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8.1-8.3 Introduction to Chemical Bonding

In the unit on electron structure, it was clearly seen how the arrangement of electrons around the nucleus of the atom of an element can be deduced from the specific place it occupies on the periodic table.

Group I (alkali) metals all react in a similar fashion. This led scientists to conclude that the similarity in the electron arrangement of their atoms must be the reason for their similar behaviour.

Scientists further concluded that the way atoms join together to form a chemical bond would also be very similar, thus a very close relationship between the way electrons are arranged in atoms, and the manner in which these atoms are joined, must exist.

With this in mind, you are going to learn about how atoms combine to form new substances, and how the bonds between the atoms come about.

You are also going to learn about different types of chemical bonding, called bonding models, the forces that hold the atoms together in the chemical bond, and the substances that are formed once the bond has been made. A bonding model is a theoretical description of how a chemical bond occurs between two atoms. The following important facts are to be noted:

Types of Bonds

- The bonding model for atoms of non-metals bonding together is called the **covalent bonding model**.
    - The covalent bonding model describes how these atoms are able to join together.
    - Covalent bonding is the result of the electrons in the outermost energy level being shared between two atoms when electron shells overlap.

- The bonding model that describes what happens when atoms of a non-metal and atoms of a metal bond together is called the **ionic bonding model**.
    - It is the result of electrons in the outermost energy levels of the atoms being transferred from the metal atom to the non-metal atom to produce atoms that now have a charge. These charged atoms are known as ions.
    - Any atom wants to achieve its most stable energy state and that can only be done if its outermost energy level is completely filled with electrons.

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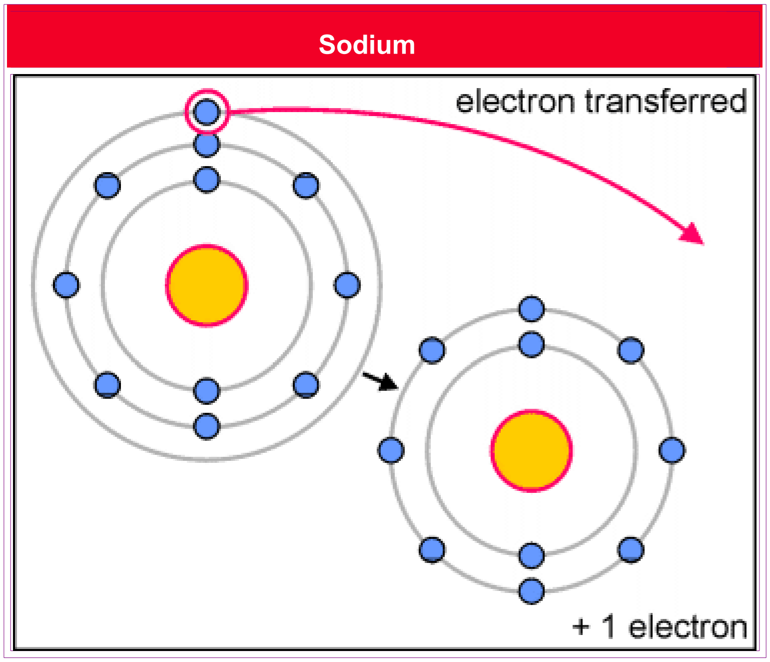
Example 1 - Consider an atom of sodium

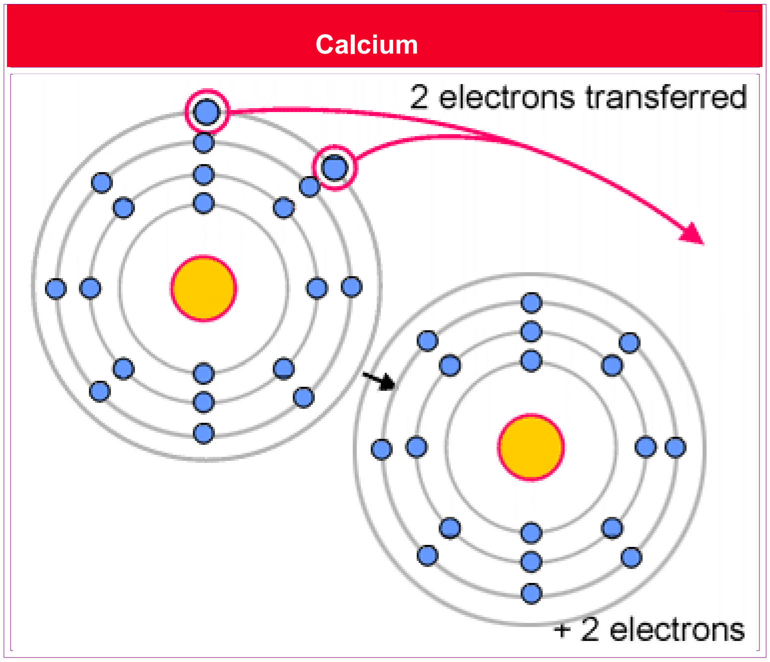
Na: Z = 11 (electron structure 2,8,1)



- Sodium has 1 electron in its outermost energy level, thus it will prefer to lose this electron to leave the completely filled energy level below.
- The **sodium cation** now has the electron structure 2, 8 and its symbol is Na^+.

Example 2 - Consider an atom of calcium

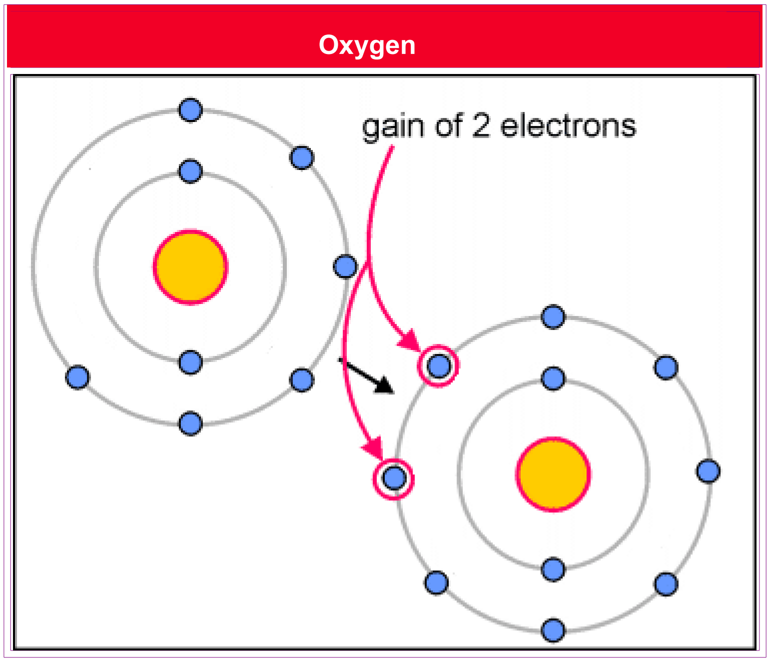
Ca: Z = 20 (electron structure 2,8,8,2)



- Calcium has 2 electrons in its outermost energy level, thus it will prefer to lose 2 electrons to leave the filled energy level below.

- The calcium cation now has the electron structure 2, 8, 8 and its symbol is Ca^{2+}.
- The opposite is true with non-metal atoms. Non-metals require to gain their electrons in the transfer process to form a negative ion called an **anion**.

Example - Consider an atom of Oxygen.
O : Z = 8 (electron structure 2, 6)



- An oxygen atom has **six** electrons in its outermost energy level, thus it will tend to **gain** two electrons to complete the outermost energy level.
- The **oxygen anion** has electron structure 2, 8 (ie has gained 2 extra electrons) and now has the symbol O^{2-}.

Now that you have been introduced to both the covalent and ionic bonding models, let us examine these models in more detail.

COVALENT BONDING

As you learnt in an earlier unit, covalent bonding occurs whenever atoms of non-metallic elements join to form a chemical bond.

The main characteristic is that this chemical bond is formed by the sharing of electrons between the atoms through the overlap of the outermost energy level.

Remember: Atoms have stable arrangements if their outer energy levels are filled.

Let us now look at examples of how this bonding is achieved.

Hydrogen atoms undergo covalent bonding. Each hydrogen atom has 1 electron in its outermost energy level. Thus to fill that energy level, the atoms approach each other until the outermost levels are able to overlap.

Once this has taken place, the electrons are **shared** between the atoms so that both outermost energy levels are now complete. The **shared electron** pair can now move around both hydrogen atoms.

The shared electron pair forms a **covalent bond**, and the new substance formed is called the **hydrogen molecule** with a symbol of H_2.

The symbol H_2 means that two hydrogen atoms are bonded together and the subscript "2" indicates the number of hydrogen atoms present in the molecule.

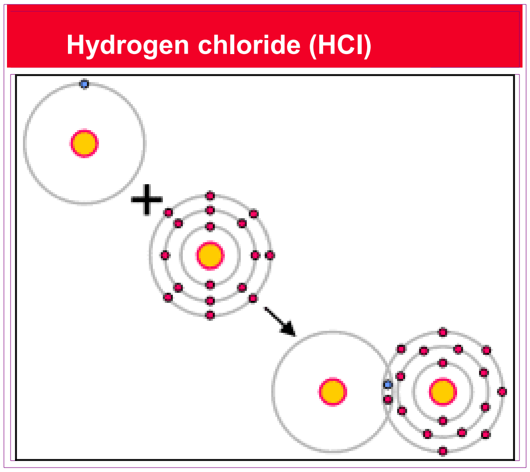
Let us now look at some other examples of covalent bonding. There is covalent bonding in a molecule of chlorine.

Each chlorine atom has 7 electrons in the outermost energy level, thus each chlorine atom requires 1 electron to complete that energy level.

Overlap occurs and electron sharing takes place. The shared electron pair can thus move around both chlorine atoms and a covalent bond is formed.

Other examples:

1. **Hydrogen chloride (HCl)**



In this **molecule**, two electrons are shared between the hydrogen and chlorine atoms to form a covalent bond.

The hydrogen atom shares its electron with the chlorine atom. Hydrogen now has a full energy level of 2 electrons.

The Cl atom shares one of its electrons with the H atom. Cl now has a full outermost energy level.

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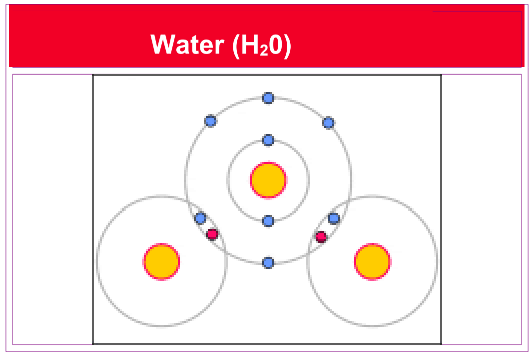
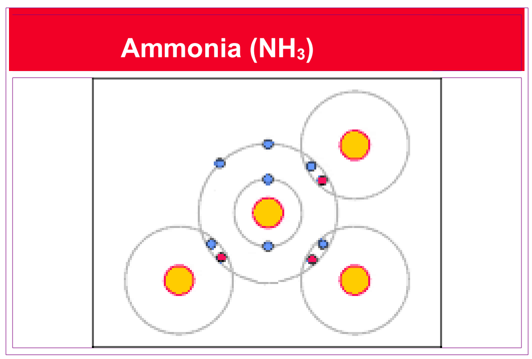
2. **Water (H_2O)**

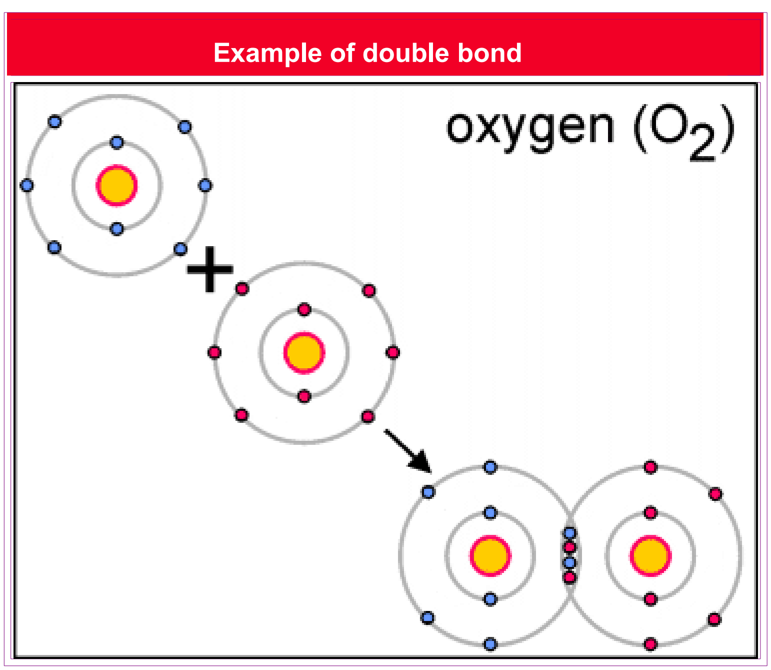


Ammonia (NH_3)



DOUBLE and TRIPLE COVALENT BONDS

If two pairs of electrons are shared, the bond is said to be a "double bond".



Note that 2 shared pairs of electrons now exist and a **double bond** is said to be present.

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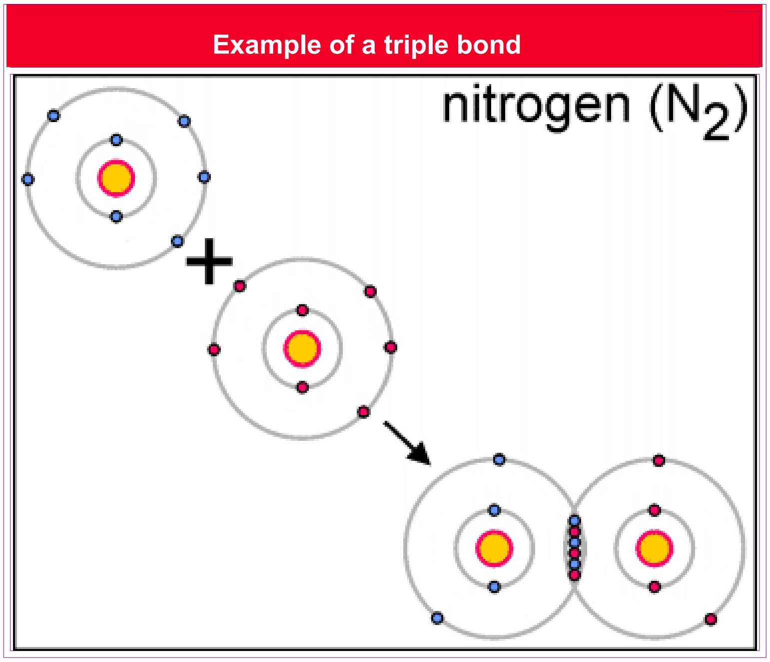
If 3 shared pair of electrons are present, then a **triple bond** is said to exist.



As can be seen in this diagram of a nitrogen molecule, 3 shared electron pairs are present, thus a triple bond exists.

PROPERTIES of COVALENT COMPOUNDS

The covalent bond formed between non-metal atoms through electron sharing is considered to be a strong bond. However, the forces holding the molecules together are **weak**.

This means that covalently bonded substances are often **gases** or **liquids** at room temperature. There are some **solids** but these solids are characterised by relatively low **melting** and **boiling** points.

They do not conduct electricity nor do they dissolve in water.

Let us examine some solids that have covalently bonded atoms making up the molecules of the solid.

There are **two** main structures that these solids can be classified into:

1. Molecular solids
2. Giant macromolecular solids

1. Molecular solids

These solids are made up of molecules in a crystalline structure

A crystalline structure is a structure which consists of a regular and orderly arrangement of particles which make up the crystal.

The molecules pack together as closely as possible and are held together by weak intermolecular forces. Intermolecular

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forces are forces that exist between molecules.

Examples of molecular solids are iodine and ice (solid H_2O).

Molecular solids are characterised by low melting and boiling points.

2. Giant macromolecular solids

These solids are made up of a network of atoms which are covalently bonded and held together by strong covalent bonds.

These giant molecules will:

- have very high melting points and boiling points
- not dissolve in water
- not conduct electricity (except for graphite)

Examples of giant macromolecular solids are:

1. Diamond
2. Graphite

PROPERTIES OF IONIC COMPOUNDS

You have already learned that ionic bonding is characterised by the following:

- Ionic bonds are made between metal and non-metal atoms.
- Electrons are transferred from the metal atom to the non-metal atom.
- Metal atoms, which lose electrons, form positive ions called cations.
- Non-metal atoms gain electrons and form negative ions called anions.

Examples

Let us consider the bonding in the substance sodium chloride.

$Na^+ + Cl^- \rightarrow NaCl$

In this reaction the sodium (Na) atom loses 1 electron while the chlorine atom accepts the electron in the transfer process. The sodium atom become an Na^+ ion and the chlorine atom becomes a Cl^- ion. The ions come together and we say that sodium chloride is an ionic compound.

Thus we can say that sodium chloride is an ionic compound.

The formula is thus NaCl.

Let us have a look at another example:

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Calcium oxide

A calcium atom transfers two electrons and becomes a Ca^{2+} ion, while an oxygen atom accepts two electrons to become an O^{2-} ion. The two ions come together to form the compound CaO.

The force of attraction between two opposite charges is called an electrostatic force of attraction.

The electrostatic attractive force is the chemical bond in the compounds sodium chloride and calcium oxide. It is called an ionic bond.

Here are some of the metal and non-metal atoms that will form ions in an ionic bond.

Metal atoms	Group	Electrons lost	Cation formed
Lithium	I	1	Li^+
Sodium	I	1	Na^+
Potassium	I	1	K^+
Magnesium	II	2	Mg^{2+}
Calcium	II	2	Ca^{2+}
Aluminium	III	3	Al^{3+}

Non-metal atoms	Group	Electrons gained	Anion formed
Oxygen	VI	2	O^{2-}
Sulphur	VI	2	S^{2-}
Chlorine	VII	1	Cl^-
Bromine	VII	1	Br^-
Iodine	VII	1	I^-

- Group I elements form cations with one positive charge.
- Group II elements form cations with two positive charges.
- Group III elements form cations with three positive charges.
- Group VII elements from anions with one negative charge.
- Group VI elements form anions with two negative charges.
- group V elements form anions with three negative charges.

IONIC SOLIDS

The solids that ionically-bonded substances form are called **ionic solids.** They are composed of ions that result in a **crystal lattice.**

A crystal lattice is described as an ordered 3 dimensional arrangement of ions held together by electrostatic forces of attraction.

A crystal of sodium chloride is neutral because the number of sodium ions (Na^+) is **equal** to the number of chloride ions (Cl^-). The forces of attraction between the ions hold them in position in the structure. They cannot move out of position.

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These crystal structures are known as **giant ionic structures.**

A compound has no overall charge because the sum of positive charges is equal to the sum of negative charges in the compound.

We use this fact to enable us to write the chemical formulae for ionic compounds.

IONIC BONDING CHEMICAL NOMENCLATURE (NAMING)

Let us now look at a number of examples of simple ionic compounds, and how their chemical formulae are determined.

a) Calcium fluoride

[Ca^{2+}] and [F^{-}]

Calcium is in Group II ∴ will form a +2 ion. Fluorine is in Group VII ∴ will form a -1 ion.

thus:

Ca^{2+}	F^{-}
___	F-
+ 2	- 2

You can see that two fluoride ions are needed to give a total charge of "-2".
This balances the "+2" charge of the Ca^{2+}.

The charges are now **balanced.**

∴ The formula is CaF_2.

2 x F^{-} are needed to balance out Ca^{2+} ∴ the subscript of "2" is used to indicate this.

b) **Sodium oxide**

Na^{+} O^{2-}

∴ We will need an extra Na^{+} to balance the O^{2-}.

Na^{+}	O^{2-}
Na^{+}	__
+2	-2

∴ the formula is Na_2O

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c) **Aluminium oxide**

Al^{3+} O^{2-}

When you have **odd** and **even** charges, the easiest way is to find the lowest common multiple of the charges in this case it is "6 ".

∴		
∴	Al^{3+}	O^{2-}
	Al^{3+}	O^{2-}
	__	$\underline{O^{2-}}$
	$\underline{+6}$	$\underline{-6}$

Thus the formula is Al_2O_3.

Now, you must be introduced to some more complicated ions that you will have to learn. You must be able to use these to write chemical formulae