalent bonds Valencies of C and O are 4 and 2 respectively.

Example 9:  Two atoms of carbon (2.4) combine with four atoms of hydrogen (1) to form **ethene C_2H_4.** The molecule can be shown as  with one carbon = carbon double bond and four carbon - hydrogen single covalent bonds The valency of carbon is still 4.

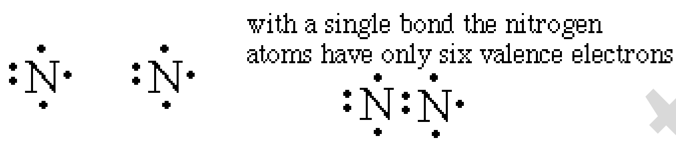
Examples 10-13: The scribbles below illustrate some more complex examples. Ex. 10 nitrogen; Ex. 11 ethane; Ex. 12 chloromethane and Ex. 13 methanol. The valencies or combining power in theses examples are N 3, H 1, C 4, Cl 1 and O 2.



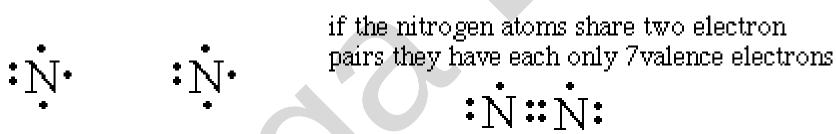
MULTIPLE BONDS

Why multiple bonds form: Nonmetal atoms bond to reach a stable or low energy condition. This happens when a main-group atom shares enough electrons to achieve a rare gas valence shell. Sometimes the number of electrons needed cannot be provided by sharing electrons simply in single bonds(pairs).

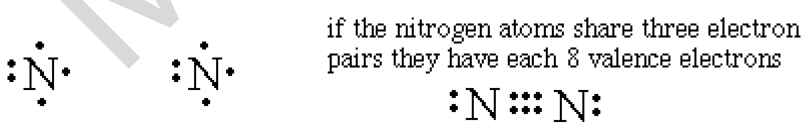
Nitrogen molecules are an example of this situation. Nitrogen atoms have five valence electrons. We know that molecules of N_2 exist based on a great deal of measurements. If two nitrogen atoms simply formed a single bond the dot structure would look like the illustraton below. Each nitrogen atom would have only six electrons not an octet. The single bond doesn't provide the two nitrogen atoms with an octet. If the octet rule is to be followed, the nitrogen atoms must share more than two electrons.



The trial and error method is used to decide how many shared electrons are needed to create a structure where the octet rule is met. Since one bond didn't work the next thing to try is a double bond where the atoms share four electrons. Unfortunately the double bond structure only provides 7 valence shell electrons not eight.



Because the single and double bonds didn't do the trick the next thing to try is a triple bond where the nitrogen atoms share six electrons. The count for both atoms when a triple bond is used in the structure shows that the octet rule is met.

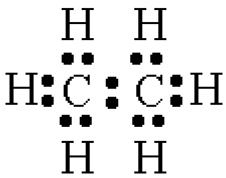
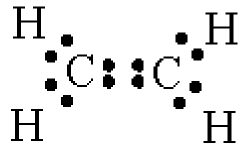
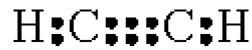


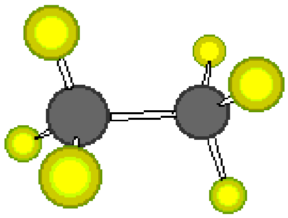
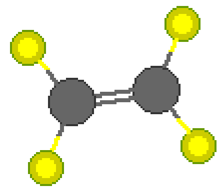
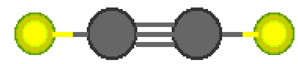
Multiple bonds are very common in carbon compounds. The carbon atom can form four bonds. These four bonds can be all single,CH_3CH_3, two single and one double, CH_2CH_2, two double, one single and one triple, CHCH.

Ethane has a single bond carbon-carbon bond

Ethene has a double carbon-carbon bond

Acetylene has a triple carbon-carbon bond

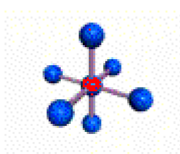
The carbon-carbon single bond has a bond length of 154 picometers.

The carbon-carbon double bond has a bond length of 133 picometers.

The carbon-carbon triple bond has a bond length of 120 picometers.

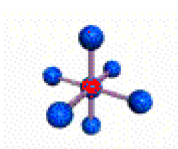
The carbon-carbon bond lengths decrease as the number of shared electrons and bonds increase. This is reasonable because there are more electrons attracted to each nucleus. The repulsions between positive nuclei decrease as more electrons are shared. The relative lengths for the three types of bonds are summarized here. REMEMBER these compare bonds between pairs of atoms like ;

C:::C, C::C, C:C or C:::N, C::N, C:N or N:::N, N::N N:N or N::O , N:O



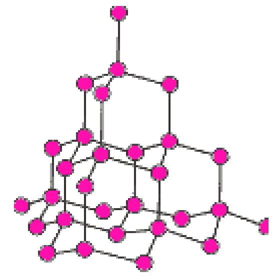
The Properties of simple covalent molecular substances - small molecules!

- The **electrical forces of attraction, that is the chemical bond*, between atoms in any molecule are strong** and most molecules do not change chemically on moderate heating.(*** sometimes referred to as the intramolecular bond**)
- **However, the electrical forces** between molecules are weak and easily weakened further on heating.**
- These weak attractions are known as ****intermolecular forces** and consequently **the bulk material is not usually very strong**.
- Consequently **small covalent molecules tend to be volatile liquids, easily vapourised, or low melting point solids**.
- On heating the inter-molecular forces are easily overcome with the increased kinetic energy gain of the particles and so have **low melting and boiling points**.
- They are also **poor conductors of electricity** because there are **no free electrons or ions** in any state to carry electric charge.
- **Most small molecules will dissolve in a solvent to form a solution**.

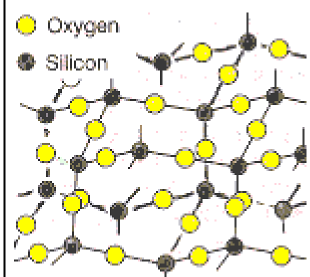


Large Covalent Molecules and their Properties

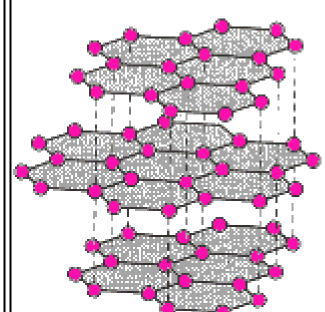
- It is possible for **many atoms** to link up to form a **giant covalent structure or lattice**. The atoms are usually non-metals.
- This produces a **very strong 3-dimensional covalent bond network or lattice**.
- **This gives them significantly different properties from the small simple covalent molecules mentioned above.**
- This is **illustrated by carbon in the form of diamond**. Carbon can form four single bonds so each carbon bonds to four others etc.
- This type of structure is thermally very stable and they have **high melting and boiling points**.
- They are usually **poor conductors of electricity** because the electrons are not usually free to move as they can in metallic structures.
- Also because of the strength of the bonding in all directions in the structure, they are often very **hard, strong** and **will not dissolve** in solvents like water.

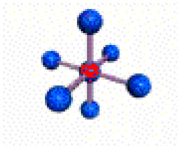


- **Silicon dioxide (silica, SiO_2)** has a similar 3D structure and properties, shown below diamond.
- **The hardness of diamond enables it to be used as the 'leading edge' on cutting tools**.
- Diamond is an **allotrope** of carbon. Two other allotropes of diamond are described below. **Allotropes are different forms of the same element in the same physical state** (**3 solid forms of carbon are described here**).



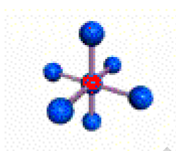
- **Carbon also occurs in the form of graphite**. The carbon atoms form joined hexagonal rings forming layers 1 atom thick.
- **There are three strong covalent bonds per carbon**, BUT, the fourth bond carbon can form from its four outer electrons, is shared between the three bonds shown (this requires advanced level concepts to fully explain, and this bonding situation also occurs in fullerenes described below).
- **The layers are only held together by weak intermolecular forces** shown by the dotted lines NOT by strong covalent bonds.
- Like diamond and silica (above) the large molecules of the layer ensure graphite **has typically very high melting point because of the strong 2D bonding network** (note: NOT 3D network)**.**
- **Graphite will not dissolve in solvents** because of the strong bonding
- BUT there are **two crucial differences compared to diamond** ...
  - **Electrons, from the 'shared bond', can move freely through each layer, so graphite is a conductor like a metal** (diamond is an electrical insulator and a poor heat conductor). Graphite is used in electrical contacts eg electrodes in electrolysis.
  - **The weak forces enable the layers to slip over each other** so where as diamond is hard material **graphite is a 'soft' crystal**, it feels slippery. **Graphite is used as a lubricant**.





Bonding in polymers and 1-3 'dimension' concepts in macromolecules

- The bonding in polymers or plastics is no different in principle to the examples described above, but there is quite a range of properties and the difference between simple covalent and giant covalent molecules can get a bit 'blurred'.
  - Bonds between atoms in molecules, eg **C-C**, are called **intra-molecular bonds**.
  - The much weaker electrical attractions between individual molecules are called **inter-molecular forces**.
- In **thermosoftening plastics like poly(ethene)** the bonding is like ethane except there are lots of carbon atoms linked together to form long chains. They are moderately strong materials but tend to soften on heating and are not usually very soluble in solvents. **The structure is basically a linear 1 dimensional strong bonding networks**.
- **Graphite structure is a layered 2 dimensional strong bond network** made of **layers of joined hexagonal rings of carbon atoms** with weak inter-molecular forces between the layers.
- **Thermosetting plastic structures like melamine** have a **3 dimensional cross-linked giant covalent structure network** similar to diamond or silica in principle, but rather more complex and chaotic! Because of the strong 3D covalent bond network they do not dissolve in any solvents and do not soften on heating and are much stronger than thermoplastics.
- **More on polymers** in **Oil Notes** and **Extra Organic Chemistry Notes**.



3. Ionic Bonding

Ionic bonds are formed by one atom **transferring electrons** to another atom to form **ions**. **Ions** are atoms, or groups of atoms, which have lost or gained electrons.

The atom **losing electrons** forms a **positive ion (a cation)** and is usually a **metal**. The overall charge on the ion is positive due to excess positive nuclear charge (protons do NOT change in chemical reactions).

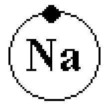
The atom **gaining electrons** forms a **negative ion (an anion)** and is usually a **non-metallic element**. The overall charge on the ion is negative because of the gain, and therefore excess, of negative electrons.

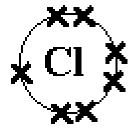
The examples below combining a metal from Groups 1 (Alkali Metals), 2 or 3, with a non-metal from Group 6 or Group 7 (The Halogens)

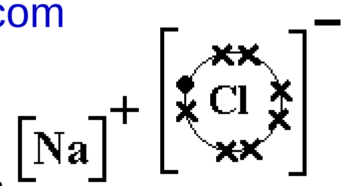
Example 1: A Group 1 metal + a Group 7 non-metal eg sodium + chlorine ➔ **sodium chloride NaCl** or ionic formula $\mathbf{Na^+Cl^-}$ In terms of electron arrangement, the sodium donates its outer electron to a chlorine atom forming a single positive sodium ion and a single negative chloride ion. The atoms have become stable ions, because electronically, sodium becomes like neon and chlorine like argon.

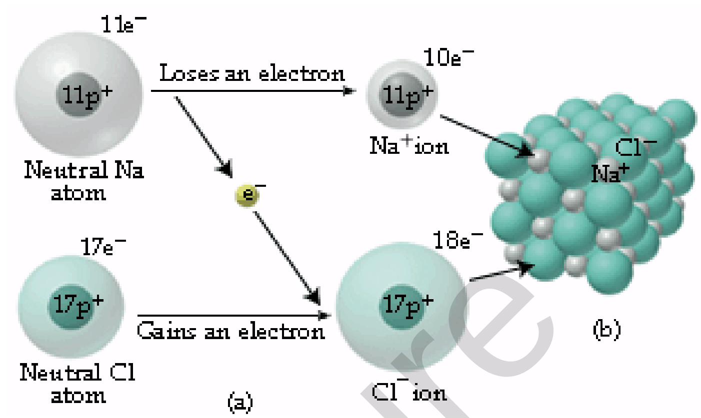
Na (2.8.1) + **Cl** (2.8.7) ➔ $\mathbf{Na^+}$ (2.8) $\mathbf{Cl^-}$ (2.8.8)

can be summarised electronically as [2,8,1] + [2,8,7] ➔ $[2,8]^+$ $[2,8,8]^-$

ONE 

combines with ONE 

to form 

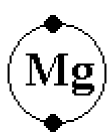


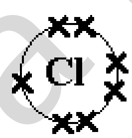
The valencies of Na and Cl are both 1, that is, the numerical charge on the ions. NaF, KBr, LiI etc. will all be electronically similar.

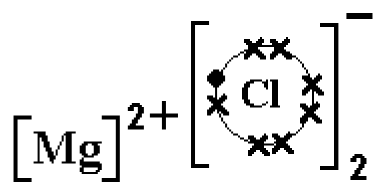
Example 2: A Group 2 metal + a Group 7 non-metal eg magnesium + chlorine ➔ **magnesium chloride $MgCl_2$** or ionic formula **$Mg^{2+}(Cl^-)_2$** In terms of electron arrangement, the magnesium donates its two outer electrons to two chlorine atoms forming a double positive magnesium ion and two single negative chloride ions. The atoms have become stable ions, because electronically, magnesium becomes like neon and chlorine like argon.

Mg (2.8.2) + 2**Cl** (2.8.7) ➔ **Mg^{2+}** (2.8) 2**Cl^-** (2.8.8)

can be summarised electronically as [2,8,2] + 2[2,8,7] ➔ $[2,8]^{2+}$ $[2,8,8]^-_2$

ONE 

combines with TWO 

to form 

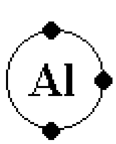
see *

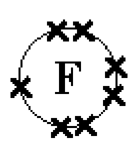
* **NOTE** you can draw two separate chloride ions, but in these examples a number subscript has been used, as in ordinary chemical formula.

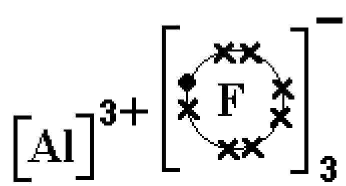
Example 3: A Group 3 metal + a Group 7 non-metal eg aluminium + fluorine ➔ **aluminium fluoride AlF_3** or ionic formula **$Al^{3+}(F^-)_3$** In terms of electron arrangement, the aluminium donates its three outer electrons to three fluorine atoms forming a triple positive aluminium ion and three single negative fluoride ions. The atoms have become stable ions, because electronically, aluminium and fluorine becomes electronically like neon. Valency of Al is, F is 1.

Al (2.8.3) + **3F** (2.7) ➔ **Al^{3+}** (2.8) 3**F^-** (2.8)

can be summarised electronically as [2,8,3] + 3[2,7] ➔ $[2,8]^{3+}$ $[2,8]^-_3$

ONE 

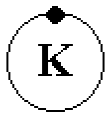
combines with THREE 

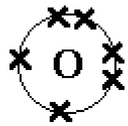
to form 

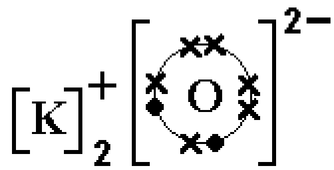
Example 4: A Group 1 metal + a Group 6 non-metal eg potassium + oxygen ➔ **potassium oxide K_2O** or ionic formula **$(K^+)_2O^{2-}$** In terms of electron arrangement, the two potassium atoms donates their outer electrons to one oxygen atom. This results in two single positive potassium ions to one double negative oxide ion. All the ions have the stable electronic structures 2.8.8 (argon like) or 2.8 (neon like). Valencies, K 1, oxygen 2. Na_2O, Na_2S, K_2S etc. will be similar.

2K (2.8.8.1) + **O** (2.6) ➔ 2**K^+** (2.8.8) **O^{2-}** (2.8)

can be summarised electronically as 2[2,8,8,1] + [2,6] ➔ $[2,8,8]^{+}_{2}$ $[2,8]^{2-}$

TWO 

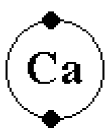
combine with ONE 

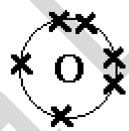
to form 

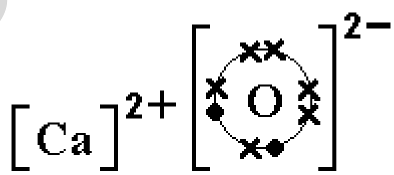
Example 5: A Group 2 metal + a Group 6 non-metal eg calcium + oxygen ➔ **calcium oxide CaO** or ionic formula **$Ca^{2+}O^{2-}$** In terms of electron arrangement, one calcium atom donates its two outer electrons to one oxygen atom. This results in a double positive calcium ion to one double negative oxide ion. All the ions have the stable electronic structures 2.8.8 (argon like) or 2.8 (neon like). the valency of both calcium and oxygen is 2. MgO, MgS, or CaS will be similar electronically (S and O both in Group 6)

Ca (2.8.8.2) + **O** (2.6) ➔ **Ca^{2+}** (2.8.8) **O^{2-}** (2.8)

can be summarised electronically as [2,8,8,2] + [2,6] ➔ $[2,8,8]^{2+}$ $[2,8]^{2-}$

ONE 

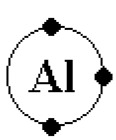
combines with ONE 

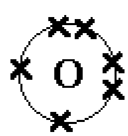
to form 

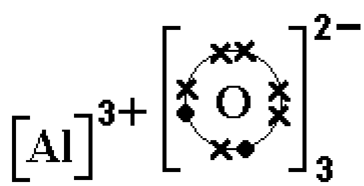
Example 6: A Group 3 metal + a Group 6 non-metal eg aluminium + oxygen ➔ **aluminium oxide Al_2O_3** or ionic formula **$(Al^{3+})_2(O^{2-})_3$** In terms of electron arrangement, two aluminium atoms donate their three outer electrons to three oxygen atoms. This results in two triple positive aluminium ions to three double negative oxide ions. All the ions have the stable electronic structure of neon 2.8. Valencies, Al 3 and O 2.

2Al (2.8.3) + 3**O** (2.6) ➔ 2**Al^{3+}** (2.8) 3**O^{2-}** (2.8)

can be summarised electronically as 2[2,8,3] + 3[2,6] ➔ $[2,8]^{3+}_{2}$ $[2,8]^{2-}_{3}$

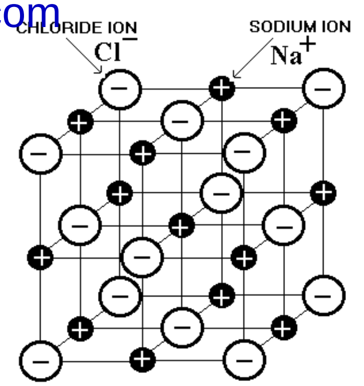
TWO 

combines with THREE 

to form 

The Properties of Ionic Compounds

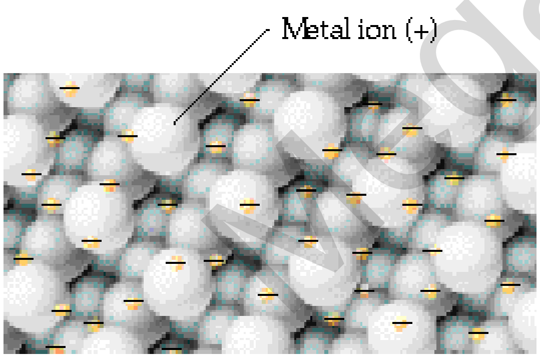
- The **alternate positive and negative ions in an ionic solid** are arranged in an orderly way in a **giant ionic lattice structure** shown on the left.
- **The ionic bond is the strong electrical attraction between the positive and negative ions** next to each other in the lattice.
- The **bonding extends throughout the crystal** in all directions.
- **Salts and metal oxides** are typical ionic compounds.
- This strong bonding force makes the structure hard (if brittle) and have **high melting and boiling points**, so they are not very volatile!
- The bigger the charges on the ions the stronger the bonding attraction eg magnesium oxide $Mg^{2+}O^{2-}$ has a higher melting point than sodium chloride $Na^{+}Cl^{-}$.
- Unlike covalent molecules, **ALL ionic compounds are crystalline solids** at room temperature.
- **They are hard but brittle**, when stressed the bonds are broken along planes of ions which shear away. They are NOT malleable like metals.
- **Many ionic compounds are soluble in water**, but not all, so don't make assumptions.
- The **solid crystals DO NOT conduct electricity** because the ions are not free to move to carry an electric current. However, if the ionic compound is **melted** or **dissolved in water**, the liquid will now **conduct electricity,** as the ion particles are now free.



SODIUM CHLORIDE LATTICE STRUCTURE

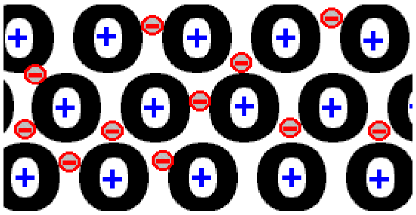
4. BONDING IN METALS

Electron-Sea Model of Metals



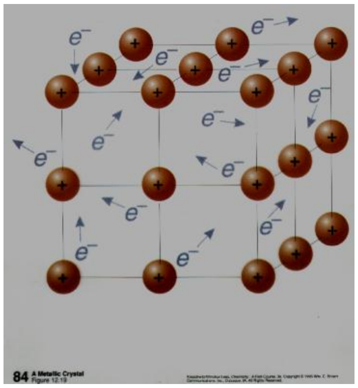
In the **electron-sea model**, a metal crystal is considered to be a three-dimensional array of metal cations immersed in a sea of valence electrons. The delocalized valence electrons are free to move throughout the crystal and are not associated with any one particular metal cation. The mobility of the electrons accounts for the high electrical conductivity of metals. Thermal conductivity can also be ascribed to the mobile electrons that conduct heat by carrying kinetic energy from one part of the crystal to another. Three-dimensional delocalized bonding allows the metal to be both malleable and ductile.

The **crystal lattice of metals consists of ions** NOT atoms surrounded by a '**sea of electrons**' forming another type of **giant lattice**.



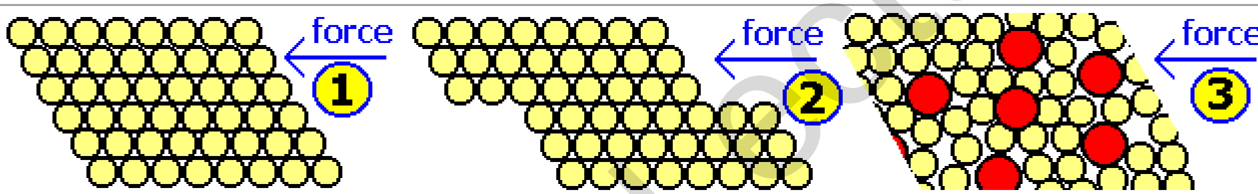
- The **outer electrons (-)** from the original metal atoms are free to move around between the positive metal ions formed **(+)**.
- These free or 'delocalised' electrons are the 'electronic glue' holding the particles together.
- There is a **strong electrical force of attraction**

between these **mobile electrons (-)** and the immobile **positive metal ions (+)** and this is the **metallic bond**.



- This **strong bonding** generally results in **dense**, **strong materials with high melting and boiling points**.
- Metals are **good conductors of electricity** because these 'free' electrons carry the charge of an electric current when a potential difference (voltage!) is applied across a piece of metal.
- Metals are also **good conductors of heat**. This is also due to the free moving electrons. Non-metallic solids conduct heat energy by hotter more strongly vibrating atoms, knocking against cooler less strongly vibrating atoms to pass the particle kinetic energy on. In metals, as well as this effect, the 'hot' high kinetic energy electrons move around freely to transfer the particle kinetic energy more efficiently to 'cooler' atoms.
- Typical metals also have a **silvery surface** but remember this may be easily tarnished by corrosive oxidation in air and water.
- Unlike ionic solids, **metals are very malleable**, they can be readily bent, pressed or hammered into shape. The layers of atoms can slide over each other without fracturing the structure. The reason for this is the **mobility of the electrons**. When planes of metal atoms are 'bent' or slide the electrons can run in between the atoms and maintain a strong bonding situation. This can't happen in ionic solids.

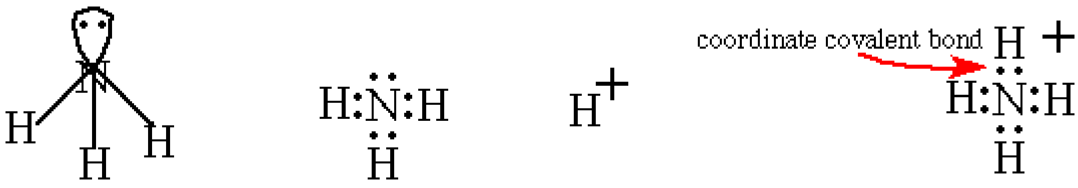
Note on Alloy Structure



1. Shows the regular arrangement of the atoms in a metal crystal and the white spaces show where the free electrons are (yellow circles actually positive metal ions).
2. Shows what happens when the metal is stressed by a strong force. The layers of atoms can slide over each other and the bonding is maintained as the mobile electrons keep in contact with atoms, so the metal remains intact BUT a different shape.
3. Shows an alloy mixture. It is NOT a compound but a physical mixing of a metal plus at least one other material (shown by red circle, it can be another metal eg Ni, a non-metal eg C or a compound of carbon or manganese, and it can be bigger or smaller than iron atoms). Many alloys are produced to give a stronger metal. The presence of the other atoms (smaller or bigger) disrupts the symmetry of the layers and reduces the 'slip ability' of one layer next to another. The result is a stronger harder less malleable metal.

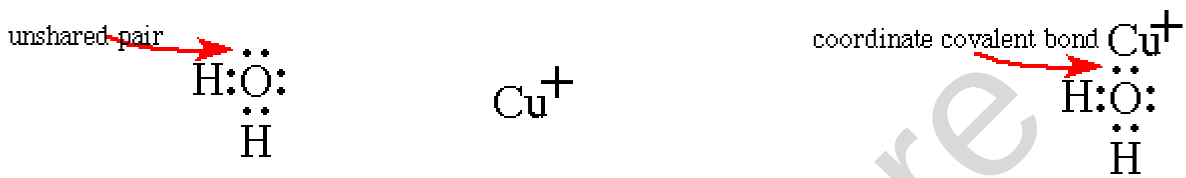
5.Coordinate Covalent Bonds(Dative Bonding)

A coordinate covalent bond is special because it involves a shared pair of electrons that came from a single atom. Ammonia had a nitrogen atom with an unshared pair of electrons. These can be shared with an electron defficient atom like H^+.



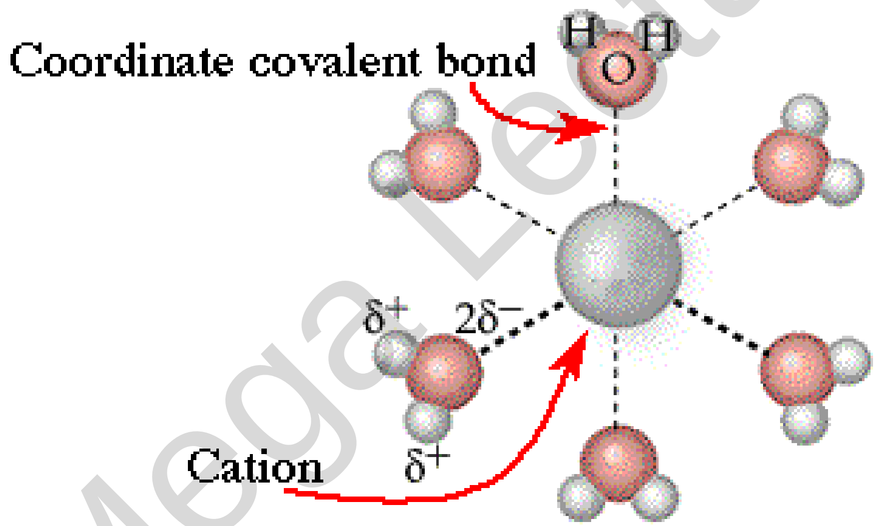
ammonia ammonia proton ammonium ion

Water molecules have two unshared pairs of electrons. These form coordinate covalent bonds with cations that are dissolved in water. This is one reason why water dissolves many ionic solids. The energy released when the water molecules bond to the cations is often enough to break up the ionic solid.



water cation water and cation

A dissolved cation will form as many as six coordinate covalent bonds.

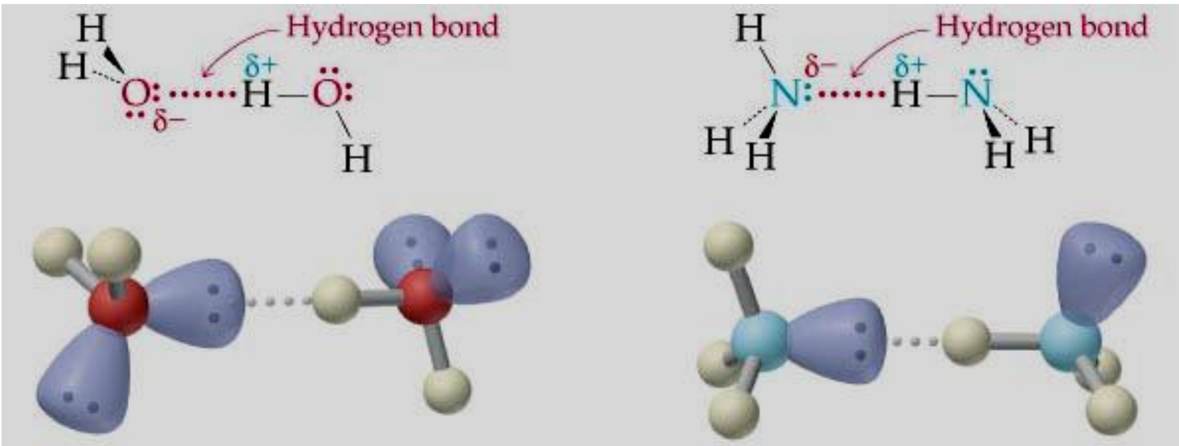


6. INTERMOLECULAR FORCES

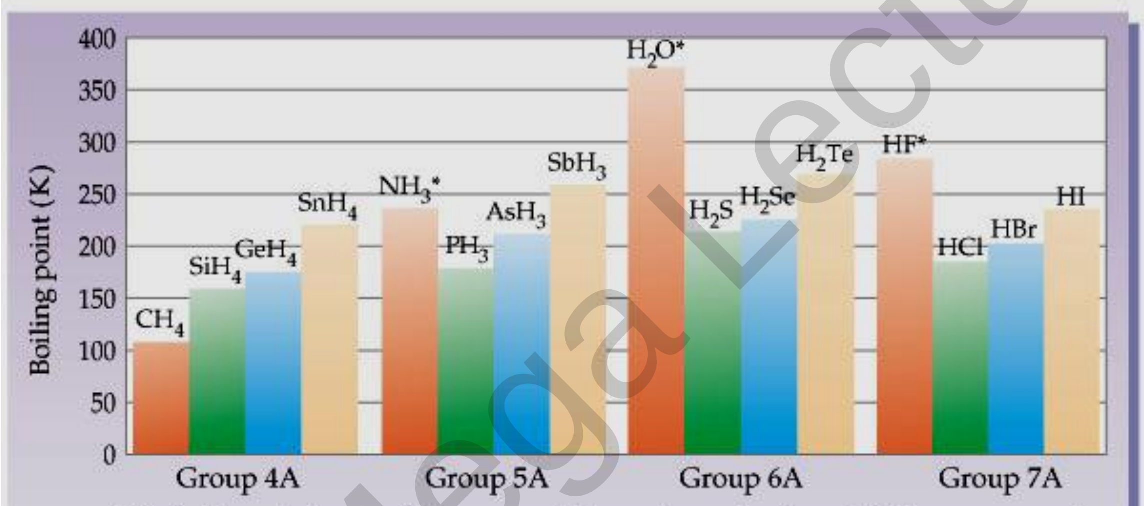
Intermolecular forces are those forces that occur between particles (molecules, atoms, or ions). The strength of these forces at a given temperature dictates whether a substance will have the properties of a solid, a liquid, or a gas. The term van der Waals forces encompasses all types of intermolecular forces. All intermolecular forces arise from electrostatic interactions governed by the basic rule that like charges repel and unlike charges attract.

Hydrogen Bonds

A hydrogen bond is an attractive interaction between a hydrogen that is bonded to a very electronegative atom (O, N, or F) and an unshared electron pair on another electronegative atom.



Hydrogen bonds can be quite strong. Substances that form hydrogen bonds have unusually high boiling points due to the extra energy that must be used to separate the molecules.



*The boiling points generally increase with increasing molecular weight down a group of the periodic table, but the hydrides of nitrogen (NH_3), oxygen (H_2O), and fluorine (HF) have abnormally high boiling points because these molecules have hydrogen bonds.

DONE