

THE NATURE OF BONDS

Name Form



Types of bond

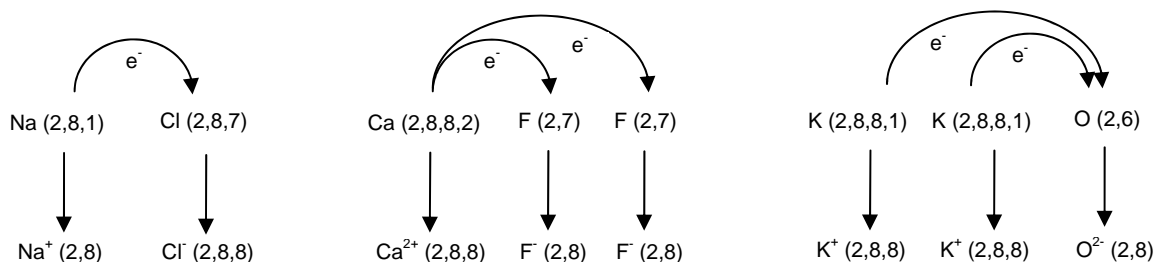
There are three types of bond – ionic, covalent and metallic.

Type of bonding	ionic	Covalent	Metallic
Nature of bonding	Electrostatic attraction between positive and negative ions.	Shared pair of electrons between atoms.	Attraction between lattice of positive metal ions and delocalised outer shell electrons.
Types of structure which have this type of bonding			
Strength of bonds	The smaller the ions and the greater the charge on the ions, the stronger the attraction between the ions (usually). This is due to a greater charge density within the structure.	The shorter the bond, the stronger the bond (usually). Double bonds are stronger than single bonds, while triple bonds are stronger than double bonds.	The smaller the metal ions, the greater the charge on the ions, and the more delocalised outer shell electrons there are, the stronger the attraction between the ions and electrons (usually). This is due to a greater charge density within the structure.

There are also three types of forces between molecules (van der Waals' forces, dipole-dipole attractions and Hydrogen bonds). These are **NOT** bonds because they are too weak.

Formation of ions

Ions can be formed when a **metal** reacts with a **non-metal**. The metal atoms lose electrons to form positive ions while non-metal atoms gain electrons to form negative ions (both obtaining full outer shells). For example (in *over-simplified* GCSE terms)



makes sodium chloride (NaCl)

makes magnesium fluoride (MgF_2)

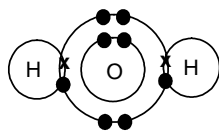
makes potassium oxide (K_2O)

Remember that ionic bonds are the attraction between positive and negative ions.

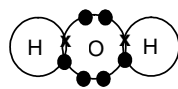
Covalent bonds

Covalent bonds can be formed when a **non-metal** reacts with a **non-metal**. The atoms share electrons to obtain full outer shells. Two shared electrons make a **single bond**, four shared electrons make a **double bond**, and six shared electrons makes a **triple bond**.

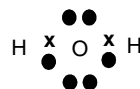
For example in water:



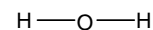
all the electrons



outer shell electrons only



outer shell electrons only
("dot-cross" diagram)



covalent bonds only
("stick" diagram)

Atoms	Group 4 atoms	Group 5 atoms	Group 6 atoms	Group 7 atoms	H
Number of covalent bonds					

How to draw "dot-cross" diagrams

- 1) Draw a stick diagram (use the help above to work out how to draw the stick diagram).
- 2) Re-draw the molecule without the sticks.
- 3) Draw a dot and a cross instead of each stick (i.e. ●x for a single bond, ●x●x for a double bond, etc.). Dots and crosses represent electrons from different atoms.
- 4) Work out how many electrons there are in the outer shell of each atom (e.g. atoms in Group 7 have 7 electrons in their outer shell), and then add in any of these electrons that are not already drawn on in the covalent bonds.

Note that elements in period 3 and beyond can fit more than 8 electrons in their outer shell (due to availability of d orbitals in the shell). For example, P and S atoms can have more than electrons in the third shell (as extra electrons can occupy the 3d orbitals).

For ions, any electric charge means that the ion has more or less electrons than the neutral atoms would have. For example, a 1+ ion has one less electron than the atoms, and a 2- ion has two extra electrons.

In your book, draw stick and then dot-cross diagrams for each of the following molecules and ions.

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|--------------------|---------------------|--------------------|------------------------------------|-----------------------------------|-----------------------------------|-------------------------------------|
| 1) F ₂ | 4) H ₂ S | 7) PF ₃ | 10) CO ₂ | 13) SO ₄ ²⁻ | 16) C ₂ H ₂ | 19) BeCl ₄ ²⁻ |
| 2) NH ₃ | 5) CH ₄ | 8) PF ₅ | 11) NH ₄ ⁺ | 14) BCl ₃ | 17) NO ₃ ⁻ | 20) IF ₄ ⁺ |
| 3) O ₂ | 6) H ₂ | 9) SO ₂ | 12) AlCl ₄ ⁻ | 15) SF ₆ | 18) BeCl ₂ | |