Bonding and Structure *(Y10)*

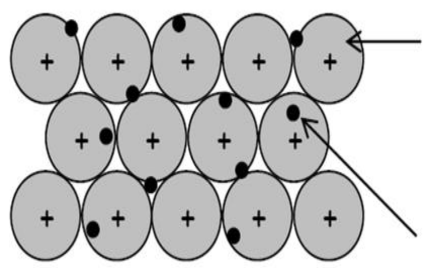
There are 3 types of chemical bonding:

Ionic bonding	**Covalent bonding**	**Metallic bonding**
- formed between *metal* and *non-metal* atoms.	- formed between *non-metal* atoms.	- formed between *metal* atoms.
- involves electrons being *transferred* to gain full outer shells.	- involves electrons being *shared* to gain full outer shells.	- involves electrons becoming **delocalised**.
- found in *ionic compounds* eg. sodium chloride ($NaCl$) and aluminium oxide (Al_2O_3).	- found in *covalent compounds* eg. carbon dioxide (CO_2) and methane (CH_4) and in some *non-metal elements* eg. H_2, Cl_2, N_2, O_2.	- found in *metals*. eg. potassium (K), copper (Cu), iron (Fe).
- results in **giant ionic structures** eg. sodium chloride ($NaCl$) and aluminium oxide (Al_2O_3).	- results in **simple molecular** or **giant covalent structures**.	- results in **giant metallic structures**.

Metallic crystals

Metals, eg. copper, consist of a regular arrangement of *positive ions surrounded by a sea of delocalised*

delocalised electrons*. Lots of energy is needed to overcome these strong electrostatic forces* so The outer shell electrons from each atom become **delocalised**. The structure is held together by the **strong electrostatic forces** between the positive ions and the sea of metals have *high*



The *layers of ions slide over each other* when a force is applied but the structure does not break apart because it is held together by the delocalised electrons. This makes metals *malleable* – they can be hammered The *delocalised electrons* are free to move throughout the structure, making metals good *conductors of electricity.*

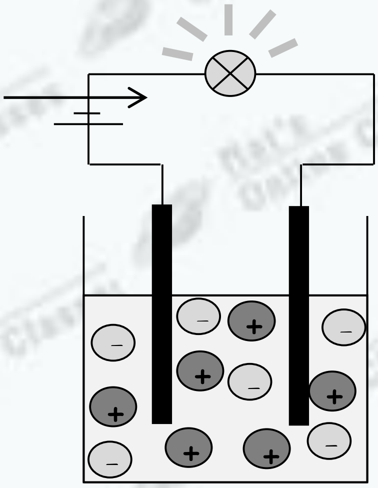
electrons.

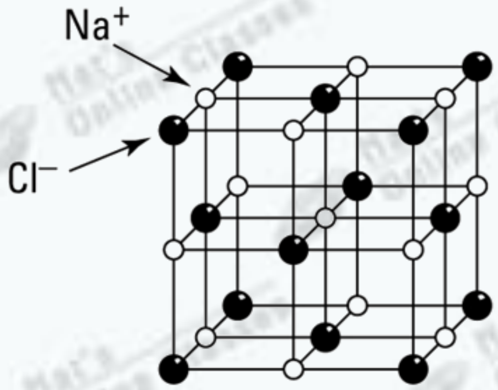
Ionic compounds

Ionic compounds, eg. NaCl, consist of a regular arrangement of *positive and negative ions*. Electrons from the metal atoms are transferred to the non-metal atoms so that all the atoms have full outer shells. The metal atoms *lose their outer shell electrons* to become *positively charged* (they are **oxidised**), and the non-metal atoms *gain electrons* to become *negatively charged* (they are **reduced**).

The oppositely charged ions are attracted together, forming a regular arrangement called a **lattice** or **giant ionic structure**.

Ionic compounds *conduct electricity* when they are *molten or in solution* because the ions are free to move. *Solid* ionic compounds *do not conduct* because the ions are *fixed in the lattice*.





Strong electrostatic forces between oppositely charged ions give ionic compounds *high melting points*. The *higher the charge* on the ions, the *higher the melting point* and *boiling point*.

Key Skill

Drawing dot and cross diagrams to represent ionic compounds

1. Draw the electron arrangement of the metal atom, representing the *electrons as dots*, and *missing off the outer shell electrons* which have been lost in ionic bonding.

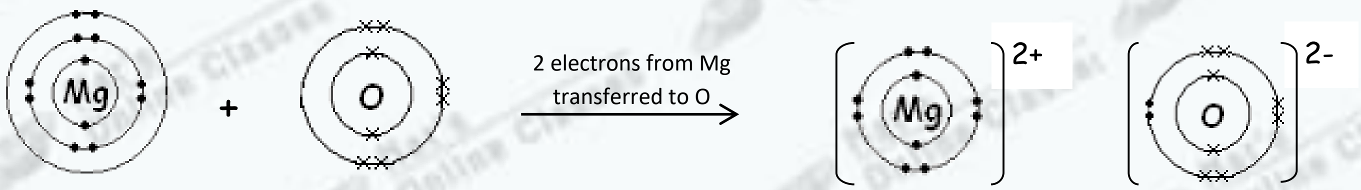
2. Draw the electron arrangement of the non-metal atom separately and to the right of the ion you have drawn, representing the *electrons as crosses*. Fill in any *spaces in the outer shell of the atom with dots* (electrons from the metal atom that have been gained in bonding).

3. Draw a set of *square brackets* around each of the ions.

4. Work out the charges on the ions. The *metal is always positive* (having lost negative electrons). The *size of the positive charge is the group no.* (the no. of outer shell electrons that have been lost). *The non-metal is always negative*, and the size of the charge is the no. of dots (electrons it has gained from the metal) in its outer shell. Write the *charges as superscripts* to the right of each set of square brackets.

Examples

i. Magnesium oxide, MgO



2, 8, 2 | 2, 6 | [2, 8]$^{2+}$ | [2, 8]$^{2-}$

Mg has 2 electrons in its outer shell so needs to lose 2 for a *full outer shell.*

O has 6 electrons in its outer shell so needs to gain 2 for a *full outer shell.*

Mg has *lost* 2 negative charges (electrons) so becomes an *ion* with a charge of +2.

O has *gained* the 2 negative charges (electrons) from Mg so becomes an *ion* with a charge of -2. Only 1O^{2-} ion is needed to balance the 2+ charge on the magnesium ion

ii. Aluminium fluoride, AlF_3

2, 8, 3

Al has 3 electrons in its outer shell so needs to lose 3 for a full outer shell.

2, 8, 7

F has 7 electrons in its outer shell so needs to gain 1 for a full outer shell. 3 F atoms are needed to accept the 3 electrons from the outer shell of Al.

$[2, 8]^{3+}$

Al has *lost* 3 negative charges (electrons) so becomes an *ion* with a charge of +3.

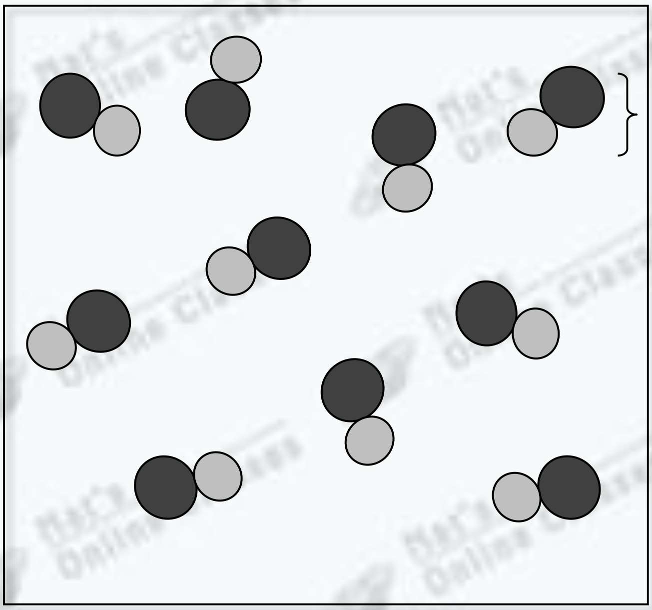
$[2, 8]^{-}$

The 3 F atoms have each *gained* 1 negative charge (electron) from Al so become *ions* with a charge of -1. 3 F^- ions are needed to balance the 3+ charge on Al and make the compound neutral.

Exam tip: *Make sure that you can draw dot and cross diagrams to represent the ionic compounds formed from any combinations of group 1,2,3 and 5,6,7 ions. eg. NaCl,* $MgCl_2$, $AlCl_3$, Na_2O, Na_3N.

Covalent substances

The vast majority of covalent substances have **simple molecular structures** – they exist as molecules, or groups of atoms chemically joined. The atoms in the molecules are held together by the strong attraction between the positive nuclei of the atoms and the negatively charged shared electrons. These are **covalent bonds**.



Strong covalent bonds hold the *atoms in the molecules*

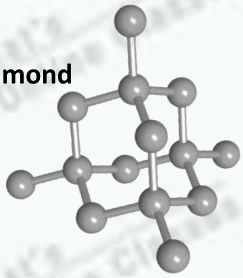
There are only *weak* **intermolecular forces** between the molecules. This means that the molecules are easily pulled away from each other so the substances have *low melting and boiling points*.

The molecules are *neutral* so cannot carry electrical charge – covalent substances made up of simple molecules *do not conduct electricity*.

A very small number of covalent substances exist as **giant covalent structures.** Diamond and graphite, different forms (or allotropes) of carbon, are particularly important examples.

Giant covalent structures have *strong covalent bonds* throughout the structure. *Lots of energy is needed* to break the many covalent bonds between atoms so they have *high melting and boiling points*.

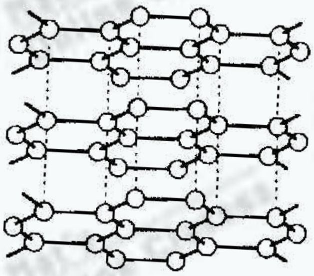
Diamond



- In *diamond*, each C atom is attached to *4* others. The structure is *hard* because there are *strong covalent bonds* throughout the structure. Diamond is therefore used for cutting.

- In *graphite*, the C atoms are arranged in *layers*. There are *strong covalent bonds within the layers* but only *weak intermolecular forces between layers*. This means they can *slide* over each other. Graphite can therefore be used as a lubricant.

- In *graphite*, each C atom is attached to *3* others, leaving 1 spare electron in the outer shell of each C atom. This electron is free to move – it is *delocalised* enabling graphite to *conduct electricity*.



Graphite

Exam tip: Make sure that you can describe and explain the differences in the properties of each type of structure. Be particularly careful with the key scientific terms you use – it is incorrect to talk about bonds being broken when simple molecules are melted or boiled, and there are only intermolecular forces in simple molecular substances and not in ionic or metallic structures.

Drawing dot and cross diagrams to represent covalent substances

1. Work out the formula of the covalent substance.

2. Use the *group number* in the periodic table to work out how many electrons each of the atoms has in its outer shell and therefore how many electrons each atom needs to gain for a full outer shell.

3. Work out how many electron pairs each atom needs to share. If an atom needs *1 more electron* for a full outer shell it needs to share *1 pair*; if it needs 2 electrons it needs to share 2 pairs etc.

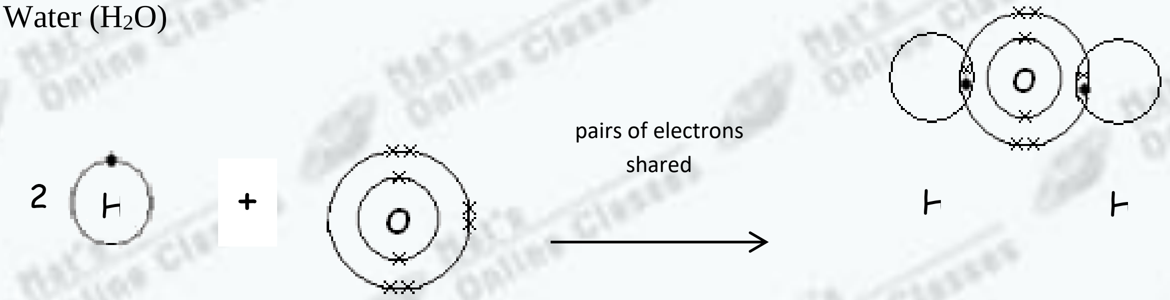
4. Use this to work out how the atoms are arranged. Draw the electronic structures of all the atoms, *interlocking the outer shells* of the atoms that are bonded together but leaving them empty of electrons.

5. Draw in the *shared pairs* of electrons where the outer shells are interlocked, ensuring that there is *one dot and one cross* in every pair.

6. Fill in the *rest of the outer shell* electrons for each atom outside the interlocking area, ensuring that the total number of crosses or dots in the outer shell is the same as the group no.

Examples

i. Water (H_2O)

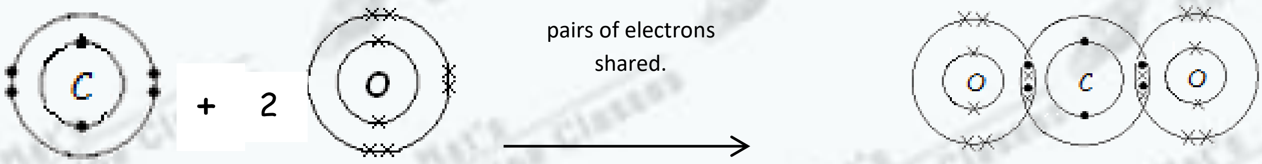


H has 1 electron in its outer shell so needs to gain 1 for a *full outer shell,* It needs to share *1 pair of electrons*.

O has 6 electrons in its outer shell so needs to gain 2 for a *full outer shell.* It needs to share *2 pairs of electrons*

The O atom and both H atoms have gained *full outer shells* by sharing electron pairs. Each electron pair is a single bond, so the structure can also be drawn as H – O – H. The H_2O *molecule* formed is *neutral*.

ii. Carbon dioxide (CO_2)



C has 4 electrons in its outer shell so needs to gain 4 for a *full outer shell*. It needs to share *4 pairs of electrons*.

O has 6 electrons in its outer shell so needs to gain 2 for a *full outer shell.* It needs to share *2 pairs of electrons*.

The C atom and both O atoms have gained *full outer shells* by sharing electron pairs. Each electron pair is a single bond, so the structure can also be drawn as O = C = O. The CO_2 *molecule* formed is *neutral*.

Exam tip: *Practise drawing dot and cross diagrams of the covalent substances named on the specification – hydrogen, chlorine, hydrogen chloride, water, methane, ammonia, oxygen, nitrogen, carbon dioxide, ethane and ethene. These molecules, with only the outer shells shown, are below:*

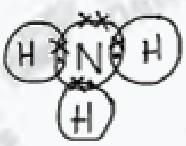
H – H

Cl – Cl

O = O

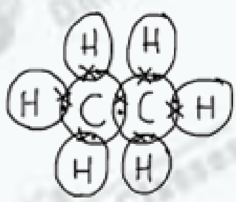
N ≡ N

H – Cl

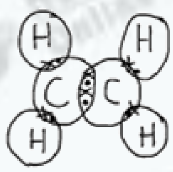


H – N – H
|
H

H
|
H – C – H
|
H



H H
| |
H – C – C – H
| |
H H



H H
\ /
C = C
/ \
H H