

# 3

## Bonding and structure

### FOCUS POINTS

- ★ In what different ways do elements combine to form compounds?
- ★ How do the bonds within a compound affect its properties?
- ★ What is a redox reaction?

In the previous chapter, you established that both metal and non-metal elements are made up of atoms, which can combine to form compounds. In this chapter you will look more closely at how different elements combine, or bond, to form compounds. You will see that, generally, when elements combine they can form either ionic bonds with one another and are known as ionic compounds, or they can form covalent bonds with one another and are known as covalent compounds. You will also look at a third type of bonding, found in metals only, called metallic bonding.

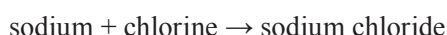
You will find out how the type of bonding affects the properties of these compounds. You will also learn more about formulae and how they can be obtained from given data and you will obtain an understanding of redox reactions and the processes associated with them.



▲ **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!

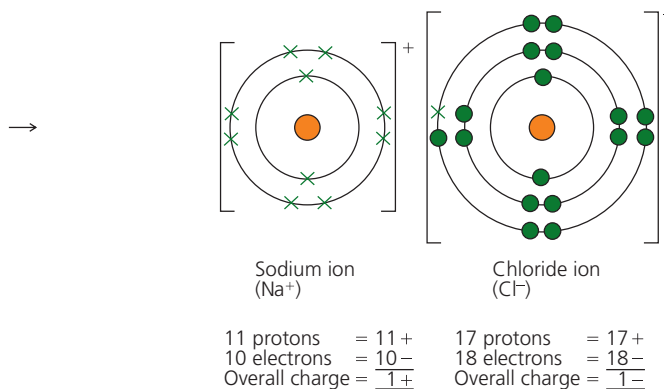
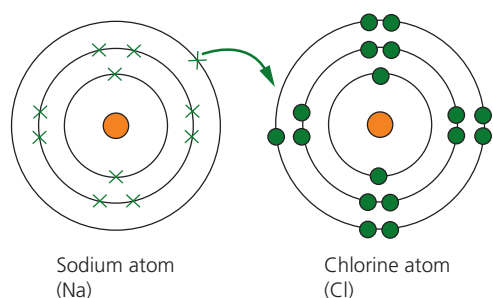
### 3.1 Ionic bonding

Ionic bonds are usually found in compounds that contain metals combined with non-metals. When this type of bond is formed, electrons are transferred from the metal atoms to the non-metal atoms during the chemical reaction. In doing this, the atoms become more stable due to their full outer shells. For example, consider what happens when sodium and chlorine react together and combine to make sodium chloride (Figure 3.1).



Sodium has just one electron in its outer shell ( $_{11}\text{Na}$  2,8,1). Chlorine has seven electrons in its outer shell ( $_{17}\text{Cl}$  2,8,7). When these two elements react, the outer electron of each sodium atom is transferred to the outer shell of a chlorine atom (Figure 3.2).

In this way both the atoms obtain full outer shells and become 'like' the noble gas nearest to them in the Periodic Table (see Figure 9.3, p. 135). One way of showing this transfer of electrons is by using a dot-and-cross diagram. In these diagrams,



▲ **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.

A further discussion of oxidation and reduction in terms of electron transfer takes place in Chapter 5 (p. 71).

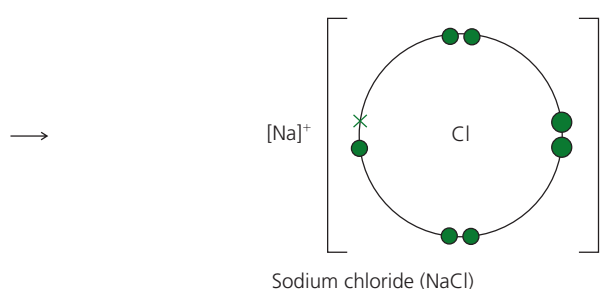
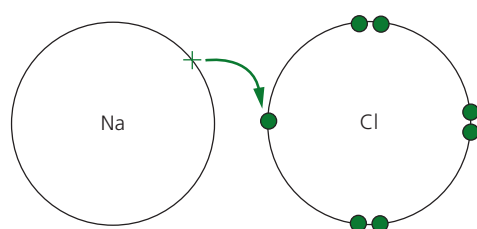
Only the outer electrons are important in bonding, so we can simplify the diagrams by missing out the inner shells (Figure 3.3).

The charges on the sodium and chloride ions are equal but opposite. They balance each other and the resulting formula for sodium chloride is NaCl. These oppositely charged ions attract each other and are pulled, or bonded, to one another by strong electrostatic forces. This type of bonding is called **ionic bonding**.

#### Key definition

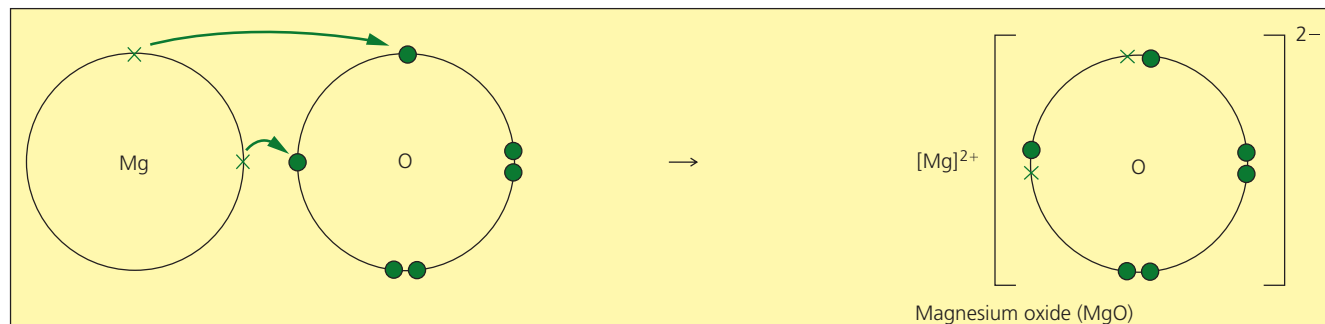
An **ionic bond** is a strong electrostatic attraction between oppositely charged ions.

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

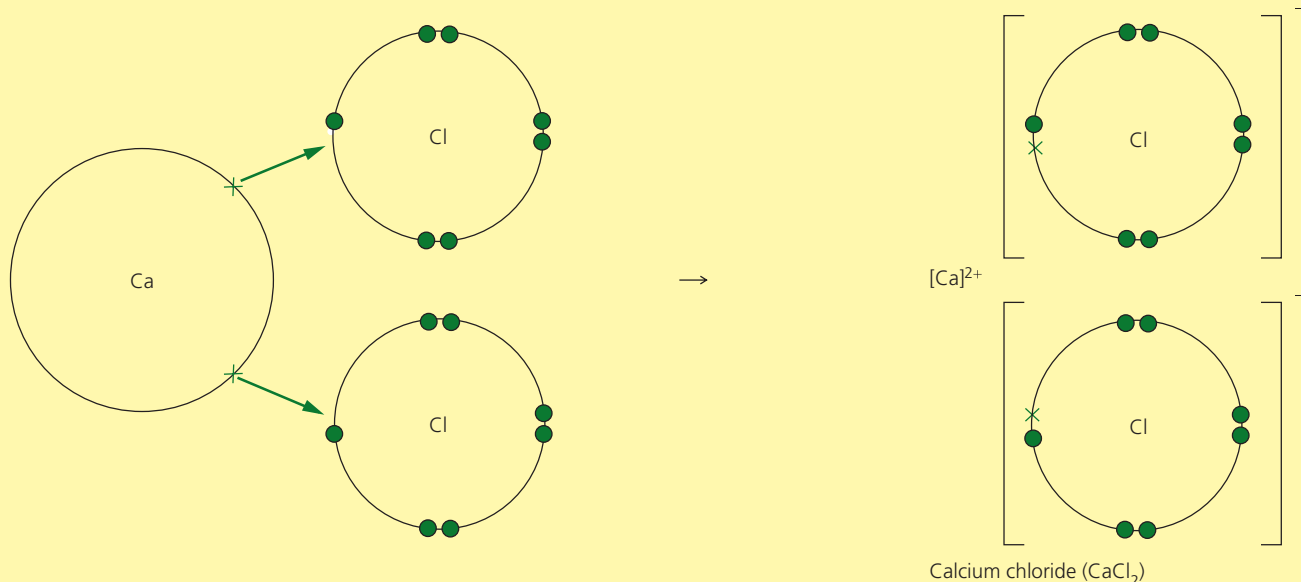
### 3 BONDING AND STRUCTURE



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

Figure 3.4 shows the electron transfers that take place between a magnesium atom and an oxygen atom during the formation of magnesium oxide.

Magnesium obtains a full outer shell by losing two electrons. These are transferred to the oxygen atom. In magnesium oxide, the  $\text{Mg}^{2+}$  and  $\text{O}^{2-}$  are oppositely charged and are attracted to one another. The formula for magnesium oxide is  $\text{MgO}$ .



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

When calcium and chlorine react, the calcium atom gives each of the two chlorine atoms one electron (Figure 3.5). In this case, a compound is

formed containing two chloride ions ( $\text{Cl}^-$ ) for each calcium ion ( $\text{Ca}^{2+}$ ). The chemical formula is  $\text{CaCl}_2$ .

#### Test yourself

- 1 Define the term 'ionic bond'.
- 2 Why do inert (noble) gases *not* bond with themselves and other elements?
- 3 Draw diagrams to represent the bonding in each of the following ionic compounds:
  - a potassium fluoride ( $\text{KF}$ )

- b lithium chloride ( $\text{LiCl}$ )
- c magnesium fluoride ( $\text{MgF}_2$ )
- d calcium oxide ( $\text{CaO}$ ).

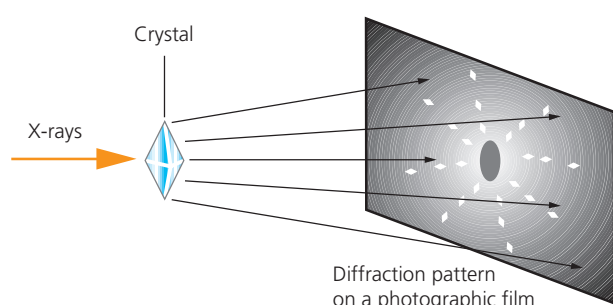
## ➔ Going further

Scientists, using **X-ray diffraction** (Figure 3.6a), have obtained photographs that indicate the way in which the ions are arranged (Figure 3.6b). The electron density map of sodium chloride is shown in Figure 3.6c.

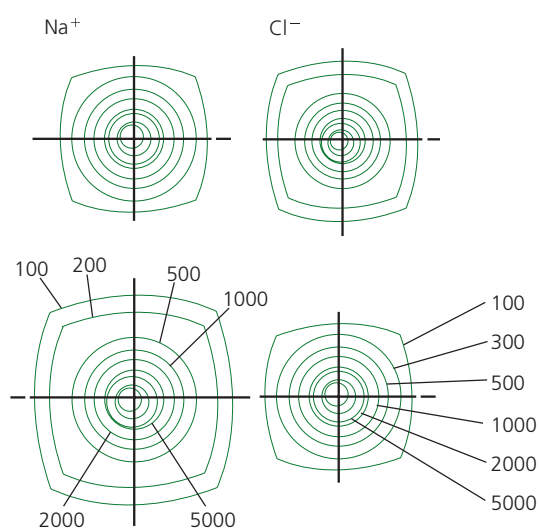
Figure 3.6d shows the structure of sodium chloride as determined by the X-ray diffraction technique. The study of crystals using X-ray diffraction was pioneered by Sir William Bragg and his son Sir Lawrence Bragg in 1912. X-rays are a form of electromagnetic radiation. They have a much shorter wavelength than light therefore it

is possible to use them to investigate extremely small structures.

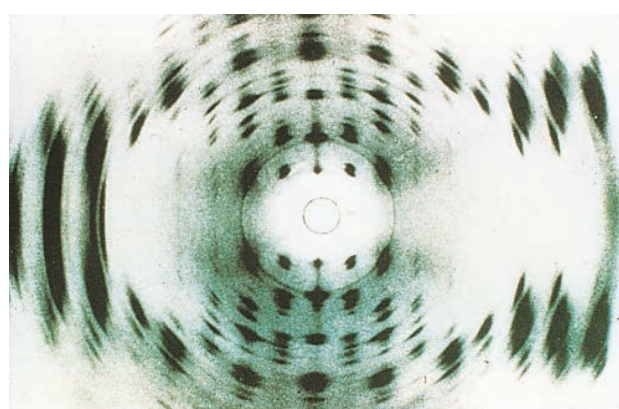
When X-rays are passed through a crystal of sodium chloride, for example, you get a pattern of spots called a diffraction pattern (Figure 3.6b). This pattern can be recorded on photographic film and used to work out how the ions or atoms are arranged in the crystal. Crystals give particular diffraction patterns depending on their structure, and this makes X-ray diffraction a particularly powerful technique in the investigation of crystal structures.



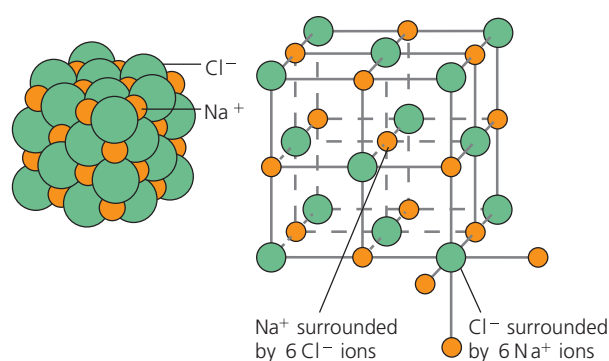
a X-ray diffraction technique



c Electron density map of sodium chloride



b X-ray diffraction photograph of sodium chloride



d The structure of sodium chloride

▲ Figure 3.6

## Ionic structures

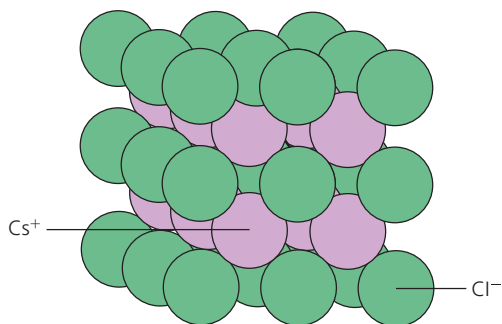
Ionic structures are solids at room temperature and have high melting and boiling points.

The ions are packed together in a regular arrangement called a **lattice**. Within the lattice, oppositely charged ions attract one another strongly.

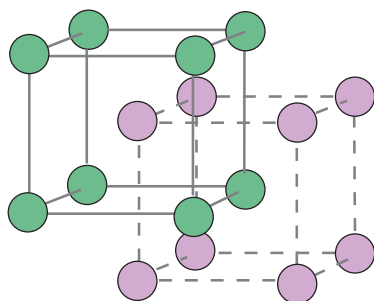
Figure 3.6d shows only a tiny part of a small crystal of sodium chloride. Many millions of sodium ions and chloride ions would be arranged in this way in a crystal of sodium chloride to make up the **giant ionic lattice** structure. Each sodium ion in the lattice is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions.

### ➔ Going further

Not all ionic substances form the same structures. Caesium chloride ( $\text{CsCl}$ ), for example, forms a different structure due to the larger size of the caesium ion compared with that of the sodium ion. This gives rise to the structure shown in Figure 3.7, which is called a body-centred cubic structure. Each caesium ion is surrounded by eight chloride ions and, in turn, each chloride ion is surrounded by eight caesium ions.



a The arrangement of ions



$\text{Cs}^+$  surrounded by 8  $\text{Cl}^-$  ions  
In the same way, each  $\text{Cl}^-$  is surrounded by 8  $\text{Cs}^+$  ions

b The unit cell

▲ **Figure 3.7** The structure of caesium chloride

## 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. A lot of energy is therefore needed to separate the ions and melt the substance.

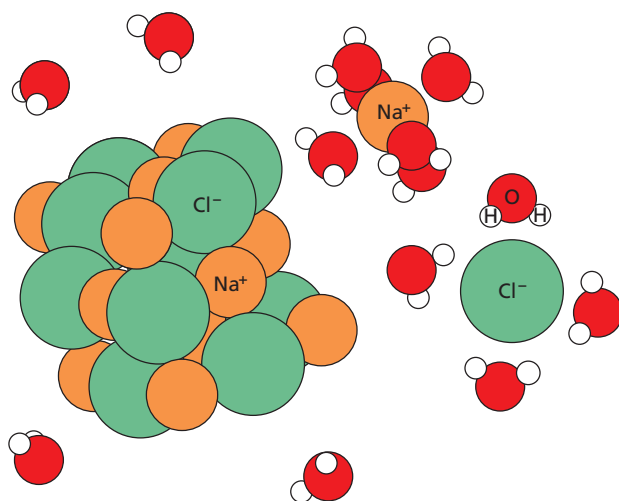
- » They are usually hard substances.
- » They mainly dissolve in water.

This is because water molecules are able to bond with both the positive and the negative ions, which breaks up the lattice and keeps the ions apart. Figure 3.8 shows the interaction between water molecules (the solvent) and sodium and chloride ions from sodium chloride (the solute). For a further discussion of the solubility of substances see, Chapter 8, p. 120.

- » 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. This allows an electric current to be passed through the molten compound (see Chapter 5, p. 71).

- » 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

## Test yourself

4 Link the terms in the boxes on the left with the definitions on the right.

|                     |  |
|---------------------|--|
| <b>Oxidation</b>    | An atom or group of atoms which has lost one or more electrons             |
| <b>Reduction</b>    | Involves loss of electrons   |
| <b>Redox</b>        | A strong electrostatic force of attraction between oppositely charged ions |
| <b>Negative ion</b> | Involves gain of electrons   |
| <b>Positive ion</b> | Involves both oxidation and reduction                                      |
| <b>Ionic bond</b>   | An atom or group of atoms which has gained one or more electrons           |

## Formulae of ionic substances

We saw previously (see p. 31) that ionic compounds contain positive and negative ions, whose charges balance. For example, sodium chloride contains one  $\text{Na}^+$  ion for every  $\text{Cl}^-$  ion, giving rise to the formula  $\text{NaCl}$ . This method can be used to write down formulae which show the ratio of the number of ions present in any ionic compound.

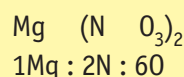
The formula of magnesium chloride is  $\text{MgCl}_2$ . This formula is arrived at by each  $\text{Mg}^{2+}$  ion combining with two  $\text{Cl}^-$  ions, and once again the charges balance. The use of oxidation numbers is useful when working out the formula of a compound. An oxidation number shows how oxidised or reduced an ion is compared to its atom.  $\text{Na}^+$  has an oxidation number of +1 because it is formed by the loss of one electron (Chapter 5, p. 71) from a sodium atom but  $\text{Mg}^{2+}$  has an oxidation number of +2 because it is formed when a magnesium atom loses two electrons (Table 3.1).  $\text{Na}^+$  can bond (combine) with only one  $\text{Cl}^-$  ion, whereas  $\text{Mg}^{2+}$  can bond with two  $\text{Cl}^-$  ions. A chloride ion has an oxidation state of -1 because it is formed when a chlorine atom gains one extra electron.

Some elements, such as copper and iron, possess two ions with different oxidation numbers. Copper

can form the  $\text{Cu}^+$  ion and the  $\text{Cu}^{2+}$  ion, with oxidation numbers of 1 and 2 respectively. Therefore it can form two different compounds with chlorine,  $\text{CuCl}$  and  $\text{CuCl}_2$ . We can also distinguish the difference by using Roman numerals in their names:  $\text{CuCl}$  is copper(I) chloride and  $\text{CuCl}_2$  is copper(II) chloride. Similarly, iron forms the  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  ions and so can also form two different compounds with, for example, chlorine:  $\text{FeCl}_2$  (iron(II) chloride) and  $\text{FeCl}_3$  (iron(III) chloride).

Table 3.1 shows the oxidation numbers of a series of ions you will normally meet in your study of chemistry.

You will notice that Table 3.1 includes groups of atoms which have net charges. For example, the nitrate ion is a single unit composed of one nitrogen atom and three oxygen atoms, and has one single negative charge. The formula, therefore, of magnesium nitrate would be  $\text{Mg}(\text{NO}_3)_2$ . You will notice that the  $\text{NO}_3$  has been placed in brackets with a 2 outside the bracket. This indicates that there are two nitrate ions present for every magnesium ion. The ratio of the atoms present is therefore:



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

|        | Oxidation numbers (valencies) |                   |            |                      |           |                      |
|--------|-------------------------------|-------------------|------------|----------------------|-----------|----------------------|
|        | 1                             |                   | 2          |                      | 3         |                      |
| Metals | Lithium                       | ( $\text{Li}^+$ ) | Magnesium  | ( $\text{Mg}^{2+}$ ) | Aluminium | ( $\text{Al}^{3+}$ ) |
|        | Sodium                        | ( $\text{Na}^+$ ) | Calcium    | ( $\text{Ca}^{2+}$ ) | Iron(III) | ( $\text{Fe}^{3+}$ ) |
|        | Potassium                     | ( $\text{K}^+$ )  | Copper(II) | ( $\text{Cu}^{2+}$ ) |           |                      |
|        | Silver                        | ( $\text{Ag}^+$ ) | Zinc       | ( $\text{Zn}^{2+}$ ) |           |                      |
|        | Copper(I)                     | ( $\text{Cu}^+$ ) | Iron(II)   | ( $\text{Fe}^{2+}$ ) |           |                      |
|        |                               |                   | Lead       | ( $\text{Pb}^{2+}$ ) |           |                      |
|        |                               |                   | Barium     | ( $\text{Ba}^{2+}$ ) |           |                      |



### 3 BONDING AND STRUCTURE

|                 | Oxidation numbers (valencies)                      |  |  |
|-----------------|--|--|--|
|                 | 1  | 2  | 3  |
| Non-metals      | Fluoride (F <sup>-</sup> )                         | Oxide (O <sup>2-</sup> )                                       |  |
|                 | Chloride (Cl <sup>-</sup> )                        | Sulfide (S <sup>2-</sup> )                                     |  |
|                 | Bromide (Br <sup>-</sup> )                         |  |  |
|                 | Hydrogen (H <sup>+</sup> )                         |  |  |
| Groups of atoms | Hydroxide (OH <sup>-</sup> )                       | Carbonate (CO <sub>3</sub> <sup>2-</sup> )                     | Phosphate (PO <sub>4</sub> <sup>3-</sup> ) |
|                 | Nitrate (NO <sub>3</sub> <sup>-</sup> )            | Sulfate (SO <sub>4</sub> <sup>2-</sup> )                       |  |
|                 | Ammonium (NH <sub>4</sub> <sup>+</sup> )           | Dichromate(VI) (Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup> ) |  |
|                 | Hydrogencarbonate (HCO <sub>3</sub> <sup>-</sup> ) |  |  |
|                 | Manganate (VII) (MnO <sub>4</sub> <sup>-</sup> )   |  |  |

## Oxidation number

Each atom in an element or compound is assigned an oxidation number to show how much it is reduced or oxidised. 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. For example, the name for FeCl<sub>3</sub> is iron(III) chloride and not iron(3) chloride.
- » The oxidation number of the free element is always 0, for example, in metals such as zinc and copper.
- » 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 I<sub>2</sub> but an oxidation number of -1 in I<sup>-</sup>.
- » Compounds have no charge overall. Hence the oxidation numbers of all the individual elements in a compound must add up to 0. The oxidation numbers of elements in compounds can vary, as seen in Table 3.1. It is possible to recognise which of the different oxidation numbers a metal element is in by the colour of its compounds (Figure 3.9).



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

- » An increase in the oxidation number, for example, from +2 to +3 as in the case of Fe<sup>2+</sup> to Fe<sup>3+</sup>, is **oxidation**.
- » However, a reduction in the oxidation number, for example, from +6 to +3 as in the case of Cr<sup>6+</sup> (in CrO<sub>4</sub><sup>2-</sup>) to Cr<sup>3+</sup>, is **reduction**.

#### Key definitions

**Oxidation** involves an increase in oxidation number.

**Reduction** involves a decrease in 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.

For example, if a dilute solution of acidified potassium manganate(VII) (pale purple) is added to a solution of iron(II) sulfate, a colour change takes place as the reaction occurs (Figure 3.10). The iron(II) sulfate (pale green) changes to pale yellow, showing the presence 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

In this reaction, the iron(II) ions have been oxidised to iron(III) ions (increase in oxidation number) and the manganate(VII) ions have been reduced to manganese(II) ions (decrease in oxidation number), which are very pale pink. Hence the manganate(VII) ions are the oxidising agent and the iron(II) ions are the reducing agent.

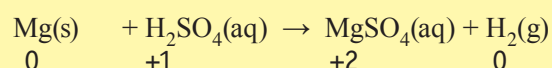
Potassium iodide is a common reducing agent. When it is oxidised, a colour change is also produced. The iodide ion ( $I^-$ ) is oxidised to iodine ( $I_2$ ). The colour of the resulting solution will change from colourless to yellow-brown. If you then add starch indicator, a test to show the presence of iodine, it will turn blue-black. Hence if potassium iodide solution is added to a solution of iron(III) solution followed by the starch indicator, then a blue-black colour will be seen as shown in Figure 3.11. This shows that the iron(III) ions must have been reduced to iron(II) ions as the iodide ions are oxidised to iodine.



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

It is possible to recognise redox processes by looking at the oxidation numbers on the two sides of the chemical equation for a reaction. For example, in the reaction between magnesium and dilute sulfuric acid, the magnesium dissolves and hydrogen gas is produced. Both magnesium metal and hydrogen gas are free elements and so have an oxidation number of 0. In sulfuric acid, hydrogen has an oxidation number of +1 since the overall charge on the sulfate ion is -2. Similarly, the oxidation number of magnesium in magnesium sulfate is +2.

magnesium + sulfuric acid  $\rightarrow$  magnesium sulfate + hydrogen



The sulfate ions are unchanged by the reaction and so can be ignored.

As you can see, the oxidation number of magnesium has increased from 0 to +2. Therefore the magnesium has been oxidised by the sulfuric acid and so sulfuric acid is the **oxidising agent**. The oxidation number of hydrogen has decreased from +1 in the sulfuric acid to 0 in the free element. Therefore the hydrogen has been reduced by the magnesium and so magnesium is the **reducing agent**.

#### 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.



### 3 BONDING AND STRUCTURE

#### Test yourself

- 5 Identify the oxidising and reducing agents in the following reactions.
- a  $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{ZnSO}_4\text{(aq)} + \text{H}_2\text{(g)}$   
b  $\text{Cu}^{2+}\text{(aq)} + \text{Zn(s)} \rightarrow \text{Cu(s)} + \text{Zn}^{2+}\text{(aq)}$   
c  $\text{Mg(s)} + \text{Cu(NO}_3)_2\text{(aq)} \rightarrow \text{Mg(NO}_3)_2\text{(aq)} + \text{Cu(s)}$   
d  $\text{Fe(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{FeSO}_4\text{(aq)} + \text{H}_2\text{(g)}$
- 6 Using the information in Table 3.1, write the formulae for:
- a copper(I) oxide                      e potassium  
b zinc phosphate                      manganate(VII)  
c iron(III) chloride                      f sodium  
d lead(II) bromide                      dichromate(VI).
- 7 Using the formulae in your answer to question 6, write down the ratio of atoms present for each of the compounds.

#### Going further

##### The 'cross-over' method

A less scientific but simpler method to work out the formulae of compounds is called the 'cross-over' method. In this method it is only necessary to 'swap' the numerical value of the oxidation number of the elements or groups of atoms (or radicals) concerned. This is easily done by 'crossing over' the numerical value of their oxidation numbers, by placing the numerical value of the oxidation number of the first element after the symbol of the second and placing the numerical value of the oxidation number of the second element or radical after the symbol of the first.

For example, in aluminium sulfide the oxidation numbers of the elements are:

$$\text{Al} = 3 \quad \text{and} \quad \text{S} = 2$$

Hence the chemical formula is  $\text{Al}_2\text{S}_3$ .

Similarly, in sodium sulfate the oxidation numbers are:

$$\text{Na} = 1 \quad \text{and} \quad \text{SO}_4 = 2$$

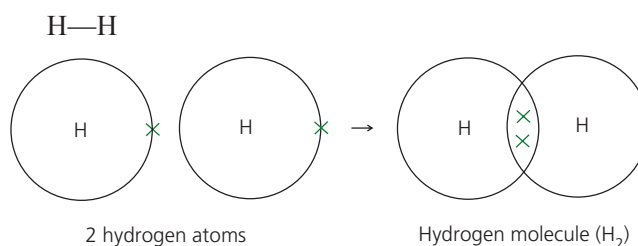
Hence the chemical formula is  $\text{Na}_2\text{SO}_4$ .

gain the stability of the noble (inert) gas electronic configuration. The simplest example of this type of bonding can be seen by considering the hydrogen molecule,  $\text{H}_2$ .

##### Key definition

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

Each hydrogen atom in the molecule has one electron. In order to obtain a full outer shell and gain an electronic configuration that is the same as the noble gas helium, each of the hydrogen atoms must have two electrons. To do this, the outer shells of the two hydrogen atoms overlap (Figure 3.12a). A molecule of hydrogen is formed, with two hydrogen atoms sharing a pair of electrons (Figure 3.12a, b). This shared pair of electrons is known as a single covalent bond and is represented by a single line as in hydrogen:



- a The electron sharing to form the single covalent bond in  $\text{H}_2$  molecules



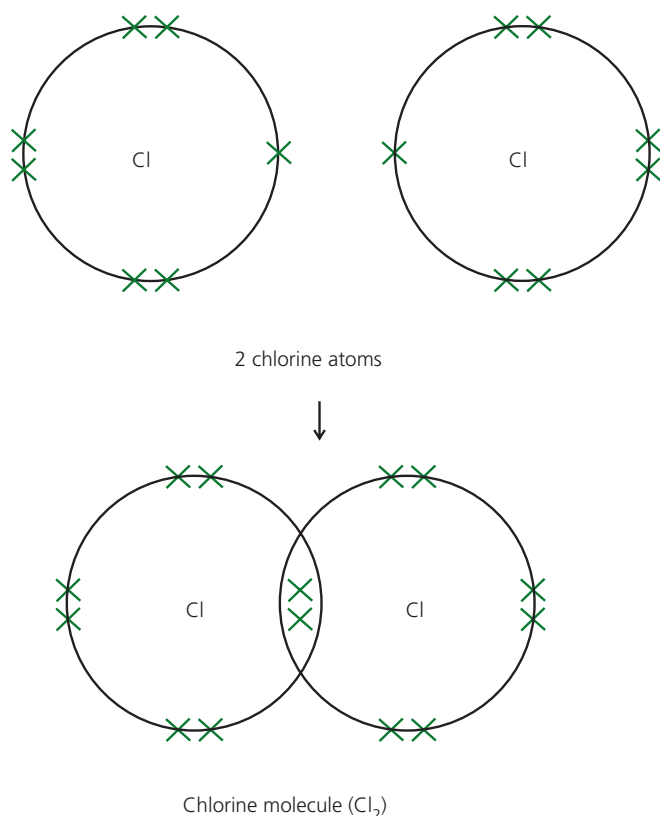
- b Model of a hydrogen molecule

#### Figure 3.12

A similar example exists in the diatomic halogen molecule chlorine,  $\text{Cl}_2$  (Figure 3.13).

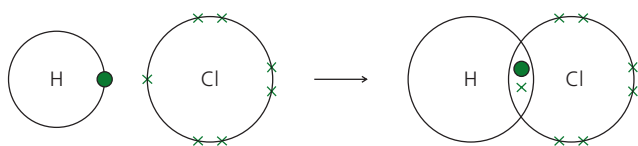
## 3.2 Covalent bonding

Another way in which atoms can form stable compounds is by sharing the electrons in their outer shells. This occurs between non-metal atoms, and the bond formed is called a **covalent bond**. During the bond formation, the atoms involved



▲ **Figure 3.13** The electron sharing to form the single covalent bond in  $\text{Cl}_2$  molecules (sharing outer electron shells only)

Another simple molecule is hydrogen chloride. This time two different elements share electrons to gain the electronic configuration of their nearest noble gas. In this case hydrogen, with only one electron in its outer shell, needs to share one more electron to gain the electronic configuration of helium. Chlorine has seven electrons in its outer shell so it also needs to share one more electron, this time to gain the electronic configuration of the noble gas argon (Figure 3.14).



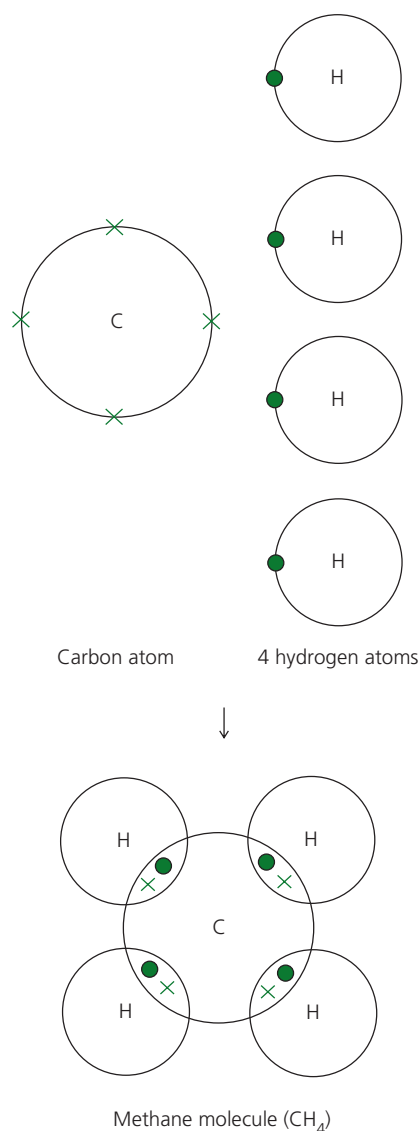
▲ **Figure 3.14** Dot-and-cross diagram to show the formation of a hydrogen chloride molecule.

## Other covalent compounds

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

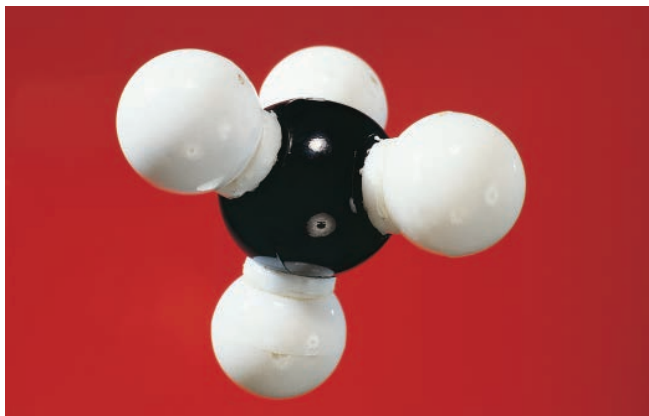


The carbon atom needs four more electrons to attain the electronic configuration of the noble gas neon. Each hydrogen atom needs only one electron to form the electronic configuration of helium. Figure 3.15 shows how the atoms gain these electronic configurations by the sharing of electrons. Note that only the outer electron shells are shown.



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

### 3 BONDING AND STRUCTURE



▲ **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:

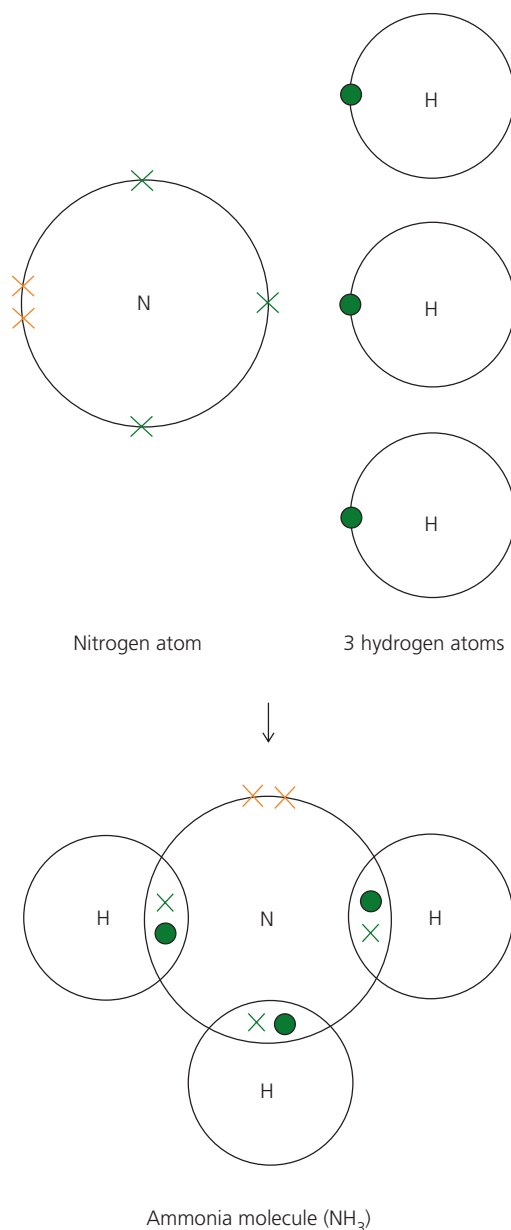


The nitrogen atom needs three more electrons to obtain the noble gas structure of neon. Each hydrogen requires only one electron to form the noble gas structure of helium. The nitrogen and hydrogen atoms share electrons, forming three single covalent bonds (Figure 3.17).

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



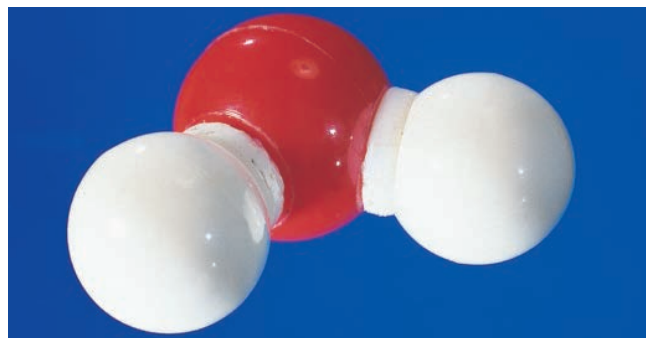
The oxygen atom needs two electrons to gain the electronic configuration of neon. Each hydrogen requires one more electron to gain the electronic configuration of helium. Again, the oxygen and hydrogen atoms share electrons, forming a water molecule with two single covalent bonds as shown in Figure 3.19.



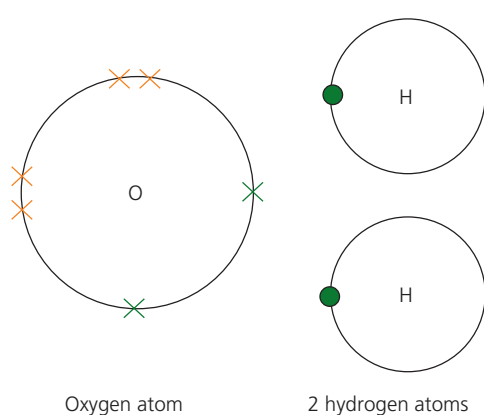
▲ **Figure 3.17** Dot-and-cross diagram to show the formation of an ammonia molecule



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



▲ **Figure 3.20** Model of a water molecule

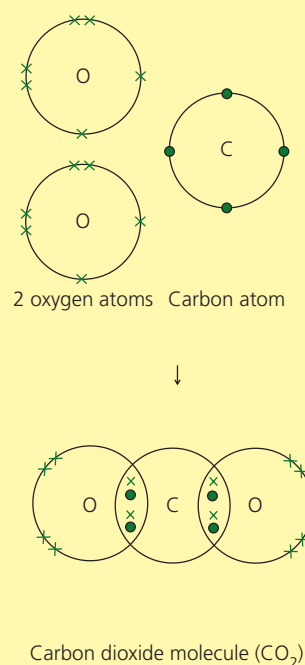


▲ **Figure 3.19** Dot-and-cross diagram to show the formation of a water molecule

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

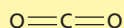


In this case each carbon atom needs to share four electrons to gain the electronic configuration of neon. Each oxygen needs to share two electrons to gain the electronic configuration of neon. This is achieved by forming two double covalent bonds in which two pairs of electrons are shared in each case, as shown in Figure 3.21. Carbon dioxide is a linear molecule (Figure 3.22).



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

### 3 BONDING AND STRUCTURE



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



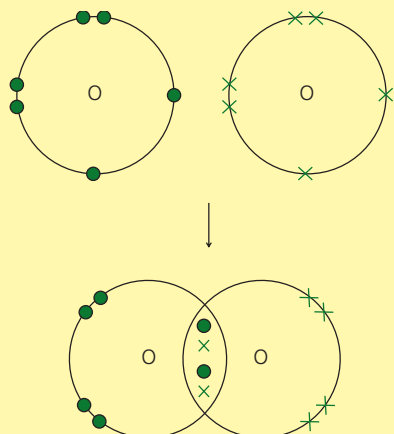
**b** Model of a linear carbon dioxide molecule

▲ **Figure 3.22**

Another molecule which contains a double covalent bond is that of oxygen gas,  $\text{O}_2$ . Each of the oxygen atoms has six electrons in its outer electron shell.



Each needs to share another two electrons to gain the electronic configuration of the noble gas neon.



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



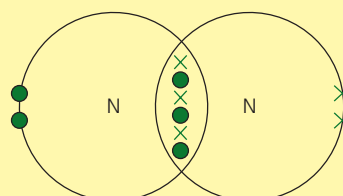
Nitrogen gas,  $\text{N}_2$ , is molecule which contains a triple covalent bond. A triple covalent bond is formed when three pairs of electrons are shared and it is represented by ' $\equiv$ '.

Each nitrogen atom has the electronic configuration shown below:



Both nitrogen atoms in the molecule need to share a further three electrons to gain the electronic configuration of the noble gas neon.

The triple bond can be shown in the molecule as  $\text{N}\equiv\text{N}$ .



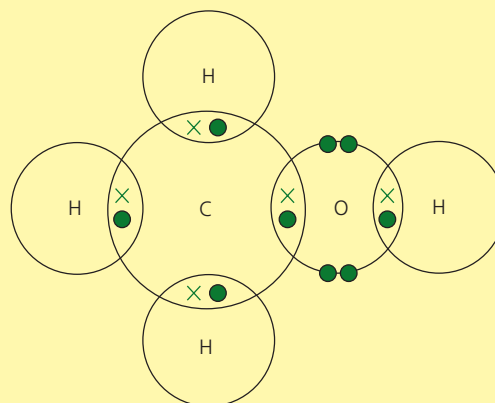
▲ **Figure 3.24** Dot-and-cross diagram to show the bonding in a nitrogen molecule

Methanol,  $\text{CH}_3\text{OH}$ , is a molecule which contains three different types of atom. When this happens and you need to draw a dot-and-cross diagram, simply make sure that atoms which bond with one another do not both have dots or crosses. In this molecule the atoms have these configurations:



The carbon atom needs to share four further electrons which it can do with the three hydrogens; each needs to share one more electron, and the oxygen atom. The oxygen atom, which shares one electron with the carbon atom, shares another electron with the remaining hydrogen atom. By doing this, the hydrogen atoms gain the electronic configuration of helium, and the carbon and oxygen atoms gain the electronic configuration of neon.

The dot-and-cross diagram for methanol is:



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

### Test yourself

8 What do you understand by the term 'covalent bond'?

9 Draw diagrams to represent the bonding in each of the following covalent compounds:

- a tetrachloromethane ( $\text{CCl}_4$ )
- b hydrogen sulfide ( $\text{H}_2\text{S}$ )

## Covalent structures

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

Simple molecular structures are simple, formed from only a few atoms. They have strong covalent bonds between the atoms within a molecule (**intramolecular bonds**) but have weak bonds between the molecules (**intermolecular forces**).

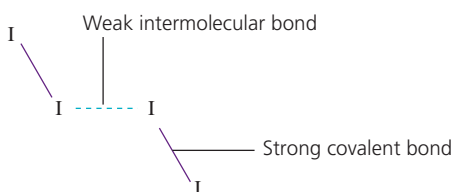


### Going further

Some of the strongest of these weak intermolecular forces occur between water molecules.

One type of weak bond between molecules is known as the van der Waals bond (or force), and these forces increase steadily with the increasing size of the molecule.

Examples of simple molecules are iodine (Figure 3.26), methane, water and ethanol.



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

Giant covalent structures contain many hundreds of thousands of atoms joined by strong covalent bonds. Examples of substances with this type of structure are diamond, graphite and silicon(IV) oxide (Figure 3.27).

## Properties of covalent compounds

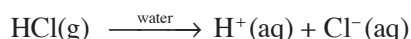
Covalent compounds have the following properties.

- » As simple molecular compounds, they are usually gases, liquids or solids with low melting and boiling points.

The melting points are low because of the weak intermolecular forces of attraction which exist between simple molecules. These are weaker compared to the strong covalent bonds. Giant covalent substances have higher melting points, because the whole structure is held together by strong covalent bonds.

It should be noted that in ionic compounds, the interionic forces are much stronger than the intermolecular forces in simple covalent substances and so the melting and boiling points are generally higher.

- » 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. For example, hydrogen chloride gas produces aqueous hydrogen ions and chloride ions when it dissolves in water:

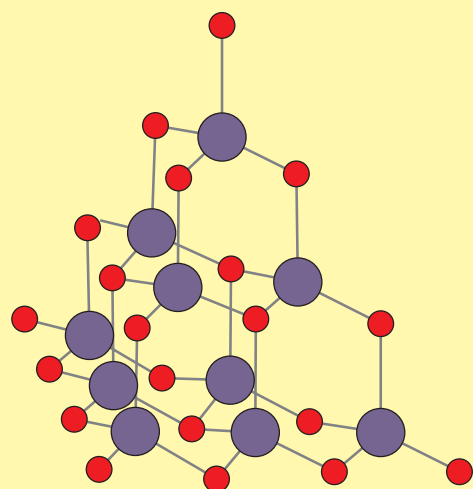


The presence of the ions allows the solution to conduct electricity.

- » 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.



### 3 BONDING AND STRUCTURE



a The silicon(IV) oxide structure in quartz



b Quartz is a hard solid at room temperature. It has a melting point of  $1610^{\circ}\text{C}$  and a boiling point of  $2230^{\circ}\text{C}$

▲ Figure 3.27



#### Going further

##### Giant covalent structures

When an element can exist in more than one physical form in the same physical state, it is said to exhibit allotropy (or polymorphism). Each of the different

physical forms is called an allotrope. Allotropy is actually quite a common feature of the elements in the Periodic Table (p. 135). An example of an element which shows allotropy is carbon.



#### Practical skills

##### Testing the electrical conductivity of ionic and covalent substances

**For safe experiments/demonstrations which are related to this chapter, please refer to the *Cambridge IGCSE Chemistry Practical Skills Workbook*, which is also part of this series.**

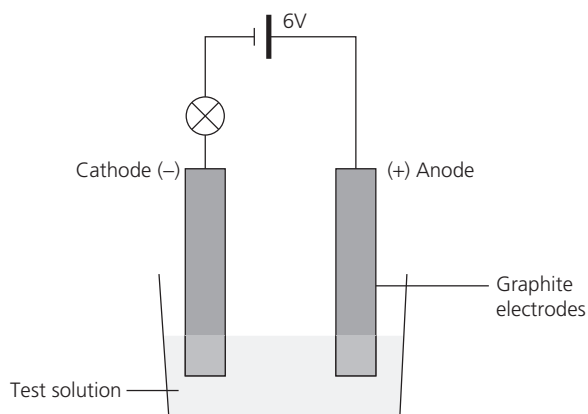
##### Safety

- Eye protection must be worn.
- Take care when handling dilute aqueous solutions: most are low hazard (and should all be  $\leq 1\text{M}$ );  $0.2\text{--}1.0\text{ mol/dm}^3$  copper(II) sulfate solution is an irritant and corrosive.



The apparatus was set up as shown in the diagram on the right to investigate the electrical conductivity of a number of compounds in aqueous solution.

- 1 Describe the experimental method you could use to carry out a fair test on a variety of ionic and covalent substances dissolved in water and using the apparatus shown in the diagram.



| Solution                   | Does the bulb light? | Does the substance in solution conduct electricity? | Does the substance in solution contain ionic or covalent bonding? |
|----------------------------|----------------------|---|---|
| Glucose ( $C_6H_{12}O_6$ ) | No                   |   |   |
| Potassium chloride         | Yes                  |   |   |
| Calcium chloride           | Yes                  |   |   |
| Sodium iodide              | Yes                  |   |   |
| Ethanol ( $C_2H_5OH$ )     | No                   |   |   |
| Copper(II) sulfate         | Yes                  |   |   |
| Pure water                 | No                   |   |   |

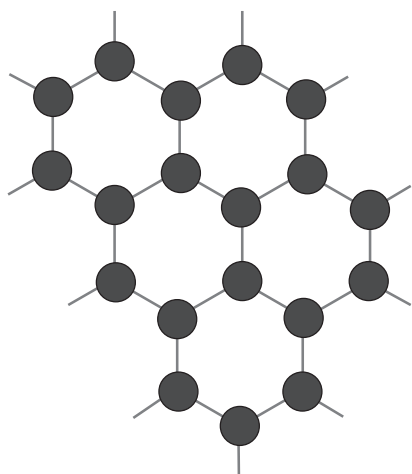
- 2 Copy and complete the results table above.
- 3 Using your answers to columns three and four in the results table, write a conclusion for this experiment. In your conclusion, state and explain any trends shown in the results.
- 4 Would the results be different if solid potassium chloride were used instead of potassium chloride solution? Explain your answer.
- 5 Predict and explain the results you would obtain for magnesium nitrate solution.
- 6 Predict and explain the results you would obtain with just distilled water – that is, no dissolved substance present.

## Different forms of carbon

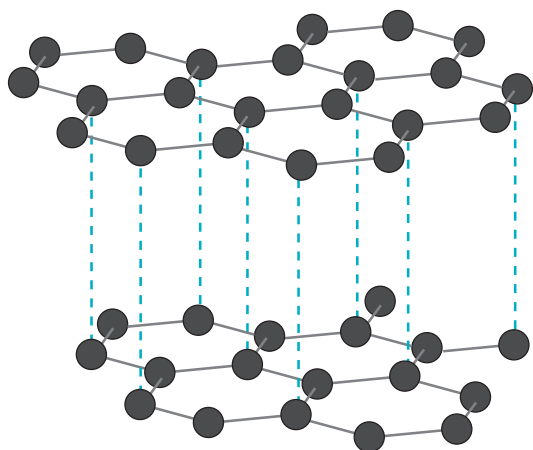
Carbon is a non-metallic element which exists in more than one solid structural form. These are graphite and diamond. Each of the forms has a different structure (Figures 3.28 and 3.29) and

so they exhibit different physical properties (Table 3.2). The different physical properties that they exhibit lead to graphite and diamond being used in different ways (Table 3.3 and Figure 3.30).

### 3 BONDING AND STRUCTURE

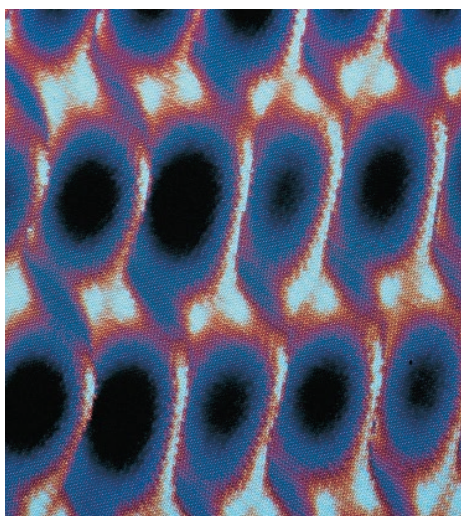


One layer



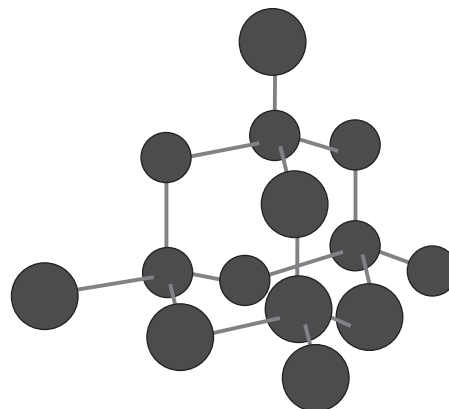
Showing how the layers fit together

**a** A portion of the graphite structure

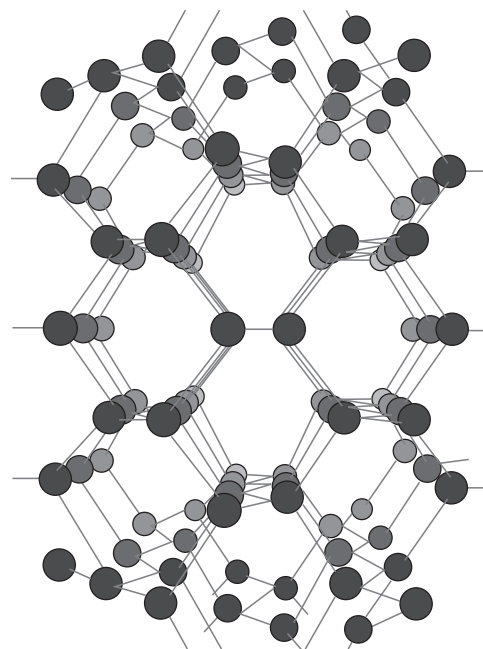


**b** A piece of graphite as imaged through a scanning tunnelling microscope

▲ Figure 3.28



A small part of the structure



A view of a much larger part of the structure

**a** The structure of diamond



**b** The Regent Diamond has been worn by Queen Elizabeth II

▲ Figure 3.29

▼ **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 <sup>3</sup> ) | 2.25                                 | 3.51   |

### Graphite

Figure 3.28a shows the structure of graphite. This is a layer structure. Within each layer, each carbon atom is bonded to three others by strong covalent bonds. Each layer is therefore like a giant molecule. Between these layers there are weak forces of attraction and so the layers will pass over each other easily.

With only three covalent bonds formed between carbon atoms within the layers, an unbonded electron is present on each carbon atom. These 'spare' (or **delocalised**) electrons form electron clouds between the layers and it is because of these spare electrons that graphite conducts electricity.

#### ➔ Going further

##### Graphitic compounds

In recent years, a set of interesting compounds known as graphitic compounds have been developed. In these compounds, different atoms have been fitted in between the layers of carbon atoms to produce a substance with a greater electrical conductivity than pure graphite. Graphite is also used as a component in certain sports equipment, such as tennis and squash rackets.

#### ➔ Going further

##### Graphene

Discovered in 2004, graphene is a so-called super material made up of single layers of graphite as shown in the upper diagram of Figure 3.28a. It is able to conduct electricity one million times better than copper metal and has enormous potential in electronics.

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

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

### Diamond

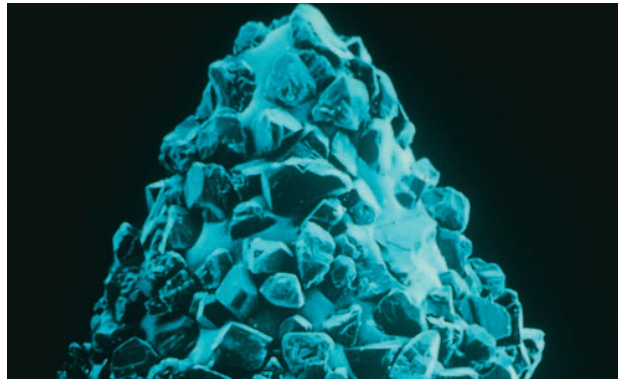
Figure 3.29a shows the diamond structure. Each of the carbon atoms in the giant structure is covalently bonded to four others. They form a tetrahedral arrangement similar to that found in silicon(IV) oxide (p. 44). This bonding scheme gives rise to a very rigid three-dimensional structure and accounts for the extreme hardness of the substances silicon(IV) oxide and diamond. All the outer shell electrons of the carbon atoms are used to form covalent bonds, so there are no electrons available to enable diamond or silicon(IV) oxide to conduct electricity.

It is possible to manufacture the different allotropes of carbon. Diamond is made by heating graphite to about 300°C at very high pressures. Diamond made by this method is known as industrial diamond. Graphite can be made by heating a mixture of coke and sand at a very high temperature in an electric arc furnace for about 24 hours.

The various uses of graphite and diamond result from their differing properties (Figure 3.30).



### 3 BONDING AND STRUCTURE



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

#### ➔ Going further

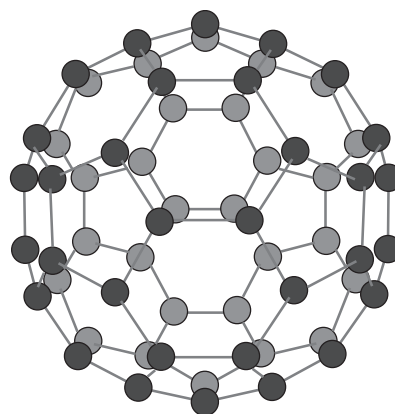
##### Buckminsterfullerene – an unusual form of carbon

In 1985, a new allotrope of carbon was obtained by Richard Smalley and Robert Curl of Rice University, Texas. It was formed by the action of a laser beam on a sample of graphite.

The structure of buckminsterfullerene can be seen in Figure 3.31.

This spherical structure is composed of 60 carbon atoms covalently bonded together. Further spherical forms of carbon, 'bucky balls', containing 70, 72 and 84 carbon atoms have been identified and the discovery has led to a whole new branch of inorganic carbon chemistry. It is thought that this type of molecule exists in chimney soot. Chemists have suggested that, due to the large surface area of the bucky balls, they may have uses as catalysts (Chapter 7, p. 106). Also they may have uses as superconductors.

Buckminsterfullerene is named after an American architect, Buckminster Fuller, who built complex geometrical structures (Figure 3.32).



▲ **Figure 3.31** Buckminsterfullerene – a 'bucky ball' ( $C_{60}$ )





▲ **Figure 3.32**  $C_{60}$  has a structure similar to a football and to the structure of the domes shown

### Test yourself

**10** Explain the difference between ionic and covalent bonding.

**11** Compare the structures of silicon(IV) oxide and diamond.

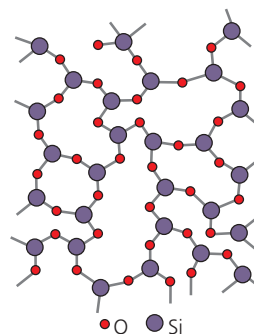
**12** Draw up a table to summarise the properties of the different types of substances you have met in this chapter. Your table should include examples from ionic substances and covalent substances (simple and giant).

### Going further

#### Glasses and ceramics

##### Glasses

Glasses are all around us. Glasses are irregular giant molecular structures held together by strong covalent bonds. Glass can be made by heating silicon(IV) oxide with other substances until a thick viscous liquid is formed. As this liquid cools, the atoms present cannot move freely enough to return to their arrangement within the pure silicon(IV) oxide structure. Instead they are forced to form a disordered arrangement as shown in Figure 3.33. Glass is called a supercooled liquid.



▲ **Figure 3.33** Two-dimensional structure of silicon(IV) oxide in glass



### 3 BONDING AND STRUCTURE

#### Ceramics

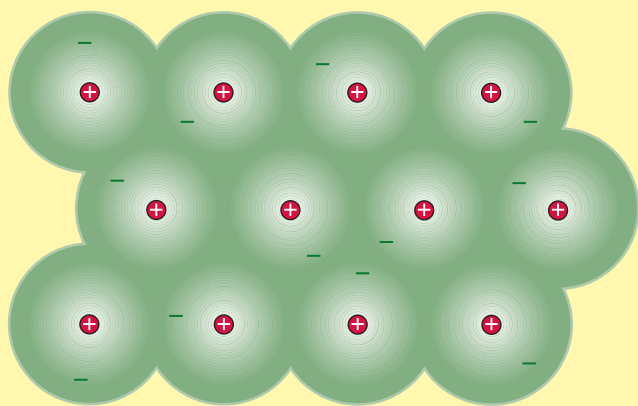
The word ceramic comes from the Greek word *keramos* meaning pottery or 'burnt stuff'. Clay dug from the ground contains a mixture of several materials including silicon(IV) oxide. During firing in a furnace, the clay is heated to a temperature of 1000°C. The material produced at the end of the firing, the ceramic, which is

a type of glass, consists of many tiny mineral crystals bonded together with glass.

There are now many uses of these ceramic materials, such as ceramic discs, which prevent short-circuits in the national grid pylons; ceramic tiles, which protect the spacecraft; and ceramic bearings that do not need lubrication.

## 3.3 Metallic bonding

Another way in which atoms obtain a more stable electronic configuration is found in metals. The electrons in the outer shell of the atom of a metal move freely throughout the structure. They are delocalised, forming a mobile 'sea' of electrons (Figure 3.34). When the metal atoms lose these electrons, they form a giant lattice of positive ions. Therefore, metals consist of positive ions embedded in a sea of moving electrons. The negatively charged electrons attract all the positive metal ions and bond them together with strong electrostatic forces of attraction as a single unit. This is the **metallic bond**.



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

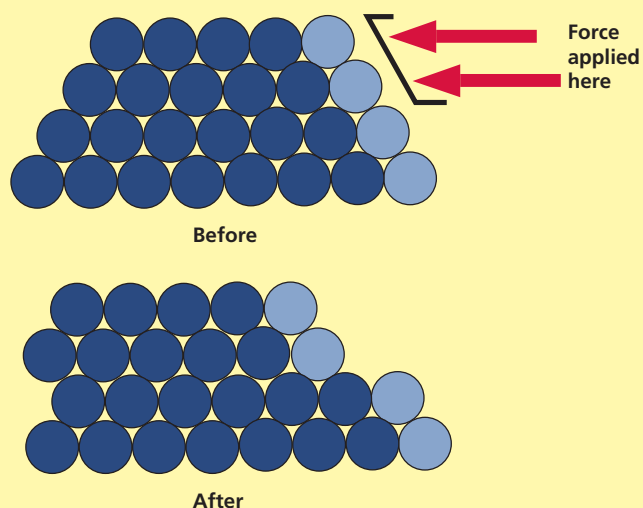
### Properties of metals

Metals have the following properties.

- » They conduct electricity due to the mobile electrons within the metal structure. When a

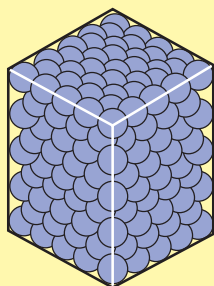
metal is connected in a circuit, the electrons move towards the positive terminal while at the same time electrons are fed into the other end of the metal from the negative terminal.

- » They are malleable and ductile. Unlike the fixed bonds in diamond, metallic bonds are not rigid, but they are still strong. If a force is applied to a metal, rows of ions can slide over one another. They reposition themselves and the strong bonds re-form as shown in Figure 3.35. 'Malleable' means that metals can be bent or hammered into different shapes. 'Ductile' means that the metals can be pulled out into thin wires.
- » They usually have high melting and boiling points due to the strong attraction between the positive metal ions and the mobile 'sea' of electrons.



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

- » 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.

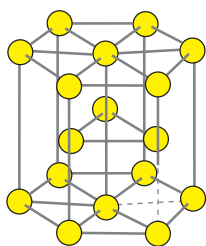


▲ **Figure 3.36** Arrangement of ions in the crystal lattice of a metal

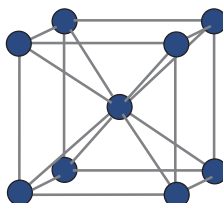
### Test yourself

- 13 Explain the terms:  
**a** malleable  
**b** ductile.
- 14 Explain why metals are able to conduct electricity.

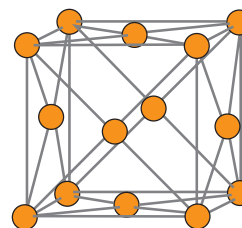
### ➔ Going further



Close-packed hexagonal structure (e.g. magnesium, density =  $1.74 \text{ g cm}^{-3}$ )



Body-centred cubic structure (e.g. iron, density =  $7.87 \text{ g cm}^{-3}$ )



Face-centred cubic structure (e.g. copper, density =  $8.92 \text{ g cm}^{-3}$ )

▲ **Figure 3.37** Relating different structures to the density of metals

#### Revision checklist

After studying Chapter 3 you should be able to:

- ✓ Explain how compounds form between metals and non-metals.
- ✓ Describe the properties of ionic compounds.
- ✓ Explain in terms of structure and bonding the properties of ionic compounds.

- ✓ Work out the formulae of ionic substances from the charges on the ions.

- ✓ Identify redox reactions as reactions involving gain and loss of oxygen.

- ✓ Identify redox reactions as reactions involving gain and loss of electrons.
- ✓ Identify redox reactions by changes in oxidation number.

- ✓ Define redox reactions as involving simultaneous oxidation and reduction.
- ✓ Identify oxidation as gain of oxygen and reduction as loss of oxygen.

- ✓ Define oxidation as loss of electrons and increase in oxidation number, and reduction as gain of electrons and reduction in oxidation number.
- ✓ Identify redox reactions by the colour changes involved when using acidified potassium manganate(VII) or potassium iodide.
- ✓ Define an oxidising agent as a substance that oxidises another substance and is itself reduced during a redox reaction.
- ✓ Define a reducing agent as a substance that reduces another substance and is itself oxidised during a redox reaction.

- ✓ Identify oxidation and reduction in redox reactions.

- ✓ Identify oxidising agents and reducing agents in redox reactions.

- ✓ Explain how simple compounds form between non-metals by single covalent bonds.

- ✓ Explain how simple compounds containing multiple covalent bonds form between non-metals.

- ✓ Describe and explain the properties of simple covalent compounds.

- ✓ Describe the giant covalent structures of graphite and diamond.

- ✓ Describe the giant covalent structure of silicon(IV) oxide,  $\text{SiO}_2$ .

- ✓ Relate the structures of graphite and diamond to their uses.

- ✓ Describe the similarity in properties between diamond and silicon(IV) oxide, related to their structures.

- ✓ Describe the arrangement of particles in a solid metal.

- ✓ Recognise that the particles in a metal are held together by metallic bonds.

- ✓ Describe metallic bonding.

- ✓ Use metallic bonding to explain some of the properties of metals.

## Exam-style questions

- 1 Define the following:
  - a positive (+) ion [2]
  - b negative (–) ion [2]
  - c ionic bond [2]
  - d electrostatic force of attraction. [2]
- 2 Using Table 3.1 on p. 35, write down the names of:
  - a three atoms which would form an ion with a charge of +2 [3]
  - b three atoms which would form an ion with a charge of +1 [3]
  - c three atoms/groups of atoms which would form an ion with a charge of –2. [3]
- 3 Sketch diagrams to show the bonding in each of the following compounds:
  - a calcium fluoride ( $\text{CaF}_2$ ) [4]
  - b oxygen ( $\text{O}_2$ ) [4]
  - c magnesium chloride ( $\text{MgCl}_2$ ) [4]
  - d tetrachloromethane ( $\text{CCl}_4$ ). [4]
- 4 Use the information given in Table 3.1 on p. 35 to deduce the formula for:
  - a silver oxide
  - b zinc chloride
  - c potassium sulfate
  - d calcium nitrate
  - e iron(II) nitrate
  - f copper(II) carbonate
  - g iron(III) hydroxide
  - h aluminium fluoride. [8]
- 5 Atoms of elements X, Y and Z have 16, 17 and 19 electrons, respectively. Atoms of argon have 18 electrons.
  - a Determine the formulae of the compounds formed by the combination of the atoms of the elements:
    - i X and Z [1]
    - ii Y and Z [1]
    - iii X with itself. [1]
  - b In each of the cases shown in a i–iii above, identify the type of chemical bond formed. [3]
  - c Give two properties you would expect each of the compounds formed in a ii and iii to have. [4]
- 6 Explain the following statements.
  - a Ammonia is a gas at room temperature. [3]
  - b The melting points of sodium chloride and iodine are very different. [8]
  - c Metals generally are good conductors of electricity. [3]
  - d Graphite acts as a lubricant but diamond does not. [3]