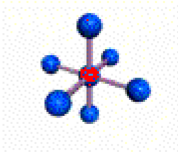
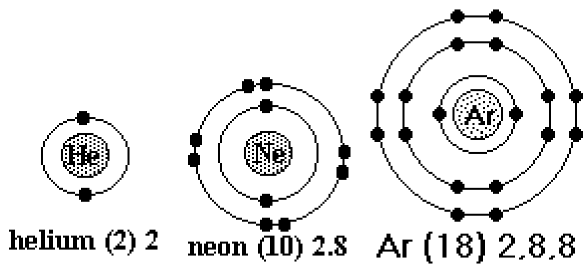
Chemical bonding



1. Why do atoms bond together? - 'electron glue'!

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** e.g. **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.



(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 and atoms can bond in two ways …

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 (see examples later).

COVALENT BONDING - **sharing electrons** to form **molecules** with **covalent bonds**, the bond is usually formed between **two non-metallic elements** **in a molecule**. **The two positive nuclei, due to the protons, of both atoms, are mutually attracted to the shared negative electrons between them.**

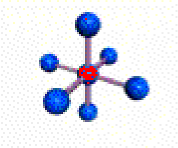
OR **IONIC BONDING** - By one atom **transferring electrons** to another atom.

- **An ion is an atom or group of atoms carrying an overall positive or negative charge**
  - **e.g.** 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 i.e. 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 on the ion, for every electron lost you get a one unit increase in the positive charge on the ion.

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. The ionic bond then consists of the attractive force between the positive and negative ions in the structure.**

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.



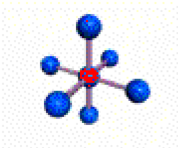
2. Covalent Bonding - electron sharing in big or small molecules!

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 i.e. 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.

Note: In the examples it is assumed you can work out the electron configuration (arrangement in shells or energy levels) given the atomic number from the Periodic Table.



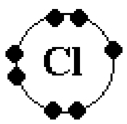
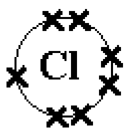
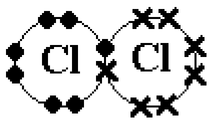
The bonding in Small Covalent Molecules

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 …

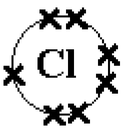
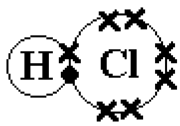
Example 1: two hydrogen atoms (1) form the molecule of the element hydrogen H_2

 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. The nuclei would be a tiny dot in the middle of where the H symbols are drawn! H valency is 1.

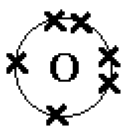
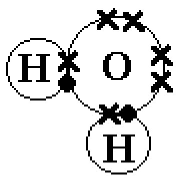
Example 2: two chlorine atoms (2.8.7) form the molecule of the element chlorine Cl_2

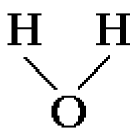
 and  combine to form  where both atoms have a pseudo argon structure of 8 outer electrons around each atom. All the other halogens would be similar eg F_2, Br_2 and $I_{2\ etc}$. Valency of halogens is 1 here.

Example 3: one atom of hydrogen (1) combines with one 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.

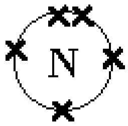
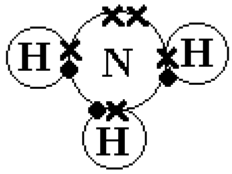
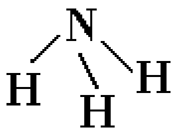
Example 4: two atoms of hydrogen (1) combine with one atom of oxygen (2.6) to form the molecule of the compound we call **water H_2O**

 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



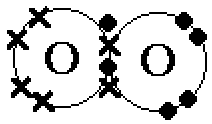
with two hydrogen - oxygen single covalent bonds (AS note: called a V or bent shape, the H-O-H bond angle is 105°). Hydrogen sulphide will be similar, since sulphur (2.8.6) is in the same Group 6 as oxygen. Valency of oxygen and sulphur is 2 here.

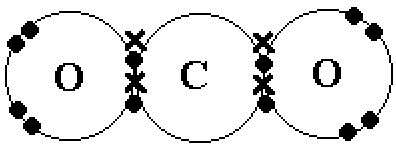
Example 5: three atoms of hydrogen (1) combine with one atom of nitrogen (2.5) to form the molecule of the compound we call **ammonia NH_3**

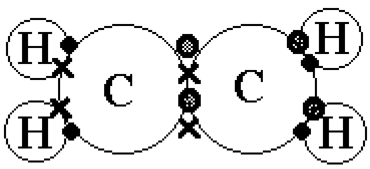
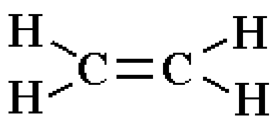
three of  and one  combine to form  so that the hydrogen atoms are electronically like helium and the nitrogen atom becomes like neon. The molecule can be shown as  with three nitrogen - hydrogen single covalent bonds (AS note: called a trigonal pyramid shape, the H-N-H bond angle is 107°). PH_3 will be similar since phosphorus (2.8.5) is in the same Group 5 as nitrogen. Valency of nitrogen or phosphorus is 3 here.

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.

More complex examples can be worked out e.g. involving C, H and O. In each case link in the atoms so that there are 2 around a H (electronically like He), or 8 around the C or O (electronically like Ne).

Example 6:  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 7:  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 (AS note: called a **linear** shape, the O=C=O bond angle is 180°). Valencies of C and O are 4 and 2 respectively.

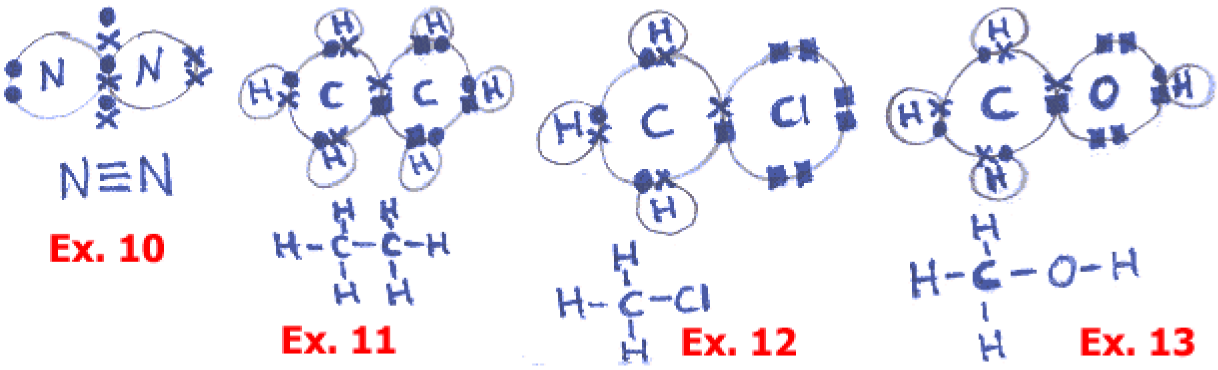
Example 8:  Two atoms of carbon (2.4) combine with four atoms of hydrogen (1) to form **ethene** C_2H_4. The molecule can be shown as  with one carbon = carbon double bond and four carbon - hydrogen single covalent bonds (called a **planar** shape, its completely flat!, the H-C=C and H-C-H bond angles are 120°). The valency of carbon is still 4.

Examples 10-13:

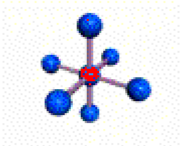
The scribbles below illustrate some more complex examples. Can you deduce them for yourself?

Ex. 10 **nitrogen** N_2; Ex. 11 **ethane** C_2H_6; Ex. 12 **chloromethane** CH_3Cl and Ex. 13 **methanol** CH_3OH.

Electronic origin of the diagrams showing the outer electrons of N, C, Cl and O: N at. no. 7 (2.5); H at. no. (1); C at. no. 6 (2.4); Cl at. no. 17 (2.8.7) and O at. no. 8 (2.6) plus a variety of crosses and blobs! The valencies or combining power in theses examples are N 3, H 1, C 4, Cl 1 and O 2.



- **AS advanced level notes on shapes and bond angles**:
  - Ex. 11 Ethane has a linked double tetrahedral shape, all H-C-H and H-C-C bond angles are 109°
  - Ex. 12 chloromethane has tetrahedral shape with H-C-H and H-C-Cl bond angles of approximately 109°
  - Ex. 13 methanol, the four bonds around the the central carbon are tetrahedrally arranged with a H 'wiggle' on the oxygen. All the H-C-H, H-C-O and C-O-H bond angles are approximately 109°
  - The blue icon eg below, represents an octahedral shape (eg SF_6, complex transition metal ions like $[Cu(H_2O)_6]^{2+}$ and the bond angles are either 90° or 180°
  - Simple molecules with a triple bond are often linear eg **H-C≡C-H ethyne** or **H-C≡N hydrogen cyanide** (methanenitrile)

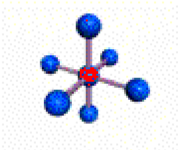


Typical properties of simple covalent substances - small molecules!

- 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 intra-molecular 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.

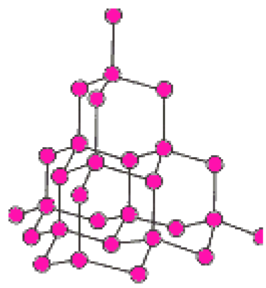
- On heating the inter-molecular forces are easily overcome with the increased kinetic energy gain of the particles and so have **low melting and boiling points**.
- They are also **poor conductors of electricity** because there are **no free electrons or ions** in any state to carry electric charge.
- **Most small molecules will dissolve in a solvent to form a solution.**
- The **properties of these simple small molecules should be compared and contrasted with those molecules of a giant covalent nature** (next section). Apart from points on the strong bonds between the atoms in the molecule and the lack of electrical conduction, all the other properties are significantly different!

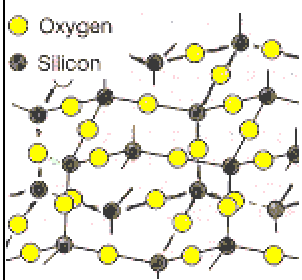
BIG!  **Large Covalent Molecules and their Properties**

(macromolecules - giant covalent networks and polymers)

The structure of the three allotropes of carbon (diamond, graphite and fullerenes), silicon and silicon dioxide (silica)	**DIAGRAMS**

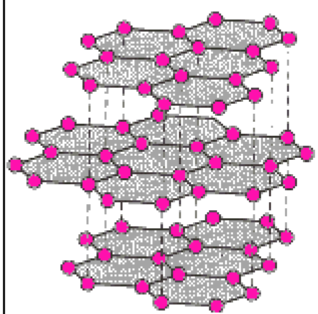
- It is possible for **many atoms** to link up to form a **giant covalent structure or lattice**. The atoms are usually non-metals.
- This produces a **very strong 3-dimensional covalent bond network or lattice**.
- **This gives them significantly different properties from the small simple covalent molecules mentioned above.**
- This is **illustrated by carbon in the form of diamond** (an **allotrope** of carbon). Carbon has four outer electrons that form four single bonds, so each carbon bonds to four others by electron pairing/sharing. **Pure silicon**, another element in Group 4, **has a similar structure**.
    - NOTE: **Allotropes** are different forms of the same element in the same physical state. They occur due to different bonding arrangements and so **diamond**, **graphite** (below) and **fullerenes** (below) are the **three solid allotropes of the element carbon**.
        - Oxygen (dioxygen), O_2, and ozone (trioxygen), O_3, are the two small gaseous allotrope molecules of the element oxygen.
        - Sulphur has three solid allotropes, two different crystalline forms based on small S_8 molecules called rhombic and monoclinic sulphur and a 3rd form of long chain (-S-S-S- etc.) molecules called plastic sulphur.
- **TYPICAL PROPERTIES of GIANT COVALENT STRUCTURES**

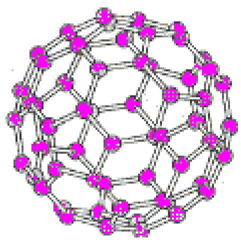


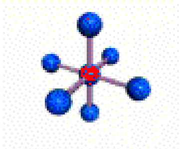


- This type of giant covalent structure is thermally very stable and has a very **high melting and boiling points**.
- They are usually **poor conductors of electricity** because the electrons are not usually free to move as they can in metallic structures.
- Also because of the strength of the bonding 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**.

- **Carbon also occurs in the form of graphite**. The carbon atoms form joined hexagonal rings forming layers 1 atom thick. It is NOT considered a giant covalent structure.
- **There are three strong covalent bonds per carbon** (3 C-C bonds in a planar arrangement from 3 of its 4 outer electrons), BUT, the fourth outer electron is '**delocalised**' or shared between the carbon atoms to form the equivalent of a 4th bond per carbon atom (this situation requires advanced level concepts to fully explain, and this bonding situation also occurs in fullerenes described below, and in aromatic compounds you deal with at advanced level).
- **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**.
- These two different characteristics described above are put to a common use with the electrical contacts in

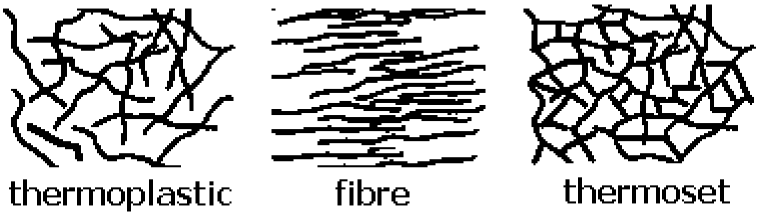


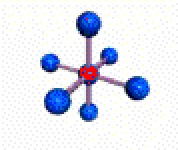
electric motors and dynamos. These contacts (called brushes) are made of graphite sprung onto the spinning brass contacts of the armature. The graphite brushes provide good electrical contact and are self-lubricating as the carbon layers slide over each other.	
• A **3rd form of carbon are fullerenes or 'bucky balls'**! It consists of hexagonal rings like graphite and alternating pentagonal rings to allow curvature of the surface. • **Buckminster Fullerene C_{60}** is shown and the bonds form a pattern like a soccer ball. Others are oval shaped like a rugby ball. • They are **NOT** giant covalent structures and are classed as simple! molecules. They do dissolve in solvents, and although solid, their melting points are not that high. • **They are mentioned here to illustrate the different forms of carbon** AND they are trying to make them as continuous tubes to form very strong fibres of 'pipe like' molecules called 'nanotubes'.	



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, e.g. **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.





3. Ionic Bonding - electron transfer

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).

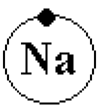
The atom **gaining electrons** forms a **negative ion (an anion)** and is usually a **non-metallic element**. The overall charge on the ion is negative because of the gain, and therefore excess, of negative electrons.

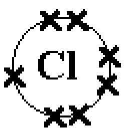
The examples below combining a metal from Groups 1 (Alkali Metals), 2 or 3, with a non-metal from Group 6 or Group 7 (The Halogens)

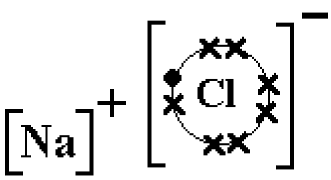
Example 1: A Group 1 metal + a Group 7 non-metal eg sodium + chlorine ==> **sodium chloride NaCl** or ionic formula **Na^+Cl^-** In terms of electron arrangement, the sodium donates its outer electron to a chlorine atom forming a single positive sodium ion and a single negative chloride ion. The atoms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

Na (2.8.1) + **Cl** (2.8.7) ==> **Na^+** (2.8) **Cl^-** (2.8.8)

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

ONE 

combines with ONE 

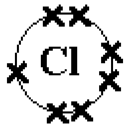
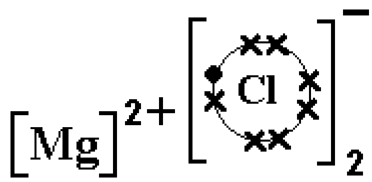
to form 

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 e.g. 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] ==> $[2,8]^{2+}$ $[2,8,8]^{-}_{2}$

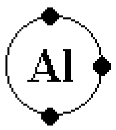
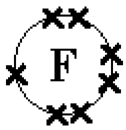
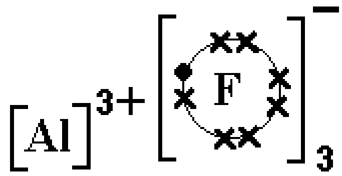
ONE  combines with TWO  to form  **see ***

*** NOTE** you can draw two separate chloride ions, but in these examples a number subscript has been used, as in ordinary chemical formula. The valency of Mg is 2 and chlorine 1, i.e. the numerical charges of the ions. BeF_2, $MgBr_2$, $CaCl_2$ or CaI_2 etc. will all be electronically similar.

Example 3: *A Group 3 metal + a Group 7 non-metal e.g.* 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)

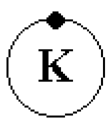
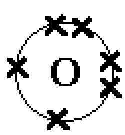
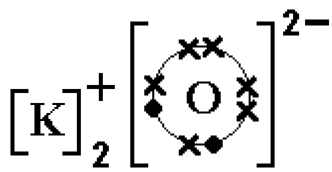
can be summarised electronically as [2,8,3] + 3[2,7] ==> $[2,8]^{3+}$ $[2,8]^{-}_{3}$

ONE  combines with THREE  to form 

Example 4: A Group 1 metal + a Group 6 non-metal e.g. 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)

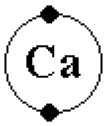
can be summarised electronically as 2[2,8,8,1] + [2,6] ==> $[2,8,8]^{+}_{2}$ $[2,8]^{2-}$

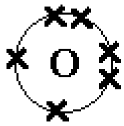
TWO  combine with ONE  to form 

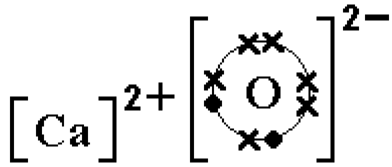
Example 5: A Group 2 metal + a Group 6 non-metal e.g. calcium + oxygen ==> **calcium oxide CaO** or ionic formula $Ca^{2+}O^{2-}$ In terms of electron arrangement, one calcium atom donates its two outer electrons to one oxygen atom. This results in a double positive calcium ion to one double negative 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 

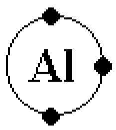
combines with ONE 

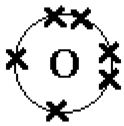
to form 

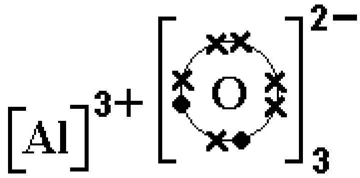
Example 6: A Group 3 metal + a Group 6 non-metal e.g. aluminium + oxygen ==> **aluminium oxide** **Al_2O_3** or ionic formula **$(Al^{3+})_2(O^{2-})_3$** In terms of electron arrangement, two aluminium atoms donate their three outer electrons to three oxygen atoms. This results in two triple positive 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.

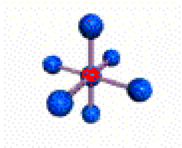
2Al (2.8.3) + 3**O** (2.6) ==> 2**Al^{3+}** (2.8) 3**O^{2-}** (2.8)

can be summarised electronically as 2[2,8,3] + 3[2,6] ==> $[2,8]^{3+}{}_2$ $[2,8]^{2-}{}_3$

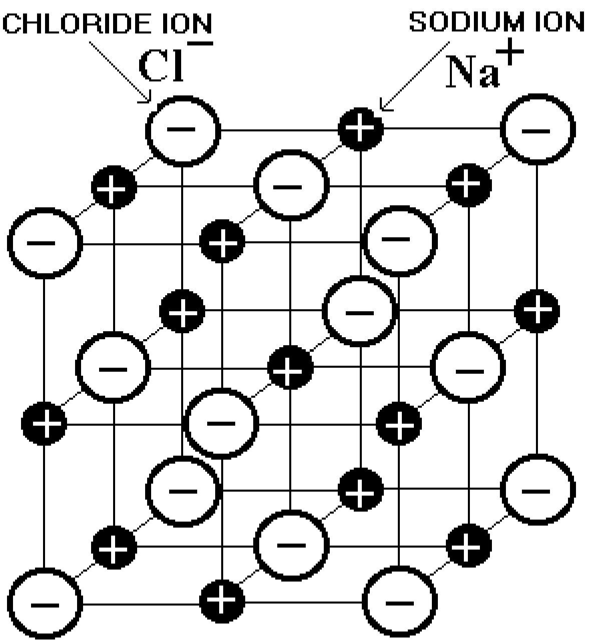
TWO 

combines with THREE 

to form 

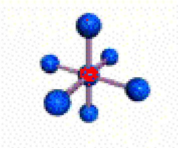


The properties of Ionic Compounds

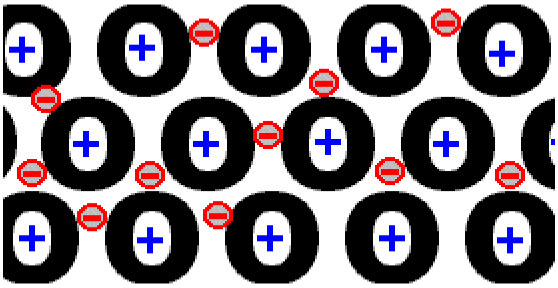


SODIUM CHLORIDE LATTICE STRUCTURE

- 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 e.g. magnesium oxide $\mathbf{Mg^{2+}O^{2-}}$ has a higher melting point than sodium chloride $\mathbf{Na^{+}Cl^{-}}$.
- 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 (see below).
- **Many ionic compounds are soluble in water,** but not all, so don't make assumptions. Salts can dissolve in water because the ions can separate and become surrounded by water molecules which weakly bond to the ions. This reduces the attractive forces between the ions, preventing the crystal structure to exist. Evaporating the water from a salt solution will eventually allow the ionic crystal lattice to reform.
- 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.



4. BONDING IN METALS

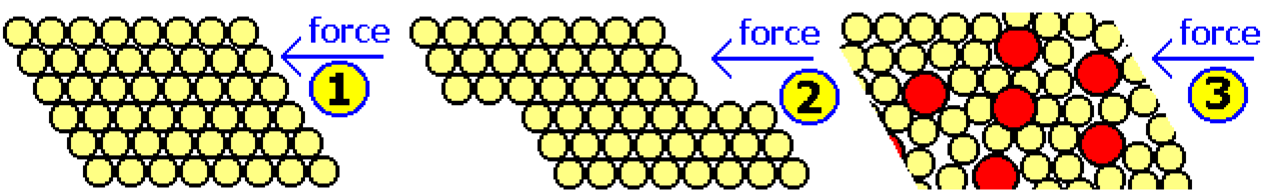


- The **crystal lattice of metals consists of ions** NOT atoms surrounded by a '**sea of electrons**' forming another type of **giant lattice**.
- The **outer electrons (-)** from the original metal atoms are free to move around between the positive metal ions formed **(+)**.
- These free or 'delocalised' 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**.

The physical properties of metals:

- 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 (see below). 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.
- For more on the properties and uses of metals see **Transition Metals** and **Extra Industrial Chemistry** pages and the note and diagram below.

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.