BioMedical Admissions Test (BMAT)

Section 2: Chemistry

Topic C6: Chemical Bonding, Structure, and Properties

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Ionic Bonding

Ionic bonding occurs between a **metal** and a **non-metal**.

- ➔ **Ions** are formed because elements always try to have a full outer shell of electrons; this is their most stable form.
- ➔ The ions are then **attracted to each other** to form an **ionic compound.**
- ➔ Atoms are **more stable when they have a full outer shell of electrons.**

In a reaction between a metal and non-metal:

- The **metal loses electrons**. This means that it now has more protons than electrons and therefore is a **positive ion**.
- The **non-metal receives electrons**. It then has more electrons than protons and therefore becomes a **negative ion**.
- The positive metal ion and the negative non-metal ion are then attracted to each other and form an ionic compound.

The metals e.g. group 1 or group 2 metals, lose their electrons because their outer shell only has 1 or 2 electrons in it.

- ➔ It is easier for the metal to lose these electrons than to gain an addition 6 or 7 electrons to reach the stability of a full outer shell.

The non-metals, e.g. group 16 or 17 non-metals, gain electrons because their outer shells are nearly full.

- ➔ It is easier for the non-metal to gain an extra 1 or 2 electrons than to lose 6 or 7 electrons to reach the stability of a full outer shell.

The number of electrons gained or lost by the atom determines the charge of its ion.

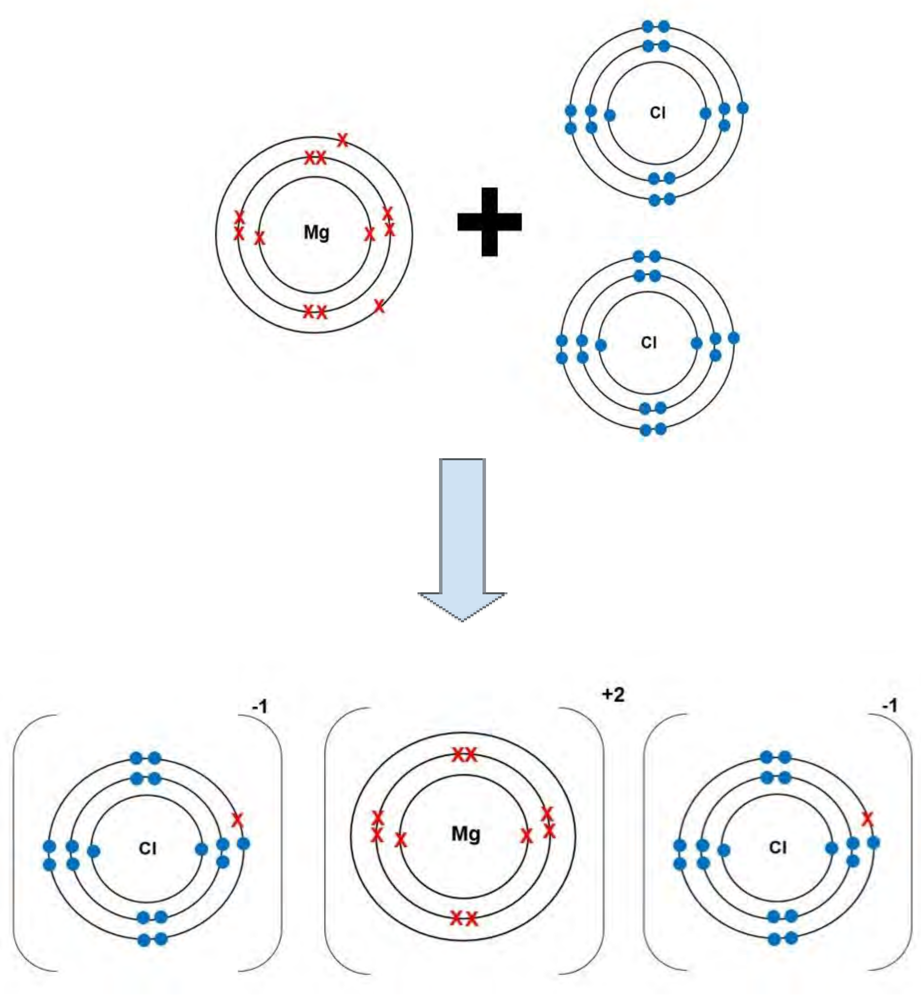
- Group 1 metals have 1 electron in their outer shell. They lose this to become a positive ion with a +1 charge.
- Group 2 metals lose 2 electrons and so have a +2 charge on their ion
- Group 17 metals (halogens) gain one electron and so have a -1 charge on their ion.

Worked Example

Both magnesium and chlorine have a partially filled outer shell of electrons:

- ➔ Magnesium has 2 electrons in its outer shell; it will lose these 2 electrons to become a +2 ion with a full outer shell of electrons.
- ➔ Chlorine has 7 electrons in its outer shell; it will gain 1 electron to become a -1 ion with a fuller outer shell of electrons.
- ➔ Because a chloride ion only has a -1 charge, 2 chloride ions are needed to balance out the +2 charge of the magnesium ion.

Therefore the formula of magnesium chloride is **$MgCl_2$**



Exam tip: When working out the formula of an ionic compound such as $MgCl_2$, remember that the overall charge on the compound must be 0. This means all the negative charges on ions in the compounds must balance all the positive charges on ions in the compounds.

Properties of Ionic Compounds

The **cations** (positive ions) and **anions** (negative ions) in ionic compounds such as NaCl are held together in an **ionic lattice.**

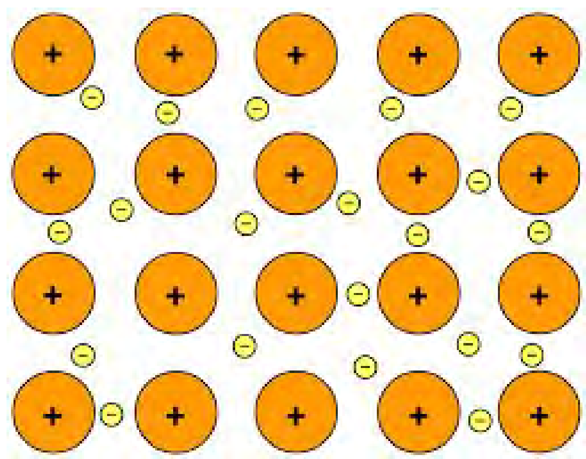
➔ There are **strong forces of electrostatic attraction** between the anions and cations which hold the ions together strongly in the regular lattice arrangement.
➔ Ionic compounds **conduct electricity** when **dissolved in solution** or **molten**, because the ions are free to move and carry a current.
➔ Ionic compounds have **high melting points** and **high boiling points** because the strong forces of electrostatic attraction between the ions of opposite charge require large amounts of energy to overcome.
➔ Ionic compounds **dissolve in water,** as it is **polar** and so the ions are attracted to it
➔ Thus the substances can then be **electrolysed**.

Metallic Bonding

Metallic bonding takes place **between metal atoms.**

Metal atoms are all arranged **regularly** in a **metallic lattice**. The **electrons in the outer shell of the metal atoms are shared by all the ions**; they are said to be **delocalised**.

- These delocalised electrons are **free to move around**.
- Due to these delocalised electrons, **metals are able to conduct electricity and heat**.



https://commons.wikimedia.org/wiki/File:Metallic_bonding.svg

The metallic bond that holds the particles in a metal together is the **electrostatic attraction** between the positive metal ions and the negative electrons. The electrostatic attraction between ions and electrons creates layers with atoms arranged regularly. The layers can slide over other layers which allows metals to be bent and shaped.

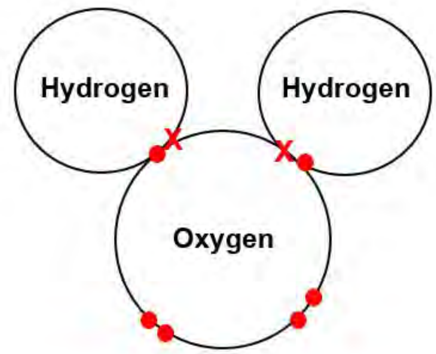
Covalent Bonding

Covalent bonds involve the **sharing** of outer shell electrons between atoms so that both atoms can benefit from the **stable full outer shell** configuration.

➔ Each covalent bond is a **pair of electrons**, providing one extra shared electron for each atom.
➔ The number of covalent bonds an atom can form is dependent on the number of electrons it has in its outer shell.

For example water (H_2O) has oxygen forming two covalent bonds each with hydrogen.

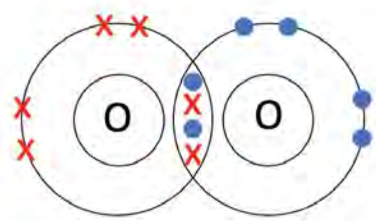
- ➔ A hydrogen atom has only one electron. To fill the first electron shell (which has a capacity of two) then it must gain 1 electron to have a full outer shell of electrons.
- ➔ An oxygen atom is in group 16 of the periodic table with 6 atoms in its outer shell and so needs to gain two more electrons for a full shell.
- ➔ Thus oxygen will form 1 covalent bond with each hydrogen atom and so each of the three atoms has a full outer shell.



Double covalent bonds can also be formed. This is where each atom shares two electrons from its outer shell.

An example of this Oxygen - a **diatomic molecule** where there are two molecules of oxygen are bonded together.

- ➔ This means that both oxygen atoms now have a full outer shell.



Covalent bonds are **strong** and require lots of energy to break them. This is because there are **strong forces of electrostatic attraction** between the pairs of negative electrons shared between the atoms and the strong positive nuclei of the atoms.

- ➔ Covalently bonded molecules can form either simple molecules or giant molecules.

Giant Covalent Structures

Diamond:

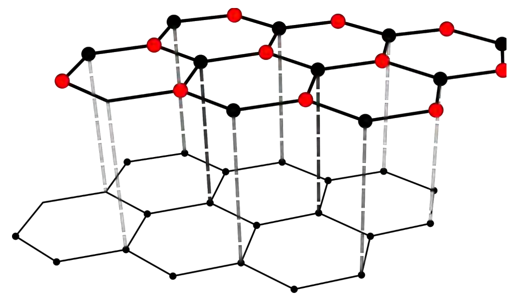
- Diamond is very **hard**. This is because each carbon atom in diamonds forms **4 covalent bonds** with other carbons atoms, making a very strong rigid giant covalent structure.

Silicon Dioxide:

- Makes up **sand**.
- Silicon and oxygen are covalent bonded into a giant structure to form SiO_2.

Graphite:

- Each carbon atom in graphite forms **only 3 covalent bonds** with other carbon atoms creating layers which are free to slide over each other.
  - This means that each carbon atom has **one delocalised electron (free electron).**
  - There are **weak intermolecular forces between the layers**.
- A carbon atom can form a maximum of 4 covalent bonds.
- The free electron means that graphite is able to **conduct heat and electricity.**
- As the layers are able to slide over each other, graphite is **soft**.



Properties of different molecules

Structure	**Type of bonding**	**Melting and Boiling points**	**Solubility in water**	**Electrical conductivity**
Ionic lattice e.g. NaCl	Ionic bonding	High	Generally soluble	Only conduct when molten or dissolved in solution
Giant covalent network e.g. SiO_2, diamond, graphite	Covalent	High	Insoluble	Do not conduct because no charged ions (except from graphite)
Simple molecular e.g. CO_2	Covalent	Low- due to weak intermolecular forces which are easily broken when the substances melt or boil	Usually insoluble	Do not conduct- no ions so no electrical charge.
Metallic lattice e.g. Na, Mg	Metallic	High	Insoluble	Conducts when solid or liquid