Covalent Bonding part 1

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Covalent bonding occurs between two **non-metal** atoms e.g. H_2O, NH_3, O_2, N_2, H_2, alkanes, alkenes etc.

sharing of electrons between **two non-metal atoms** resulting in an electrostatic attraction between the positive nucleus and the electrons.

Each of the two atoms contributes one electron to a covalent bond.

Dot and cross diagrams

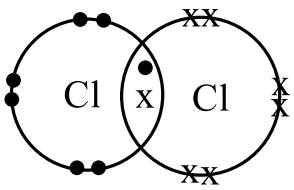
To represent covalent bonding we can use a dot and cross diagram that shows the electrons on each atom and the bond.

Go to the periodic table to find which **group** each element is in to get the **number of outer electrons**. The outer electrons are important, as they are involved in the bonding.

✓ Be careful as they may ask you to draw ALL the electrons including the inner shells.

Chlorine, Cl_2

Using chlorine as an example. Chlorine is in group 7 and has 7 outer electrons:



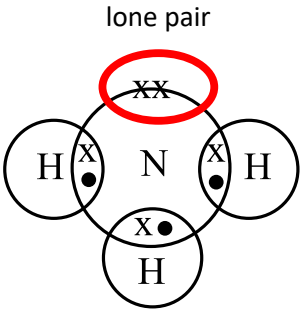
Make sure you draw the two "circles" overlapping. The two electrons in the "overlap region" is the covalent bond.

The circles represent the **outer shells**. It is actually different orbitals that overlap forming the bond but for the purposes of dot and cross, just draw it as above.

The aim is to get 8 electrons around **each** atom i.e. a stable arrangement. This means you count the two electrons in the bond for each atom i.e. the two electrons in the bond belong to both atoms.

✓ There are examples of molecules where there are **not** 8 electrons around each atom e.g. BF_3, which only has 6 electrons around the boron. It's quite rare and not worth worrying about, but it does exist.

Ammonia, NH_3



Nitrogen has **5 outer electrons** and hydrogen just has the **1 electron**.

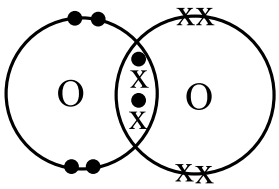
There are three N-H bonds, each consisting of one electron from the nitrogen and one electron from the hydrogen to complete the stable full shell arrangement.

✓ There are 8 electrons around the nitrogen and 2 electrons per hydrogen atom. It is ok to have 2 around hydrogen as this is still a full shell.

There are two electrons left over on the nitrogen, which are not involved in any bonding. These are called **lone pairs** or non-bonding pairs.

Oxygen, O_2

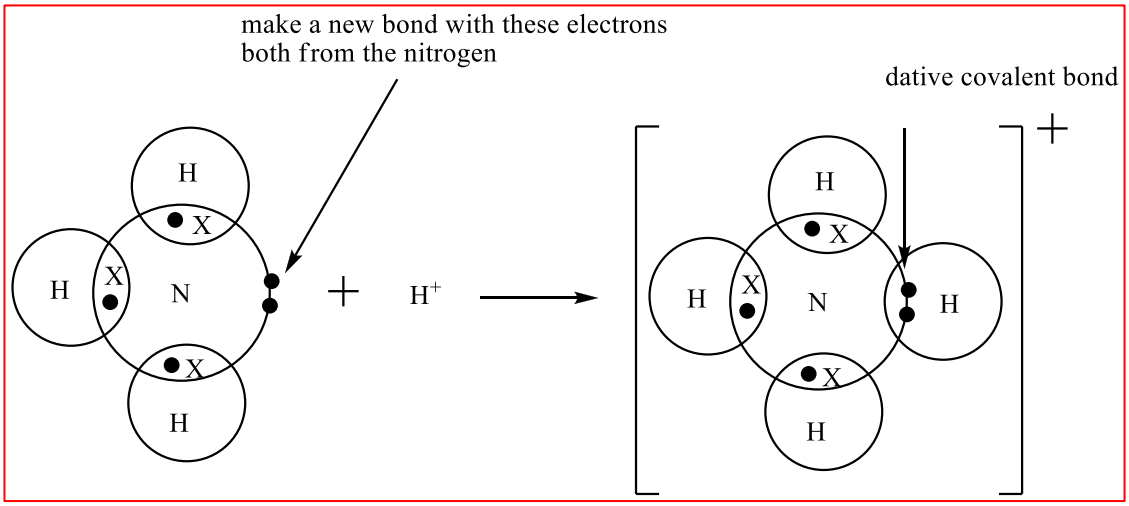
Oxygen has a **double bond** and therefore **4 bonding electrons** as shown below:



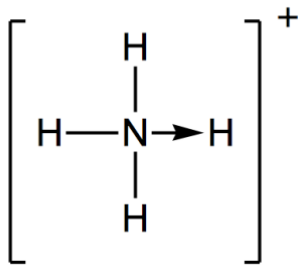
Dative Covalent

Another type of covalent bond is where the two electrons come from the **same** atom rather than one from each of two different atoms. This is known as a dative covalent bond.

The NH_4^+ (ammonium) ion is a classic example of this:



The two electrons are still shared between two atoms and it is still a covalent bond as before.



Another way to represent a dative covalent bond is by drawing an arrow from the nitrogen to the hydrogen where the covalent bond is made.

Bond length/strength

Different covalent bonds are different lengths i.e. a C-H bond length is different from a C-C bond etc. This is due to different degrees of orbital overlap. The more they overlap then the stronger the bond.

bond length **decreases** as the bond gets stronger

Multiple bonds are shorter than single bonds. For example, C=C is shorter than C-C and therefore stronger. But it is not twice as strong as some people assume.

There are different types of covalent structures that we will look at in part 2 of this tutorial.

Click here for more covalent bonding