Cambridge OL

Chemistry

CODE: (5070)

Chapter 03

Bonding structure

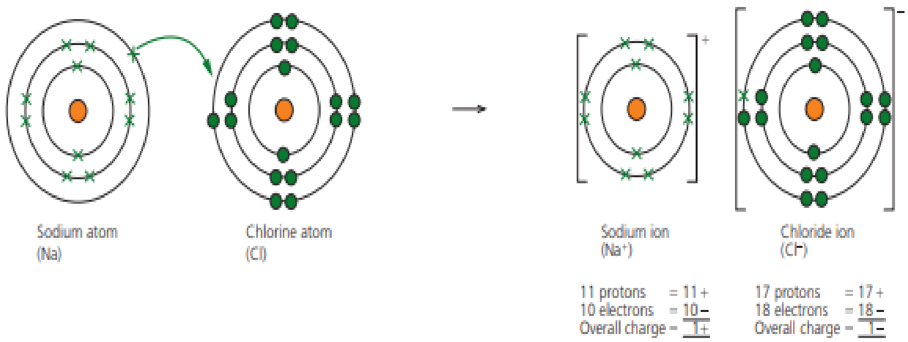
3.1 Ionic compounds

Ionic bonds form when metals combine with non-metals, transferring electrons during chemical reactions, enhancing stability due to full outer shells. Examples include sodium chloride and chlorine.

sodium + chlorine → sodium chloride

Atoms in a reaction obtain full outer shells and become 'like' the noble gas nearest to them in the Periodic Table. This is demonstrated using a dot-and-cross diagram, where electrons are crossed and dots.

▲ **Figure 3.1** The properties of salt are very different from those of the sodium and chlorine it was made from. To get salt into your diet you would not eat sodium or inhale chlorine!



▲ **Figure 3.2** Ionic bonding in sodium chloride

To lose electrons in this way is called **oxidation**. The chlorine atom has become a chloride ion with an electronic configuration like argon. To gain electrons in this way is called **reduction**. In the chemical process producing sodium chloride both oxidation and reduction have taken place and so this is known as a **redox reaction**

Key definitions

Oxidation involves loss of electrons.

Reduction involves gain of electrons.

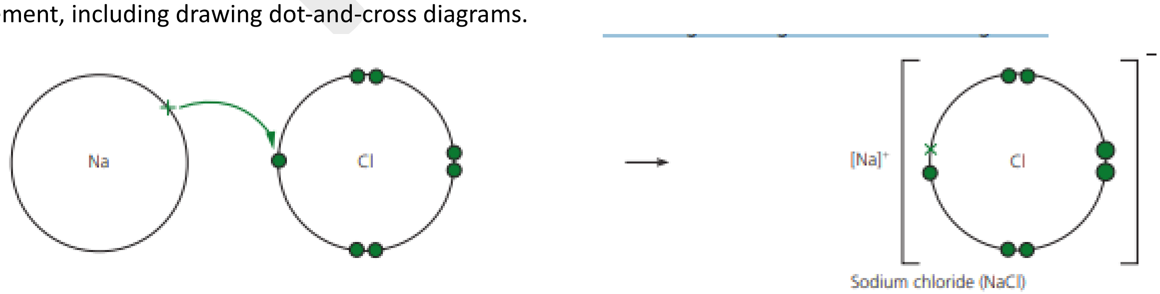
Redox reactions involve simultaneous oxidation and reduction.

Bonding involves the attraction of oppositely charged ions, such as sodium and chloride ions, through strong electrostatic forces, resulting in the formation of sodium chloride. This type of bonding is called ionic bonding

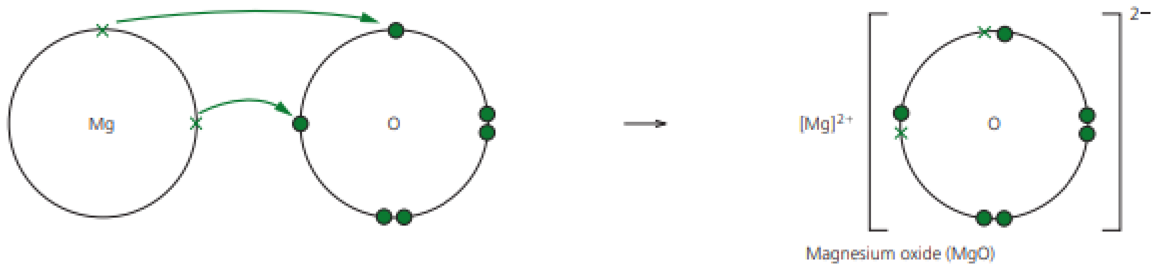
Key definition

An **ionic bond** is a strong electrostatic attraction between oppositely charged ions.

All students should be able to describe the formation of ionic bonds between elements of Group I and Group VII and draw dot-and-cross diagrams. Extended learners should be able to do this for any metallic and non-metallic element, including drawing dot-and-cross diagrams.

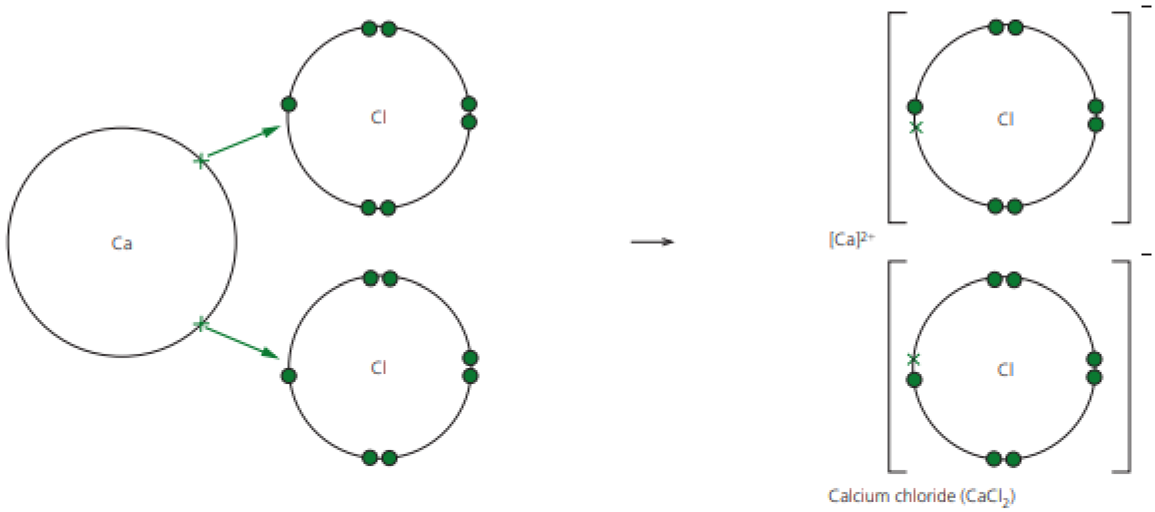


▲ **Figure 3.3** Simplified diagram of ionic bonding in sodium chloride



▲ **Figure 3.4** Simplified diagram of ionic bonding in magnesium oxide

Figure 3.4 illustrates electron transfers between magnesium and oxygen atoms in magnesium oxide formation, where oppositely charged Mg^{2+} and O^{2-} atoms attract each other.



▲ **Figure 3.5** The transfer of electrons that occurs during the formation of calcium chloride

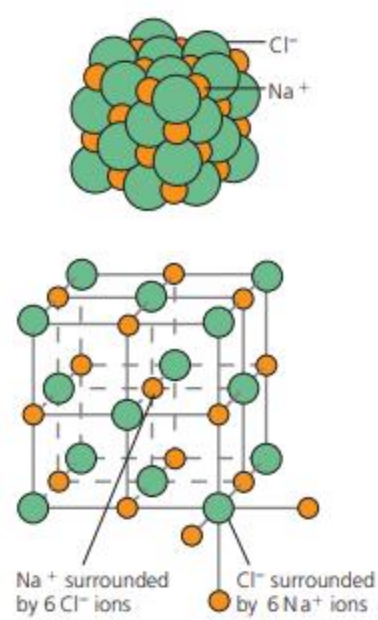
Ionic structures

Ionic structures are solids at room temperature and have high melting and boiling points. The ions are packed together in a regular arrangement called a **lattice**. Within the **lattice**, oppositely charged ions attract one another strongly.

Figure 3.6 shows only a tiny part of a small crystal of sodium chloride. Many millions of sodium ions and chloride ions would be arranged in this way in a crystal of sodium chloride to make up the **giant ionic lattice structure**.

Properties of ionic compounds Ionic compounds have the following properties.

» They are usually solids at room temperature, with high melting and boiling points. This is due to the strong electrostatic forces holding the crystal lattice together.

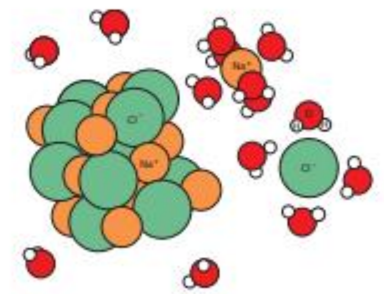


▲ **Figure 3.6** The structure of sodium chloride

» They are usually hard substances.
» They mainly dissolve in water.

» They usually conduct electricity when in the molten state or in aqueous solution. The forces of attraction between the ions are weakened and the ions are free to move to the appropriate electrode.

» They usually cannot conduct electricity when solid, because the ions are not free to move.



▲ **Figure 3.8** Salt (an ionic compound) dissolving in water

Formulae of ionic substances

Ionic compounds contain positive and negative ions with balanced charges. For example, sodium chloride has a formula NaCl, while magnesium chloride has a formula $MgCl_2$. Oxidation numbers indicate the oxidation or reduction of an ion compared to its atom. For example, Na+ has an oxidation number of +1, while Mg2+ has an oxidation number of +2. Some elements, like copper and iron, have two ions with different oxidation numbers, forming different compounds with chlorine. Table 3.1 shows the oxidation numbers of a series of ions and includes groups of atoms with net charges. For example, magnesium nitrate has a formula of $Mg(NO_3)_2$, with two nitrate ions for every magnesium ion.

$$Mg\,(N\,O_3)_2$$

$$1Mg : 2N : 6O$$

▼ **Table 3.1** Oxidation numbers (valencies) of some elements (ions) and groups of atoms

	Oxidation number (valency)					
	1		**2**		**3**	
Metals	Lithium	(Li^+)	Magnesium	(Mg^{2+})	Aluminium	(Al^{3+})
	Sodium	(Na^+)	Calcium	(Ca^{2+})	Iron(III)	(Fe^{3+})
	Potassium	(K^+)	Copper(II)	(Cu^{2+})		
	Silver	(Ag^+)	Zinc	(Zn^{2+})		
	Copper(I)	(Cu^+)	Iron(II)	(Fe^{2+})		
			Lead	(Pb^{2+})		
			Barium	(Ba^{2+})		
Non-metals	Fluoride	(F^-)	Oxide	(O^{2-})		
	Chloride	(Cl^-)	Sulfide	(S^{2-})		
	Bromide	(Br^-)				
	Hydrogen	(H^+)				
Groups of atoms	Hydroxide	(OH^-)	Carbonate	(CO_3^{2-})	Phosphate	(PO_4^{3-})
	Nitrate	(NO_3^-)	Sulfate	(SO_4^{2-})		
	Ammonium	(NH_4^+)	Dichromate(VI)	($Cr_2O_7^{2-}$)		
	Hydrogencarbonate	(HCO_3^-)				
	Manganate(VII)	(MnO_4^-)				

Oxidation number

The following points should be remembered when using oxidation numbers.

» Roman numerals (I, II, III, IV, V, VI, VII, VIII) are used in writing the oxidation number of an element.

» This number is placed after the element that it refers to.

» The oxidation number of the free element is always 0,

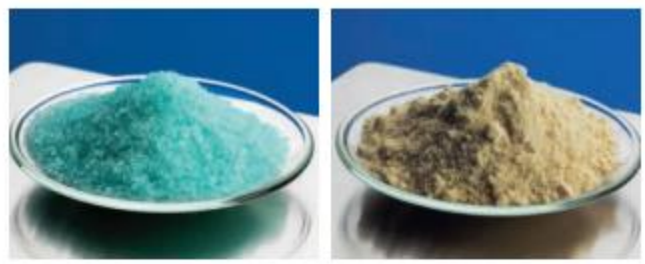
» In simple monatomic ions, the oxidation number is the same as the charge on the ion. So iodine has an oxidation number of 0 in I2 but an oxidation number of −1 in I−.

» Compounds have no charge overall. Hence the oxidation numbers of all the individual elements.

» An increase in the oxidation number, for example from +2 to +3 as in the case of Fe^{2+} to Fe^{3+}, is oxidation

» However, a reduction in the oxidation number

During a redox reaction, the substance that brings about **oxidation** is called an **oxidising agent** and is itself reduced during the process. A substance that brings about **reduction** is a **reducing agent** and is itself oxidised during the process.



▲ **Figure 3.9** Iron(II) sulfate is pale green, while iron(III) sulfate is yellow

Key definitions

An **oxidising agent** is a substance that oxidises another substance and is itself reduced.

A **reducing agent** is a substance that reduces another substance and is itself oxidised.

Key definitions

Oxidation involves an increase in oxidation number.

Reduction involves a decrease in oxidation number.

The reaction involves the oxidation of manganate(VII) ions to iron(III) ions, and the reduction of manganate(VII) ions to pale pink manganese(II) ions. Potassium iodide, a common reducing agent, produces a color change from yellow-brown to blue-black when added to a solution of iron(III) solution, indicating the reduction of iron(III) ions.

▲ **Figure 3.10** Manganate(VII) ions (oxidising agent) and iron(II) ions (reducing agent) are involved in a redox reaction when mixed

▲ **Figure 3.11** The blue-black colour shows the presence of iron(III) ions

Redox processes can be identified by examining the oxidation numbers on the two sides of a chemical equation. For instance, magnesium dissolves in sulfuric acid, producing hydrogen gas with an oxidation number of 0 and +1 respectively.

magnesium + sulfuric acid → magnesium sulfate + hydrogen

$$\underset{0}{Mg(s)} + \underset{+1}{H_2SO_4(aq)} \rightarrow \underset{+2}{MgSO_4(aq)} + \underset{0}{H_2(g)}$$

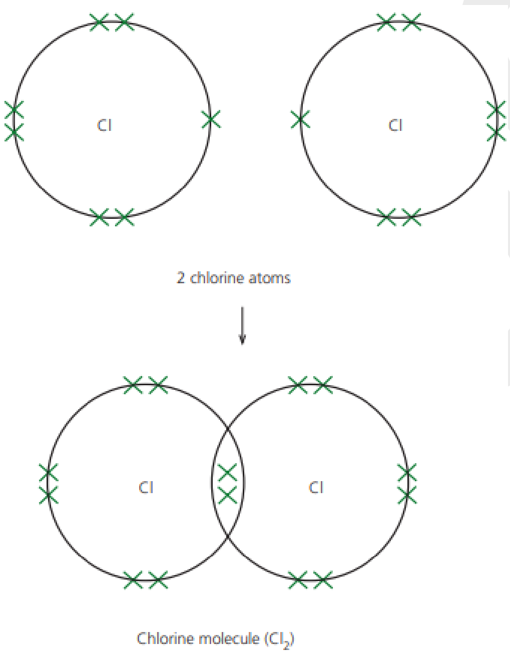
The **reducing agent**, magnesium, undergoes oxidation by sulfuric acid, resulting in an increase in its oxidation number, while the reducing agent, magnesium, reduces the oxidation number of hydrogen.

3.2 Covalent bond

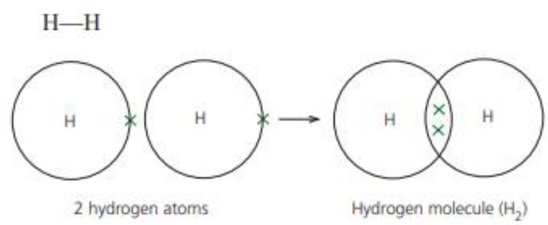
A **covalent bond** is a stable atom-to-atom interaction where electrons are shared in their outer shells, forming a noble gas electronic configuration, as seen in hydrogen molecule H_2.

> **Key definition**
> A **covalent bond** is formed when a pair of electrons is shared between two atoms leading to noble gas electronic configurations.

A single covalent bond is formed between two hydrogen atoms, resulting in a molecule of hydrogen. This bond, represented by a single line, ensures that the two hydrogen atoms share a pair of electrons, forming an electronic configuration similar to helium.



▲ **Figure 3.13** The electron sharing to form the single covalent bond in Cl_2 molecules (showing outer electron shells only)



a The electron sharing to form the single covalent bond in H_2 molecules

b Model of a hydrogen molecule

▲ **Figure 3.12**

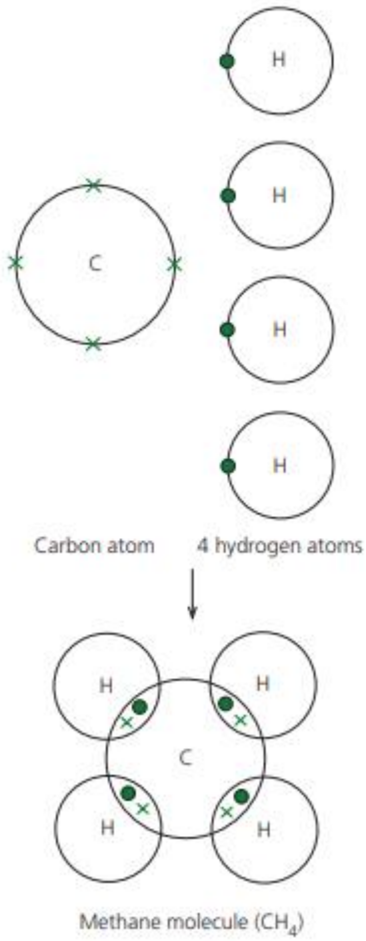
Hydrogen chloride and helium share electrons to obtain their nearest noble gas electronic configurations. Hydrogen shares one electron to gain helium's electronic configuration, while chlorine shares one to gain argon's electronic configuration.

Other covalent compounds

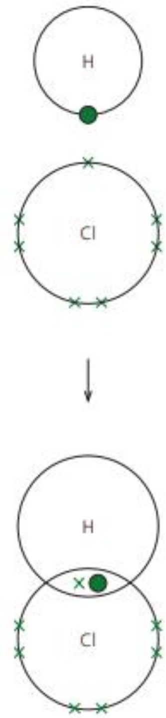
Methane (natural gas) is a gas whose molecules contain atoms of carbon and hydrogen. The electronic configurations are:

C 2,4 H 1

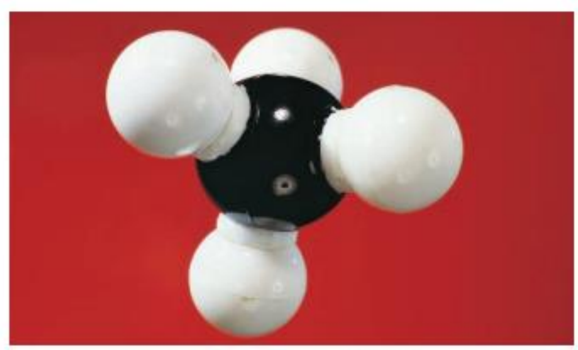
Carbon atoms need four more electrons for noble gas neon, and one for helium. Sharing electrons forms these configurations, as shown in Figure 3.15 and Figure 3.16.



▲ **Figure 3.15** Dot-and-cross diagram to show the formation of a methane molecule



▲ **Figure 3.14** Dot-and-cross diagram to show the formation of a hydrogen chloride molecule



▲ **Figure 3.16** Model of a methane molecule

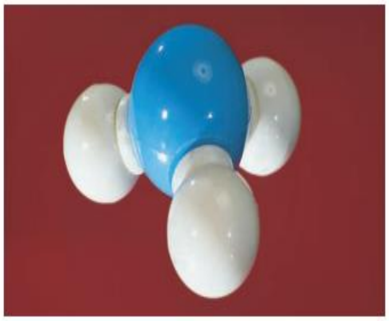
Ammonia is a gas containing the elements nitrogen and hydrogen. It is used in large amounts to make fertilisers. The electronic configurations of the two elements are;

N 2,5 H 1

Water is a liquid containing the elements hydrogen and oxygen. The electronic configurations of the two elements are:

O 2,6 H 1

The oxygen and hydrogen atoms share electrons, forming a water molecule with two single covalent bonds, as shown in Figure 3.19 and Figure 3.20.



▲ **Figure 3.18** Model of the ammonia molecule

▲ **Figure 3.20** Model of a water molecule – this is a V-shaped molecule

Carbon dioxide is a gas containing the elements carbon and oxygen. The electronic configurations of the two elements are:

C 2,4 O 2,6

Carbon and oxygen atoms share four and two electrons to obtain neon's electronic configuration, achieved through double covalent bonds, respectively, while carbon dioxide is a linear molecule.

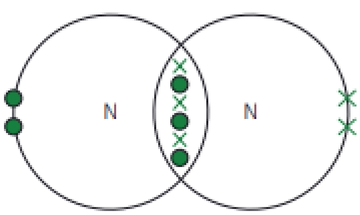
As with the carbon dioxide molecule the double bond can be shown by a double line between the two oxygen atoms, representing two shared pairs of electrons.

O=O

Each nitrogen atom has the electronic configuration shown below:

N 2,5

The triple bond can be shown in the molecule as N≡N.



▲ **Figure 3.24** Dot-and-cross diagram to show the bonding in a nitrogen molecule

Methanol, CH_3OH, has three distinct atom configurations, requiring a dot-and-cross diagram to ensure that atoms bonding do not have both dots or crosses.

C 2,4 H 1 O 2,6

The carbon atom shares four electrons with three hydrogens, each sharing one more. The oxygen atom shares one with the carbon atom, resulting in the electronic configuration of helium and neon.

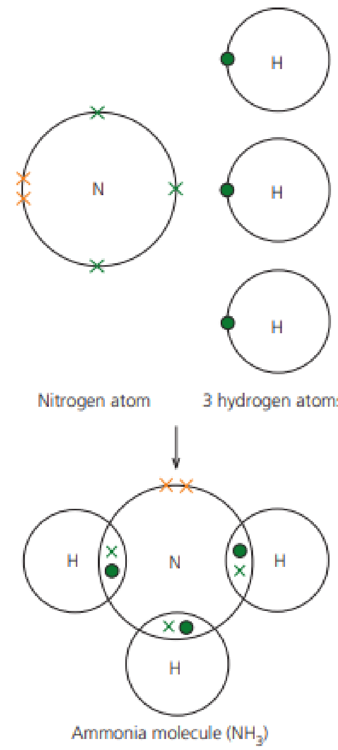


Figure 3.17 Dot-and-cross diagram to show the formation of an ammonia molecule

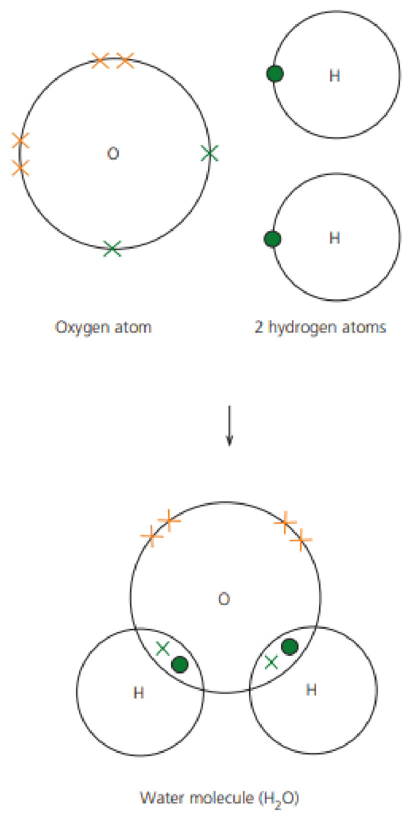
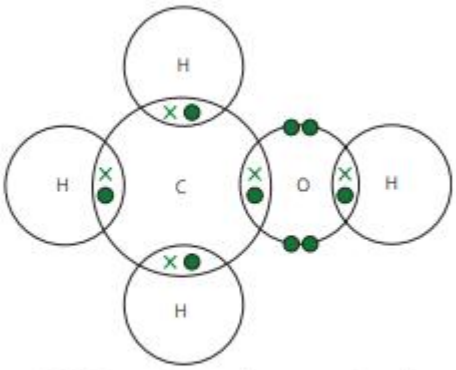
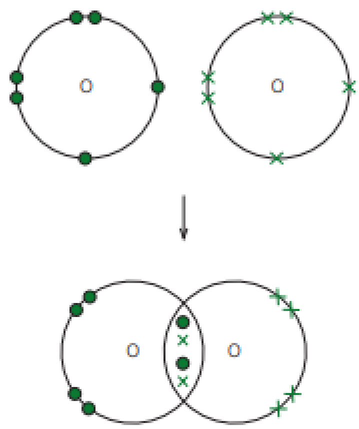


Figure 3.19 Dot-and-cross diagram to show the formation of a water molecule



▲ **Figure 3.25** Dot-and-cross diagram to show the formation of a methanol molecule



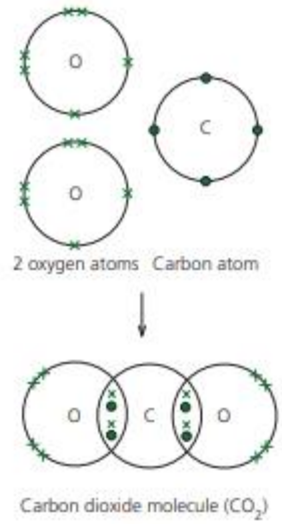
▲ **Figure 3.23** Dot-and-cross diagram to show the formation of an oxygen molecule

Covalent structures

Compounds containing covalent bonds have molecules whose structures can be classified as either **simple molecular** or **giant covalent.**

Properties of covalent compounds

Covalent compounds have the following properties



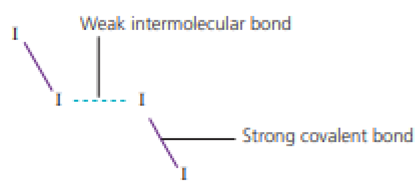
▲ **Figure 3.21** Dot-and-cross diagram to show the formation of a carbon dioxide molecule

O═C═O

a Carbon dioxide molecule. Note the double covalent bond is represented by a double line

b Model of the linear carbon dioxide molecule

▲ **Figure 3.22**



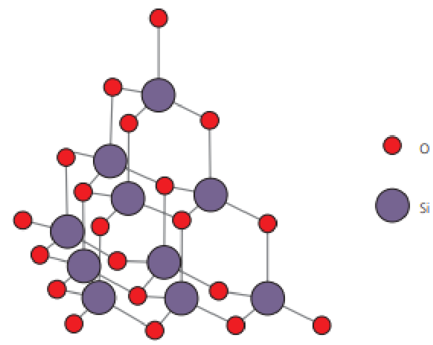
▲ **Figure 3.26** Strong covalent and weak intermolecular forces in iodine

»Simple molecular compounds, such as gases, liquids, or solids, have low melting and boiling points due to weak intermolecular forces of attraction. Covalent substances have higher melting points due to strong bonds. Ionic compounds have stronger interionic forces.

» Generally, they do not conduct electricity when molten or dissolved in water. This is because they do not contain ions. However, some molecules react with water to form ions

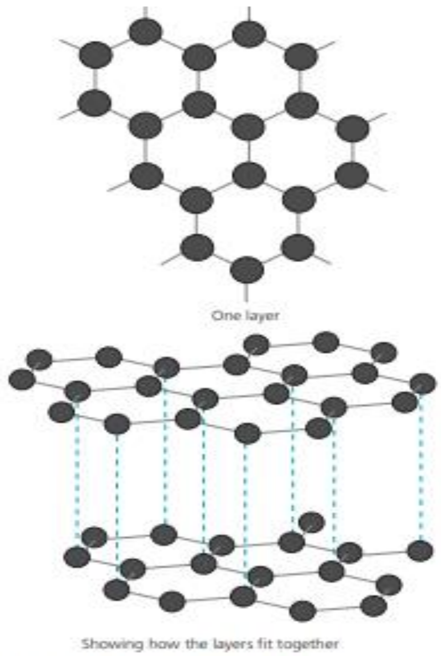
» Generally, they do not dissolve in water. However, water is an excellent solvent and can interact with and dissolve some covalent molecules better than others. Covalent substances are generally soluble in organic solvents

b Quartz is a hard solid at room temperature. It has a melting point of 1610°C and a boiling point of 2230°C

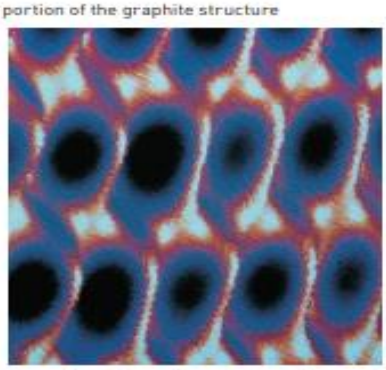


a The silicon(IV) oxide structure in quartz

▲ Figure 3.27

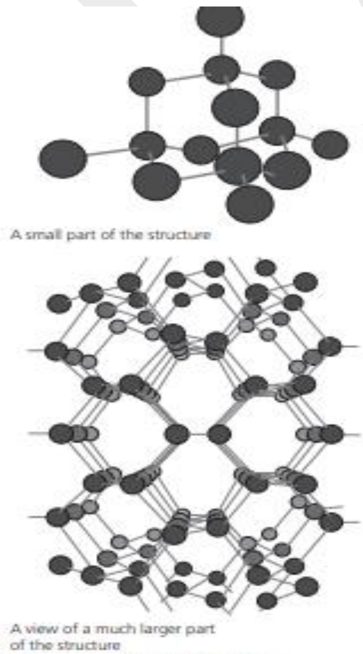


a A portion of the graphite structure



b A piece of graphite as imaged through a scanning tunnelling microscope

▲ Figure 3.28



a The structure of diamond

b A 100-carat, flawless diamond

▲ Figure 3.29

Graphite

Figure 3.28a demonstrates graphite's layer structure, with each layer bonded to three carbon atoms through strong covalent bonds. These layers pass easily, but there are weak attraction forces, allowing unbonded electrons to move freely, enabling graphite to conduct electricity.

Diamond

The diamond structure is a three-dimensional structure formed by covalent bonding of carbon atoms, similar to silicon(IV) oxide. This rigidity accounts for the extreme hardness of both substances. Diamond can be made by heating graphite to 300°C at high pressures, while graphite can be made by heating coke and sand in an electric arc furnace.

▼ **Table 3.2** Physical properties of graphite and diamond

Property	Graphite	Diamond
Appearance	A dark grey, shiny solid	A colourless transparent crystal which sparkles in light
Electrical conductivity	Conducts electricity	Does not conduct electricity
Hardness	A soft material with a slippery feel	A very hard substance
Density (g/cm^3)	2.25	3.51

▼ **Table 3.3** Uses of graphite and diamond

Graphite	Diamond
Pencils	Jewellery
Electrodes	Glass cutters
Lubricant	Diamond-studded saws
	Drill bits
	Polishers

▲ **Figure 3.30** Uses of graphite (as a pencil 'lead' and as a lock lubricant) and diamond (as a toothed saw to cut marble and on a dentist's drill)

3.3 Metallic bonding

Metals have a stable electronic configuration due to the movement of electrons in their outer shell. When lost, these electrons form a lattice of positive ions, forming a metallic bond with strong electrostatic forces.

Properties of metals

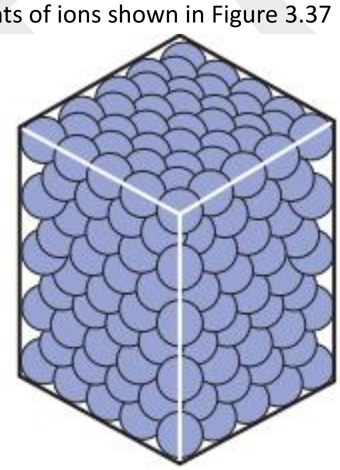
Metals have the following properties.

» They conduct electricity due to the mobile electrons within the metal structure. When a metal is connected in a circuit, the electrons move towards the positive terminal while at the same time electrons are fed into the other end of the metal from the negative terminal

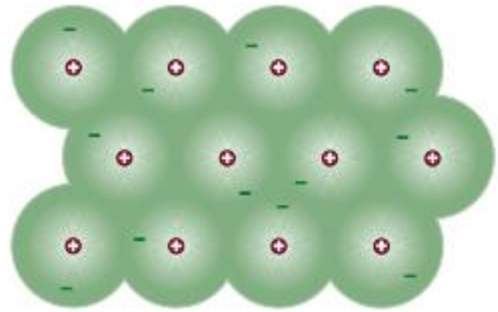
»Metals are malleable and ductile, with strong bonds that can be bent or hammered into various shapes. These bonds are not rigid but still re-form when a force is applied.

» They usually have high melting and boiling points due to the strong attraction between the positive metal ions and the mobile 'sea' of electrons.

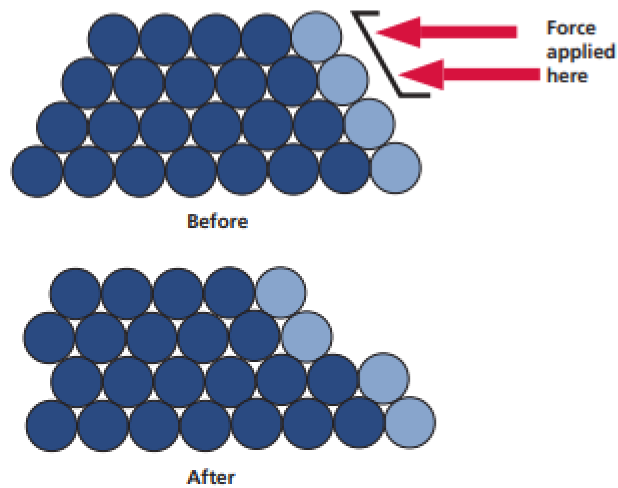
» They have high densities because the atoms are very closely packed in a regular manner, as can be seen in Figure 3.36. Different metals have different types of packing of atoms and in doing so they produce the arrangements of ions shown in Figure 3.37



▲ **Figure 3.36** Arrangement of ions in the crystal lattice of a metal



▲ **Figure 3.34** Metals consist of positive ions surrounded by a 'sea' of delocalised electrons



▲ **Figure 3.35** The positions of the positive ions in a metal before and after a force has been applied

Revision questions

1) Potassium iodide is an ionic compound.
i) Describe what happens, in terms of electron loss and gain, when a potassium atom reacts with an iodine atom.
ii) Describe the structure of solid potassium iodide. You may draw a diagram.
iii) Explain why potassium iodide has a high melting point.

b) Potassium iodide and lead nitrate are both soluble. Lead iodide is insoluble.
i) Describe how a pure dry sample of lead iodide could be made from solid potassium iodide and solid lead nitrate.
ii) Write an ionic equation for the formation of lead iodide, PbI2, when potassium iodide and lead nitrate react with each other.
State symbols are not required.

c) When chlorine gas is bubbled through an aqueous solution of potassium iodide, a redox reaction takes place.

$$2I^- + Cl_2 \rightarrow I_2 + 2Cl^-$$

i)State the colour change expected in this reaction.
ii) Identify the reducing agent in this reaction. Explain your answer.

2) a) Iron pyrite, FeS_2, is known as Fool's Gold because it is a shiny yellow solid which is similar in appearance to gold. Iron pyrite is an ionic compound. Gold is a
metallic element.
Iron pyrite, FeS_2, contains positive and negative ions. The positive ion is Fe^{2+}. Deduce the formula of the negative ion.

b) A student is provided with a sample of iron pyrite and a sample of gold. Suggest how the student could distinguish between the two substances.

c) Sulfur dioxide is produced on a large scale by heating iron pyrite strongly in air. The iron pyrite reacts with oxygen in the air producing iron (III) oxide, Fe_2O_3, and sulfur dioxide.
Construct a chemical equation for the reaction between iron pyrite and oxygen.

3) Nitrogen can form ionic compounds with reactive metals and covalent compounds with non-metals.
Nitrogen reacts with lithium to form the ionic compound lithium nitride, Li_3N.

i)Write the equation for the reaction between lithium and nitrogen.

ii) Lithium nitride is an ionic compound. Draw a diagram which shows its formula, the charges on the ions and the arrangement of the valency electrons around the negative ion.
Use x for an electron from a lithium atom.
Use o for an electron from a nitrogen atom.

4) The table below includes information about some of the elements in Period 2.

Element	Carbon	Nitrogen	Fluorine	Neon
Symbol	C	N	F	Ne
Structure	macromolecular	simple molecules N_2	simple molecules F_2	single atoms Ne
Boiling point / °C	4200	-196	-188	-246

Why does neon exist as single atoms but fluorine exists as molecules?

b) What determines the order of the elements in a period?

c)When liquid nitrogen boils the following change occurs.

$$N_2\,(l) \rightarrow N_2\,(g)$$

The boiling point of nitrogen is very low even though the bond between the atoms in a nitrogen molecule is very strong. Suggest an explanation.

d) Draw a diagram showing the arrangement of the outer shell electrons in a molecule of nitrogen.

5) Strontium and sulfur chlorides both have a formula of the type XCl_2 but they have different properties.

Property	Strontium chloride	Sulfur chloride
Appearance	white crystalline solid	red liquid
Melting point / °C	87	-8
Particles present	ions	molecules
Electrical conductivity of solid	poor	poor
Electrical conductivity of liquid	good	poor

The formulae of the chlorides are similar because both elements have a valency of 2.
Explain why Group II and Group VI elements both have a valency of 2.

b) Draw a diagram showing the arrangement of the outer electrons in one covalent molecule of sulfur chloride. Use x to represent an electron from a sulfur atom. Use o to represent an electron from a chlorine atom.

Explain the difference in electrical conductivity between the following.
i)Solid and liquid strontium chloride
ii) Liquid strontium chloride and liquid sulfur chloride

6) a) Selenium and sulfur are in Group VI. They have similar properties.
One of the main uses of selenium is in photoelectric cells. These cells can change light into electrical energy.
i)Name a process which can change light into chemical energy.
ii) Name a device which can change chemical energy into electrical energy.

b) The electron distribution of a selenium atom is 2+8+18+6.

i)Selenium forms an ionic compound with potassium. Draw a diagram which shows the formula of this ionic compound, the charges on the ions and the arrangement of the outer electrons around the negative ion.
Use o to represent an electron from an atom of potassium.
Use x to represent an electron from an atom of selenium.

ii) Draw a diagram showing the arrangement of the outer electrons in one molecule of the covalent compound selenium chloride.
Use x to represent an electron from an atom of selenium. Use o to represent an electron from an atom of chlorine.

iii) Predict two differences in the physical properties of these two compounds.

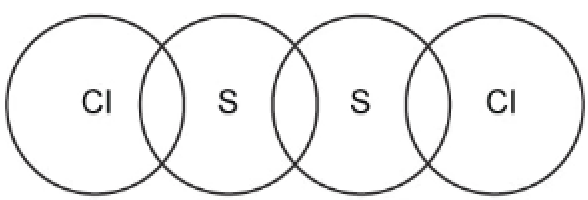
c) The selenide ion reacts with water.

$$Se^{2-} + H_2O \rightarrow HSe^- + OH^-$$

What type of reagent is the selenide ion in this reaction? Give a reason for your choice.

7) a) Sulfur monochloride, S_2Cl_2, has a simple molecular structure.
It was suggested that the structure of sulfur monochloride included a double covalent bond between the two sulfur atoms.
Suggest what is meant by the term double covalent bond?

b) After analysis of the structure of sulfur monochloride, S2Cl2, it was found that only single covalent bonds were present in its structure.
Complete the diagram below to show the bonding arrangement in a molecule of sulfur monochloride.
Show the outer shell electrons only.



c) Predict two properties of sulfur monochloride.
3d) Sulfur monochloride can be produced in the laboratory by feeding chlorine gas into a flask containing sulfur in its elemental form, S_8
Write a balanced symbol equation for this reaction.

8) a) Methane, CH_4, and propane, C_3H_8, are both simple molecular compounds that are gases at room temperature. Predict which of the two compounds has the higher boiling point.
Explain your answer.

b) A carbon atom can form up to four covalent bonds. A hydrogen atom cannot form four hydrogen bonds. State how many covalent bonds a hydrogen atom can form and explain why there is a difference.

c) Germanium also will form up to four covalent bonds.
Draw the dot-and-cross diagram of the simple molecule digermane, Ge2H6.

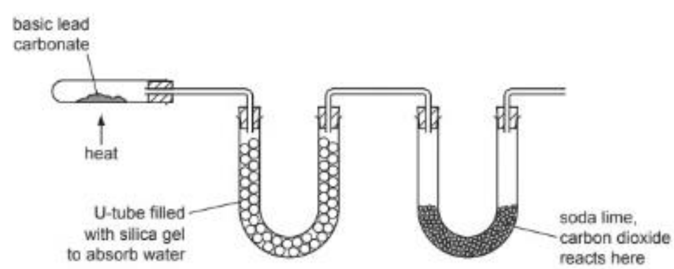
d)Explain why digermane does not conduct electricity.

9) a) Lead is an excellent roofing material. It is malleable and resistant to corrosion. Lead rapidly becomes coated with basic lead carbonate which protects it from further corrosion.
Lead has a typical metallic structure which is a lattice of lead ions surrounded by a 'sea' of mobile electrons. This structure is held together by attractive forces called a metallic bond.

i) Explain why there are attractive forces in a metallic structure.

ii)Explain why a metal, such as lead, is malleable.

b) Basic lead (II) carbonate is heated in the apparatus shown below. Water and carbon dioxide are produced.



Silica gel absorbs water. Silica gel often contains anhydrous cobalt (II) chloride.
When this absorbs water, it changes from blue to pink.
Suggest a reason.

c) Basic lead(II) carbonate has a formula of the type xPbCO3.yPb(OH)2 where x and y are whole numbers.

Determine x and y from the following information.

$$PbCO_3 \rightarrow PbO + CO_2$$
$$Pb(OH)_2 \rightarrow PbO + H_2O$$

When heated, the basic lead(II) carbonate gave 2.112 g of carbon dioxide and 0.432 g of water.

Mass of one mole of CO_2 = 44 g
Mass of one mole of H_2O = 18 g
Number of moles of CO_2 formed =
Number of moles of H_2O formed.
X= and y =
Formula of basic lead (ii) carbonate is …………..

10) a) An ore of the important metal zinc is zinc blende, ZnS. This is changed into zinc oxide which is reduced to the impure metal by carbon reduction.

i)How is zinc oxide obtained from zinc sulfide?

ii)Write a balanced equation for the reduction of zinc oxide by carbon.

iii) The major impurity in the zinc is cadmium. The boiling point of zinc is 907°C and that of cadmium is 767 °C. Name a technique which could be used to separate these two metals.

b) In common with most metals, zinc is a good conductor of electricity. It is used as an electrode in cells.

i)Give two other uses of zinc.

ii) Describe the metallic bonding in zinc and then explain why it is a good conductor of electricity.