3.5 Metal cations and non-metal anions combine to form ionic compounds

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Metals form **cations** when they lose electrons to achieve a full, stable valency shell. Non-metals form **anions** when they gain electrons to achieve a full, stable valency shell. Polyatomic ions form when two or more atoms combine to form a charged ion. Positive cations are attracted to negative anions and form **ionic compounds**. The properties of ionic compounds reflect the ionic bonds that hold the ions together.

Forming ions

Electron shells are most stable when they are full – containing eight valence electrons. The behaviour of valence electrons can be explained by the atom seeking a stable state. The atom may gain or lose electrons in an attempt to gain a full outer electron shell. In certain cases, electrons are shared between atoms to achieve this balance.

The easiest way to achieve stability for atoms with only a few (1–3) valence electrons is to lose these electrons. For example, it is easier for an atom with two electrons in its outer shell to lose two electrons than to gain six electrons. In contrast, if the valence shell is almost full (seven electrons), it is more likely that atom will gain an electron to fill the gap in the shell. The number of positive protons does not change, even when the electrons move. Therefore, if an atom gains a negative electron, the overall charge of the atom becomes negative (more negative electrons than positive protons). If an atom loses electrons, then it becomes positively charged (more positive protons than negative electrons). In both these cases, the atom is then referred to as an **ion** – a charged atom.

Metals are usually found on the left-hand side of the periodic table. This means they have fewer than four electrons in their valency shell. Therefore, metals tend to lose electrons and become positively charged. Positively charged metals are called cations.

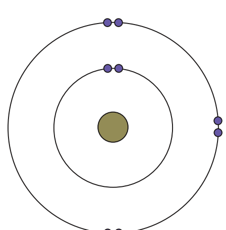

Figure 3.24 Oxygen tends to gain two electrons to fill its valence shell. Overall, there will be two more negative charges than positive charges (from the protons in the nucleus), so the ion is written as O^{2-}.

In contrast, most non-metals have almost full valency shells. This means they need to gain electrons to achieve a full valency shell of electrons. As a result, non-metals will become negatively charged. Negatively charged ions are called anions.

Positively charged cations are attracted to negatively charged anions. A cation with a 2+ charge is likely to combine (bond) with an anion of 2– charge or with two anions each with a charge of 1–. The positive charge is balanced by an equal negative charge. The bonds that are formed when ions interact are referred to as **ionic bonds**.

Properties of ionic compounds

Compounds that are held together by ionic bonds are called ionic compounds. As an ionic compound forms, the like-charged ions repel each other and the oppositely charged ions attract each other. After all the pushing and pulling, the ions settle into alternating positions, as shown in Figure 3.27, because this is the most stable arrangement. The particles are held together by strong electrostatic forces of attraction between the positively charged ions and the negatively charged ions. Because these forces bind the ions together, this is known as ionic bonding.

A lot of energy is required to move the ions out of their positions. This means that ionic compounds are hard to melt. At room

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temperature, they are in the form of hard, brittle crystals. The most commonly known example of an ionic compound is sodium chloride (table salt). Its melting point is 801°C. If you use a salt grinder at home, you will be aware of how hard and brittle salt crystals are.

Polyatomic ions

A number of ions are made up of more than one atom. These are termed **polyatomic ions**. Figure 3.28 shows some examples of polyatomic ions.

These clusters of atoms have a charge because the total number of protons does not equal the total number of electrons present. For example, in the hydroxide ion, which is made up of two atoms (one each of oxygen and hydrogen), there are nine protons and ten electrons. This means the two atoms that form the ion have an overall charge of 1–.

Calcium carbonate, the main constituent of marble, is an example of an ionic compound that contains a polyatomic ion. Calcium carbonate contains calcium ions (Ca^{2+}) and carbonate ions (CO_3^{2-}). These ions must be present in the ratio 1 : 1 so that the total positive charge equals the total negative charge. The formula of calcium carbonate is $CaCO_3$.

Ammonium carbonate is used in smelling salts. It contains ammonium ions (NH_4^+) and carbonate ions (CO_3^{2-}). In this case, the ions need to be present in the ratio 2 : 1. The formula of ammonium carbonate is $(NH_4)_2CO_3$.

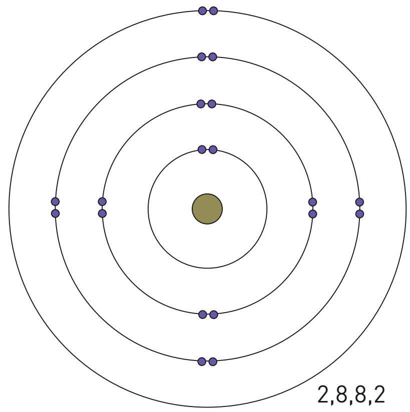

Figure 3.25 Calcium has two electrons in its valence shell, so it tends to lose them to achieve stability. The calcium ion formed is then written as Ca^{2+} to show that there are two extra protons compared with the number of electrons.

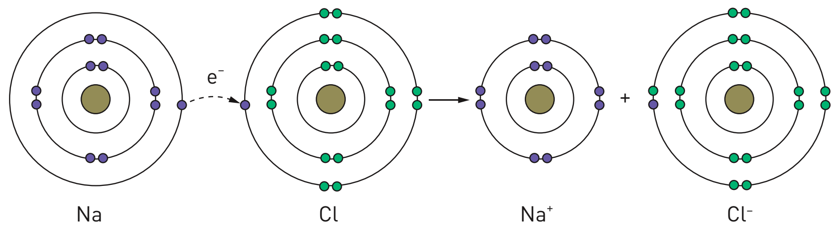

Figure 3.26 Sodium chloride is formed when sodium donates an electron to chlorine.

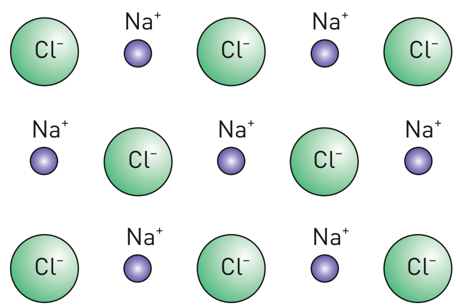

Figure 3.27 In an ionic compound, such as sodium chloride, the ions are arranged in alternating positions.

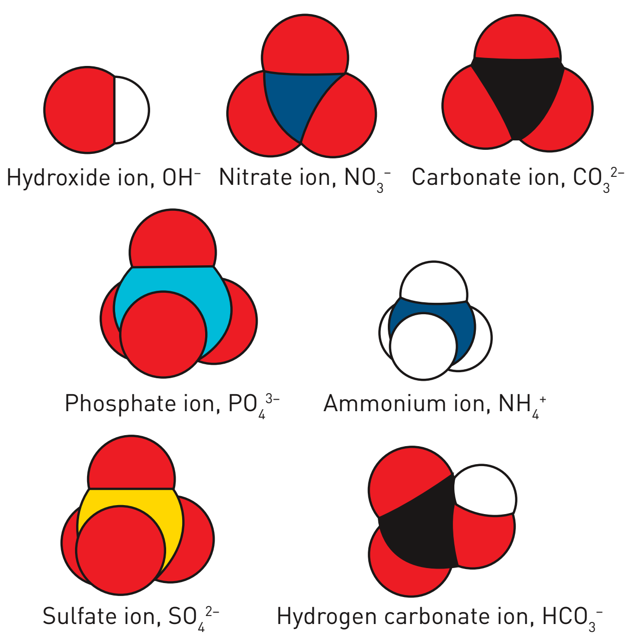

Figure 3.28 Some common polyatomic ions.

Check your learning 3.5

Remember and understand

1 Carefully examine the periodic table.
   a Which elements are likely to form positively charged ions?
   b Which elements are likely to form negatively charged ions?
   c What does this tell you about which elements will combine to form ionic compounds?

2 What kinds of particles are present in ionic compounds?

Apply and analyse

3 How does the group in which an element is found in the periodic table quickly identify one or more of its properties?

4 What is the maximum number of electrons that can be gained or lost by an atom? Why?

5 Use your knowledge of atomic structure and valence electrons to explain why many ionic compounds are made up of a metal and a non-metal.

3.6 Non-metals combine to form covalent compounds

Two non-metals need a full outer shell of electrons to remain stable. As a result, they merge their valency shells to share two electrons (one from each atom). This sharing of pairs of electrons between atoms is called a **covalent bond**. Covalent compounds form when two or more elements share pairs of electrons so that each has a full valency shell. The bonds between these atoms help explain the compound's properties.

Electrons and molecules

You have seen that when electrons are transferred from one atom to another, positive and negative ions are produced and ionic compounds are formed. However, two non-metals that complete their outer shells of electrons by gaining electrons can also bond together.

We can see this with the smallest, lightest atom: hydrogen.

Hydrogen molecules

An uncharged atom of hydrogen has just one electron in the first shell. If it could gain one more electron, this shell would contain its maximum number of electrons – two. If hydrogen was in contact with a reactive metal such as lithium, the hydrogen atom could gain that extra electron from a lithium atom. An ionic compound would form as a result. But what if only other uncharged hydrogen atoms were present? The only way each hydrogen atom can gain an extra electron is by sharing its electron with another.

As two uncharged hydrogen atoms come close together, the electrons are drawn into the region between the two nuclei. The atoms partially merge into one another, with the nuclei of both atoms now sharing the two electrons. The electrons travel in the spaces surrounding the nuclei of each atom. In effect, each atom now has a stable electron configuration because its outer shell is full.

The particle produced has two hydrogen atoms bonded strongly together and is called a molecule. A molecule is a particle produced when two or more atoms combine so that the atoms share electrons. A molecule has no overall charge because the total number of electrons and the total number of protons is the same.

The hydrogen molecule is given the formula H_2 because there are two hydrogen atoms present in the cluster.

The hydrogen molecule is an example of a molecule of an element. It is called a **diatomic molecule** because it is made up of two atoms. Other examples of diatomic molecules of non-metals are fluorine (F_2), chlorine (Cl_2), oxygen (O_2) and nitrogen (N_2).

In a molecule such as the hydrogen molecule, there is strong electrostatic attraction between each positively charged nucleus and the negatively charged electrons that they share.

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The electrons spend a considerable part of their time between the two nuclei. This electrostatic attraction is termed covalent bonding. The two shared electrons create a strong bond between the two atoms.

Molecular compounds

Like elements, compounds also form molecules. Water is an example of a **molecular compound**. Its formula is H_2O. You are now in a position to understand why it has this formula. To gain a more stable electron configuration, an:

> uncharged hydrogen atom, which has one valence electron, requires one more electron
> uncharged oxygen atom, which has six valence electrons, requires two more electrons.

A single hydrogen atom cannot supply the two electrons the oxygen atom needs, but two hydrogen atoms can. This is why there are two hydrogen atoms and just one oxygen atom in a water molecule. An oxygen atom now effectively has eight electrons in its valence shell and each hydrogen atom has two electrons. This is shown in Figure 3.30. Notice that each atom now has a full outer shell of electrons.

There are other ways of representing the structure of molecules, including with three-dimensional models. However, remember that in any representation, a single chemical bond holding the molecule together is actually a pair of negative electrons, shared between two atoms, attracted to the positive nuclei of both of these atoms.

Properties of molecular substances

Almost all molecular substances do not conduct electricity, even in the liquid state, because the molecules do not have free charged particles and so they cannot carry a current. There are no strong forces of attraction between molecules, so most molecular substances are liquids or gases at room temperature. It does not take much energy to separate the individual molecules and get them to move around.

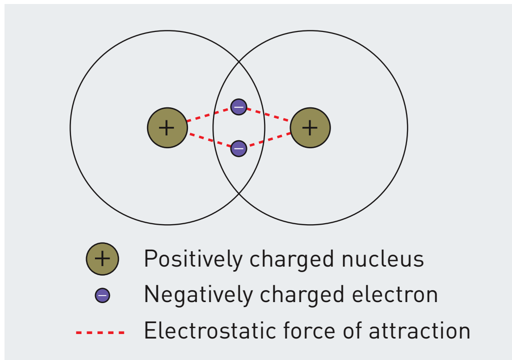

Figure 3.29 Covalent bonding within a hydrogen molecule.

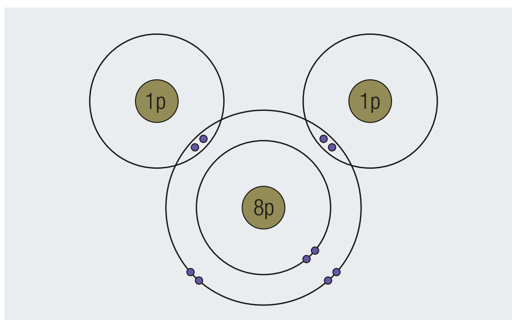

Figure 3.30 A shell diagram of a water molecule

Check your learning 3.6

Remember and understand

1 What is a diatomic molecule?
2 What types of atoms form covalent bonds?

Apply and analyse

3 Explain why molecular substances cannot conduct electricity.
4 In terms of the structure of the substance, why is it easier to turn liquid water into a gas than to break the bonds between the hydrogen and oxygen atoms?
5 a When chlorine atoms combine to form molecules, how many electrons need to be shared between the two chlorine atoms?
  b Would this be the same for two oxygen atoms combining to form a molecule? Explain your reasoning.

3.7 Metals form unique bonds

Metals form the largest collection of elements in the periodic table. They have many properties in common. All metals arrange their atoms into layers that can easily slide over each other. Some valence electrons are **delocalised**, and are able to freely move from one atom to another. This enables most metals to be good conductors. Metal alloys are mixtures of two or more metals that are stronger than pure metals. Smart alloys are metal mixtures that are able to retain a memory of their original shape.

Metallic structure

Many metals are malleable (can be bent into any shape). This property of metals is a result of the arrangement of atoms. Metal atoms arrange themselves into layers. When the metal is bent or hammered into shape, the atoms slide over one another (Figure 3.31).

Metals and conductivity

Remember that metals are found on the left-hand side of the periodic table. Metals do not have many electrons in their outer shells, and it does not take much energy for these outer electrons to move from one atom to another. This is the clue as to why metals are so good at conducting electricity.

A substance will conduct electricity if it contains charged particles that are free to move around the structure. In metals, these charged particles are electrons. Scientists refer to the outer-shell electrons as delocalised **electrons** because they are not 'stuck' in one locality. (Most electrons in metal atoms are not delocalised because they move about the nucleus of each metal atom in the inner electron shells.) Metals are good electrical conductors because the outer-shell electrons are free to move from nucleus to nucleus along the metal.

Table 3.3 gives the electrical conductivity of a number of elements at 25 °C.

All metals conduct electricity in the solid state, some better than others. They continue to conduct electricity when molten, but more weakly. The higher the temperature, the lower their electrical conductivity.

Only some metals are used for their electrical conductivity. For example, power lines have a core of steel and an outside layer of aluminium. Household wiring is usually copper, which is coated with a special kind of plastic. Metals like silver and gold are used in more specialised applications, such as in electronic devices.

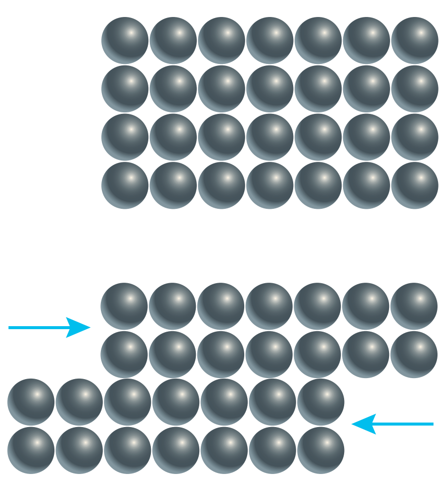

Figure 3.31 The arrangement of atoms in metals allows them to slide over each other when bent or hammered into shape.

Table 3.3 Electrical conductivities of some common elements at 25°C

ELEMENT	ELECTRICAL CONDUCTIVITY ($\times 10^6$ ohm^{-1} cm^{-1})
Aluminium	0.37
Silver	0.63
Carbon (graphite)	0.100
Copper	0.596
Gold	0.452
Iron	0.093
Lead	0.048
Magnesium	0.226
Sodium	0.210

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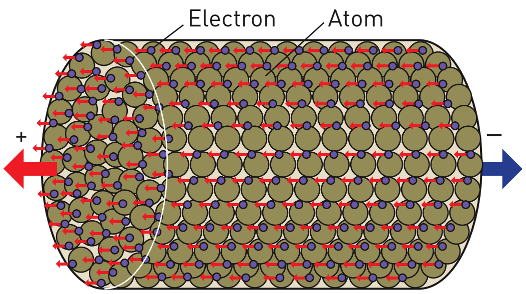

Figure 3.32 Delocalised electrons move about randomly in a metal, but move towards the positive terminal of the power source when connected into a circuit.

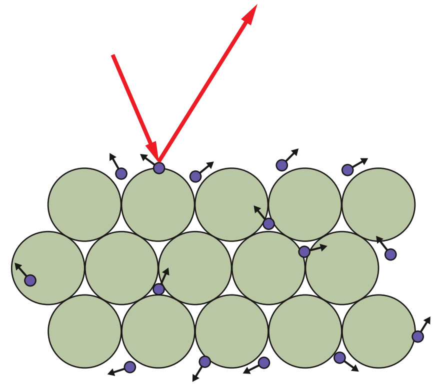

Figure 3.33 The delocalised electrons in the surface of a metal reflect light and cause it to be lustrous.

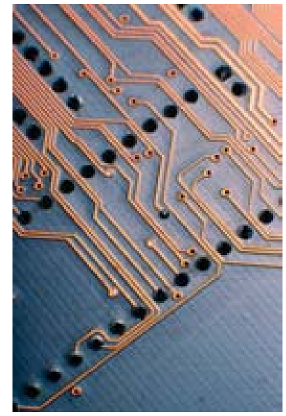

Figure 3.34 Gold bonding wire is used in integrated circuits.

Delocalised electrons are responsible for a pure metal being so lustrous (shiny). The delocalised electrons in its surface reflect light exceptionally well (Figure 3.33).

Metal alloys

An alloy is a mixture of two or more metals. Because the metal atoms are different sizes, the atoms are not arranged in the usual arrangement. This means there are no layers of atoms to slide over one another. As a result, alloys are harder than pure metals.

Soft metals such as copper, gold and aluminium are often mixed with other metals to make them hard enough for everyday use.

Brass (70% copper and 30% zinc) is used in electrical fittings and hinges.

Jewellery is often made of 18-carat gold (75% gold and 25% copper and other metals).

SMART ALLOYS

Some alloys have unique properties. When nitinol (a mixture of nickel and titanium) is cast into a particular shape and heated to 500°C, the atoms arrange themselves into a compact and regular pattern. This allows it to create a memory of this shape. If the alloy is bent out of shape, heat or an electric current can cause it return to its original shape. These metals are often called memory alloys (Figure 3.35).

An example of such memory wires are those used in orthodontic wires. The wires will constantly return to their original shape reducing the need to retighten or adjust the wire.

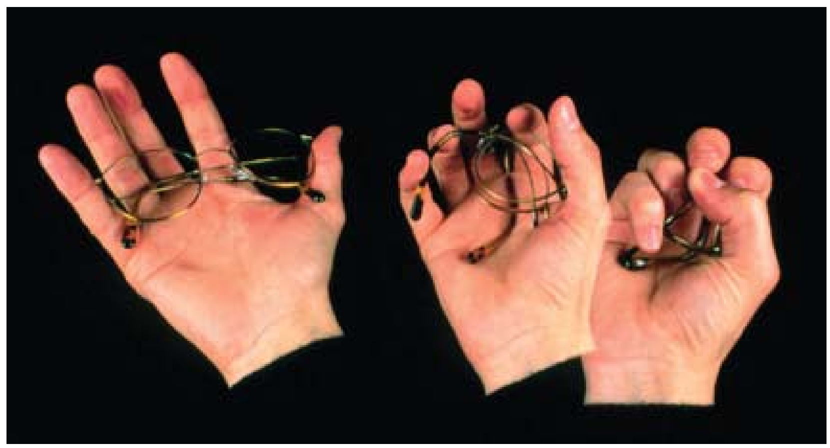

Figure 3.35 Memory wire is useful in eyeglass frames, allowing them to be bent out of shape without breaking.

Check your learning 3.7

Remember and understand

1 Describe the structure of a metal.
2 Identify the arrangement of atoms in a metal that enables each of the following properties.
  a Malleability
  b Conductivity
  c Shiny appearance
3 What is an alloy?

Apply and analyse

4 Compare the properties of an alloy with those of pure metal.
5 Memory alloys have been used to repair broken bones. Suggest why a memory alloy would be beneficial in this situation.