

4.1

Compounds, mixtures, and chemical change

Elements: a reminder

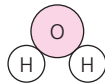
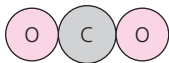
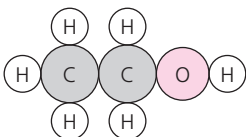
An **element** contains only one kind of atom. For example the element sodium contains only sodium atoms.

Compounds

A **compound** is made of atoms of different elements, bonded together.

The compound is described by a **formula**, made from the symbols of the atoms in it. (The plural of formula is **formulae**.)

There are millions of compounds. This table shows three common ones.

Name of compound	Elements in it	How the atoms are joined	Formula of compound
water	hydrogen and oxygen		H ₂ O
carbon dioxide	carbon and oxygen		CO ₂
ethanol	carbon, hydrogen, and oxygen		C ₂ H ₅ OH

Water has two hydrogen atoms joined or bonded to an oxygen atom. So its formula is H₂O. Note where the 2 is written. Now check the formulae for carbon dioxide and ethanol. Are they correct?

Compounds and mixtures: the difference

A **mixture** contains different substances that are *not* bonded together.

So you can usually separate the substances quite easily, using methods like those you met in Chapter 2. For example:



This is a **mixture** of iron powder and sulfur. You could separate them by dissolving the sulfur in methylbenzene (a solvent), and filtering the iron off.



But if you heat the end of a metal rod in a Bunsen burner, and push it into the mixture, the mixture starts to glow brightly. A **chemical change** is taking place.



The result is a black **compound** called iron(II) sulfide. It is made of iron and sulfur atoms bonded together. Its formula is FeS. It will not dissolve in methylbenzene.

The signs of a chemical change

When you heat a mixture of iron and sulfur, a chemical change takes place. The iron and sulfur atoms bond together to form a compound.

You can tell when a chemical change has taken place, by these three signs:

1 One or more new chemical substances are formed.

You can describe the change by a **word equation** like this:

iron + sulfur $\longrightarrow$ iron(II) sulfide

The + means *reacts with*, and the $\longrightarrow$ means *to form*.

The new substances usually look different from the starting substances. For example sulfur is yellow, but iron(II) sulfide is black.

2 Energy is taken in or given out, during the reaction.

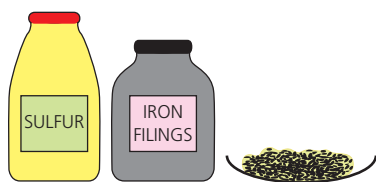
Energy was needed to start off the reaction between iron and sulfur, in the form of heat from the hot metal rod. But the reaction gave out heat once it began – the mixture glowed brightly.

3 The change is usually difficult to reverse.

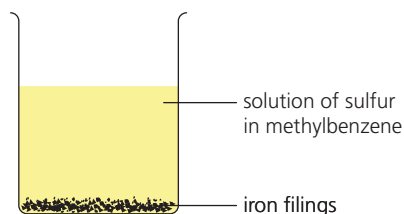
You would need to carry out several reactions to get the iron and sulfur back from iron sulfide. (But it can be done!)

A chemical change is usually called a chemical reaction.

It is different from physical change



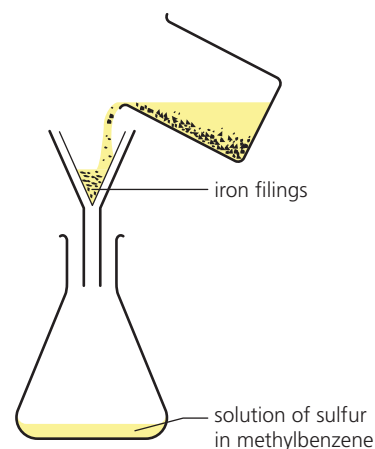
When you mix iron powder with sulfur, that is a **physical change**. No new substance has formed. If you then dissolve the sulfur ...



... in methylbenzene, that is also a physical change. The solvent could be removed again by distilling it. (Danger! It is highly flammable.)



▲ Burning gas, to fry eggs. Are chemical changes taking place?



Now separate the iron by filtering. That is a physical change. You can reverse it by putting the iron back into the filtrate again.

No new chemical substances are formed in these changes.

If no new chemical substance is formed, a change is a physical change.

Unlike chemical changes, a physical change is usually easy to reverse.

Q

- 1 Explain the difference between a *mixture* of iron and sulfur and the *compound* iron sulfide.
- 2 When you light a piece of magnesium ribbon, it burns with a dazzling white light. A white ash forms. What signs are there that a chemical change has taken place?

- 3 Is it a chemical change or a physical change? Give reasons.
 - a a glass bottle breaking
 - b butter and sugar being made into toffee
 - c cotton being woven to make sheets
 - d coal burning in air

4.2 Why do atoms form bonds?

The reaction between sodium and chlorine



Sodium and chlorine are both **elements**. When sodium is heated and placed in a jar of chlorine, it burns with a bright flame.



The result is a white solid that has to be scraped from the sides of the jar. It looks completely different from the sodium and chlorine.

So a chemical reaction has taken place. The white solid is **sodium chloride**. Atoms of sodium and chlorine have bonded (joined together) to form a compound. The word equation for the reaction is:



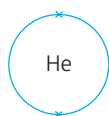
Why do atoms form bonds?

Like sodium and chlorine, the atoms of most elements form bonds.

Why? We get a clue by looking at the elements of Group 0, **the noble gases**. Their atoms *do not* form bonds.

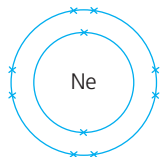
This is because the atoms have a very stable arrangement of electrons in the outer shell. This makes the noble gases **unreactive**.

helium atom:
full outer shell of 2
electrons – *stable*



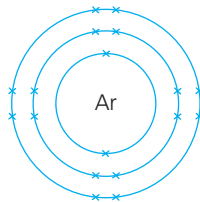
2

neon atom:
full outer shell of 8
electrons – *stable*



2 + 8

argon atom:
outer shell of 8
electrons – *stable*



2 + 8 + 8

And that gives us the answer to our question:

Atoms bond with each other in order to gain a stable arrangement of outer-shell electrons, like the atoms of Group 0.

In other words, they bond in order to gain 8 electrons in their outer shell (or 2, if they have only one shell).



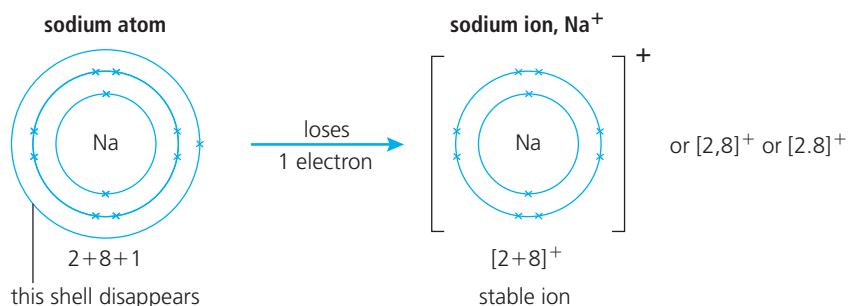
▲ Neon: the unreactive gas used in light tubes for advertising.



▲ Welding is often carried out in an atmosphere of argon, which will not react with hot metals (unlike oxygen).

How sodium atoms gain a stable outer shell

A sodium atom has just 1 electron in its outer shell. To obtain a stable outer shell of 8 electrons, it loses this electron to another atom. It becomes a **sodium ion**:



The sodium ion has 11 protons but only 10 electrons, so it has a charge of $1+$, as you can see from the panel on the right.

The symbol for sodium is Na, so the symbol for the sodium ion is Na^+ .

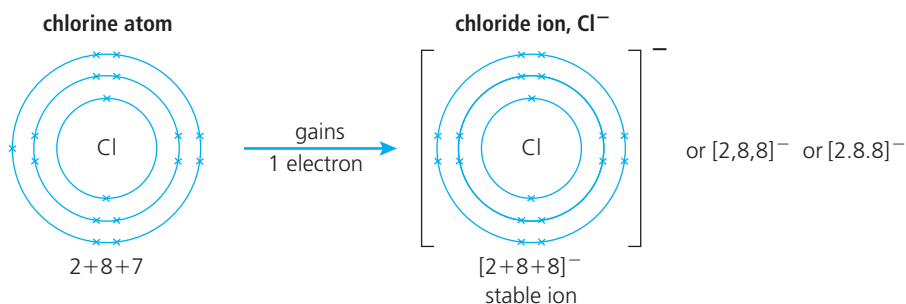
The $+$ means *1 positive charge*. Na^+ is a **positive ion**.

The charge on a sodium ion

charge on 11 protons	11+
charge on 10 electrons	10-
total charge	1+

How chlorine atoms gain a stable outer shell

A chlorine atom has 7 electrons in its outer shell. It can reach 8 electrons by accepting 1 electron from another atom. It becomes a chloride ion:



The chloride ion has a charge of $1-$, so it is a **negative ion**. Its symbol is Cl^- .

The charge on a chloride ion

charge on 17 protons	17+
charge on 18 electrons	18-
total charge	1-

Ions

An atom becomes an ion when it loses or gains electrons.

An ion is a charged particle. It is charged because it has an unequal number of protons and electrons.

Q

- 1 Why are the atoms of the Group 0 elements unreactive?
- 2 Explain why all other atoms are reactive.
- 3 Draw a diagram to show how this atom gains a stable outer shell of 8 electrons:
 - a a sodium atom
 - b a chlorine atom

- 4 Explain why
 - a a sodium ion has a charge of $1+$
 - b a chloride ion has a charge of $1-$
- 5 Explain what an *ion* is, in your own words.
- 6 Atoms of Group 0 elements do *not* form ions. Why not?

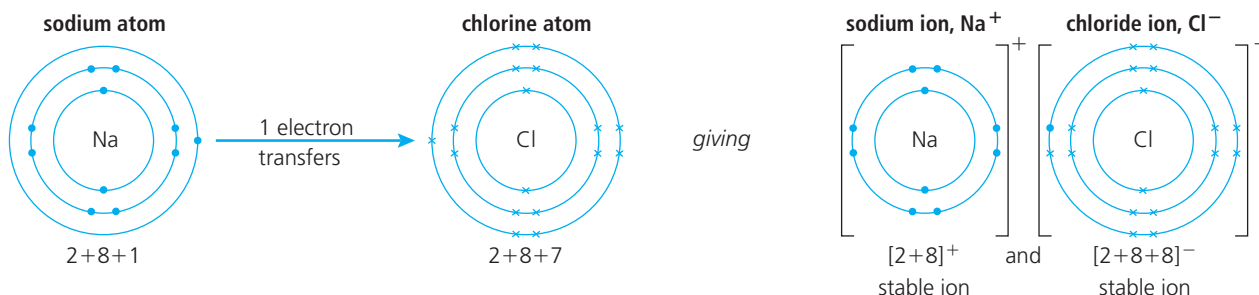
4.3 The ionic bond

How sodium and chlorine atoms bond together

As you saw on page 49, a sodium atom must lose one electron, and a chlorine atom must gain one, to obtain stable outer shells of 8 electrons.

So when a sodium atom and a chlorine atom react together, the sodium atom loses its electron *to the chlorine atom*, and two ions are formed.

Here, sodium electrons are shown as • and chlorine electrons as ×:

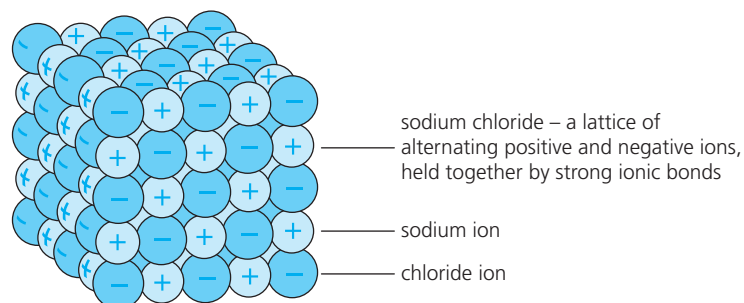


The two ions have opposite charges, so they attract each other. The force of attraction between them is strong. It is called an **ionic bond**.

The ionic bond is the bond that forms between ions of opposite charge.

How solid sodium chloride is formed

When sodium reacts with chlorine, billions of sodium and chloride ions form. But they do not stay in pairs. They form a regular pattern or **lattice** of alternating positive and negative ions, as shown below. The ions are held together by strong ionic bonds.



The lattice grows to form a giant 3-D structure. It is called 'giant' because it contains a very large number of ions. This giant structure is the compound **sodium chloride**, or **common salt**.

Since it is made of ions, sodium chloride is called an **ionic compound**. It contains one Na⁺ ion for each Cl⁻ ion, so its formula is **NaCl**.

The charges in the structure add up to zero:

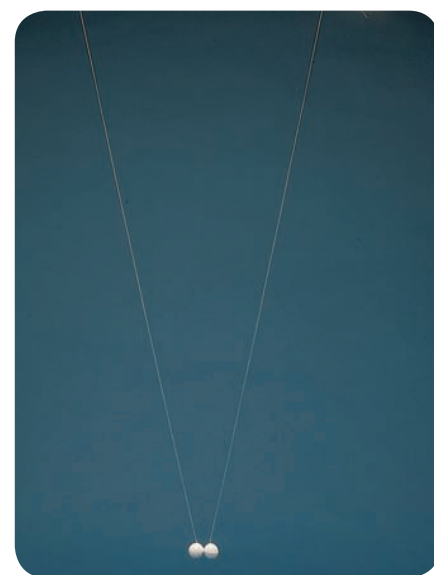
the charge on each sodium ion is	1+
the charge on each chloride ion is	1-
total charge	0

So the compound has no overall charge.

Bonding diagrams

To show the bonding clearly:

- use dots and crosses (o, •, and ×) for electrons from atoms of different elements
- write the symbol for the element in the centre of each atom.



▲ These polystyrene balls were given opposite charges. So they are attracted to each other, and cling together. The same happens with ions of opposite charge.

Other ionic compounds

Sodium is a metal. Chlorine is a non-metal. They react together to form an ionic compound. Other metals and non-metals follow the same pattern.

A metal reacts with a non-metal to form an ionic compound.

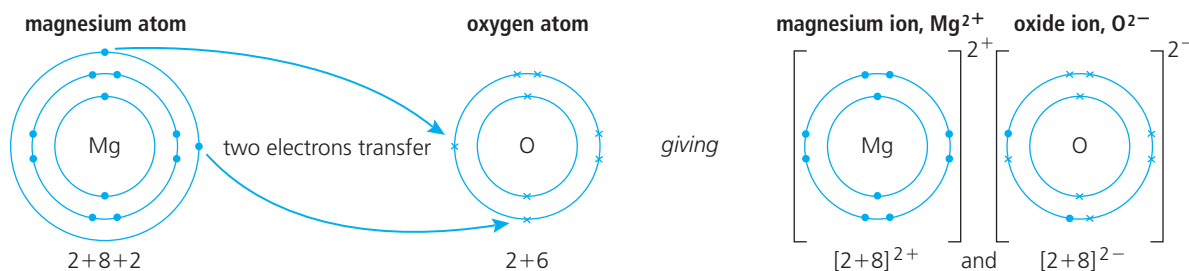
The metal atoms lose electrons. The non-metal atoms gain them.

The ions form a lattice. The compound has no overall charge.

Below are two more examples.

Magnesium oxide

A magnesium atom has 2 outer electrons and an oxygen atom has 6. When magnesium burns in oxygen, each magnesium atom loses its 2 outer electrons to an oxygen atom. Magnesium and oxide ions are formed:

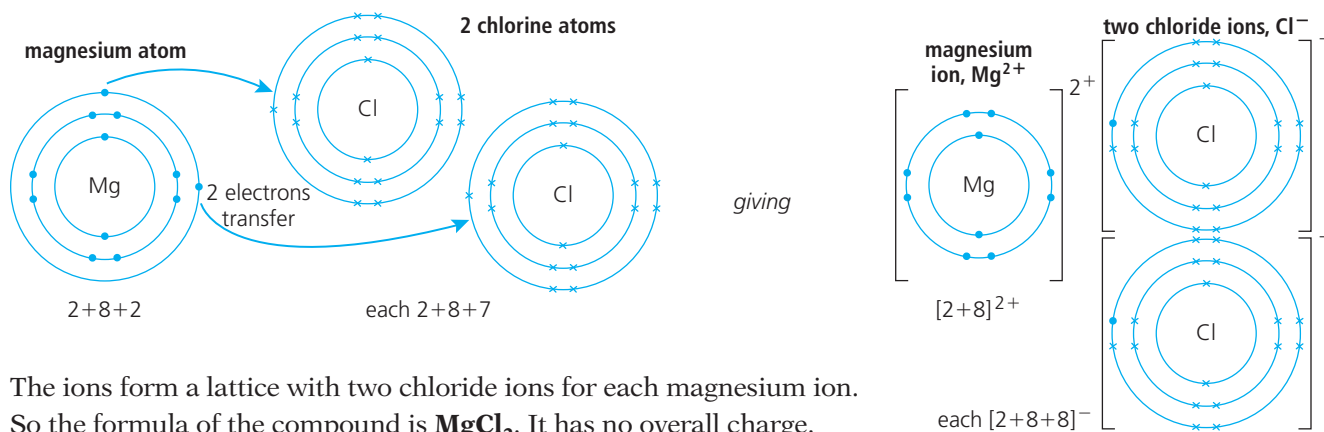


The ions attract each other because of their opposite charges. Like the sodium and chloride ions, they group to form a lattice.

The resulting compound is called **magnesium oxide**. It has one magnesium ion for each oxide ion, so its formula is **MgO**. It has no overall charge.

Magnesium chloride

When magnesium burns in chlorine, each magnesium atom reacts with *two* chlorine atoms, to form **magnesium chloride**. Each ion has 8 outer electrons:



The ions form a lattice with two chloride ions for each magnesium ion. So the formula of the compound is **MgCl₂**. It has no overall charge.

The charge on magnesium oxide

charge on a magnesium ion	2+
charge on an oxide ion	2-
total charge	0

Q

- 1 Draw a diagram to show what happens to the electrons, when a sodium atom reacts with a chlorine atom.
- 2 What is an *ionic bond*?
- 3 Describe in your own words the structure of solid sodium chloride, and explain why its formula is NaCl.
- 4 Explain why:
 - a a magnesium ion has a charge of 2+
 - b the ions in magnesium oxide stay together
 - c magnesium chloride has no overall charge
 - d the formula of magnesium chloride is MgCl₂.

4.4 More about ions

Ions of the first twenty elements

Not every element forms ions during reactions. In fact, out of the first twenty elements in the Periodic Table, only twelve easily form ions. These ions are given below, with their names.

Group I		II	H ⁺ hydrogen		III	IV	V	VI	VII	0
Li ⁺ lithium	Be ²⁺ beryllium							O ²⁻ oxide	F ⁻ fluoride	none
Na ⁺ sodium	Mg ²⁺ magnesium				Al ³⁺ aluminium			S ²⁻ sulfide	Cl ⁻ chloride	none
K ⁺ potassium	Ca ²⁺ calcium	transition elements								

Note that:

- Hydrogen and the metals lose electrons and form **positive ions**. The ions have the same names as the atoms.
- Non-metals form **negative ions**, with names ending in **-ide**.
- The elements in Groups IV and V do not usually form ions, because their atoms would have to gain or lose several electrons, and that takes too much energy.
- Group 0 elements do not form ions: their atoms already have stable outer shells, so do not need to gain or lose electrons.

The names and formulae of ionic compounds

The names To name an ionic compound, you just put the names of the ions together, with the positive one first:

Ions in compound	Name of compound
K ⁺ and F ⁻	potassium fluoride
Ca ²⁺ and S ²⁻	calcium sulfide

The formulae The formulae of ionic compounds can be worked out using these four steps. Look at the examples that follow.

- Write down the name of the ionic compound.
- Write down the symbols for its ions.
- The compound must have no overall charge, so balance the ions until the positive and negative charges add up to zero.
- Write down the formula without the charges.

Example 1

- Lithium fluoride.
- The ions are Li⁺ and F⁻.
- One Li⁺ is needed for every F⁻, to make the total charge zero.
- The formula is LiF.

Example 2

- Sodium sulfide.
- The ions are Na⁺ and S²⁻.
- Two Na⁺ ions are needed for every S²⁻ ion, to make the total charge zero: Na⁺ Na⁺ S²⁻.
- The formula is Na₂S. (What does the ₂ show?)



▲ Bath time. Bath salts contain ionic compounds such as magnesium sulfate (Epsom salts) and sodium hydrogen carbonate (baking soda). Plus scent!

Some metals form more than one type of ion

Look back at the Periodic Table on page 31. Look for the block of **transition elements**. These include many common metals, such as iron and copper.

Some transition elements form only one type of ion:

- silver forms only Ag^+ ions
- zinc forms only Zn^{2+} ions.

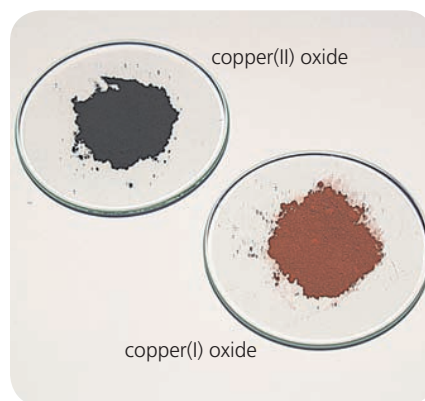
But most transition elements can form more than one type of ion.

For example, copper and iron can each form two:

Ion	Name	Example of compound
Cu^+	copper(I) ion	copper(I) oxide, Cu_2O
Cu^{2+}	copper(II) ion	copper(II) oxide, CuO
Fe^{2+}	iron(II) ion	iron(II) chloride, FeCl_2
Fe^{3+}	iron(III) ion	iron(III) chloride, FeCl_3

The (II) in the name tells you that the ion has a charge of $2+$.

What do the (I) and (III) show?



▲ The two oxides of copper.

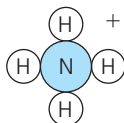
Compound ions

All the ions you met so far have been formed from single atoms. But ions can also be formed from a **group** of bonded atoms. These are called **compound ions**.

The most common ones are shown on the right. Remember, each is just one ion, even though it contains more than one atom.

The formulae for their compounds can be worked out as before. Some examples are shown below.

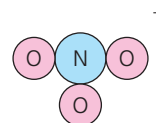
NH_4^+ , the ammonium ion



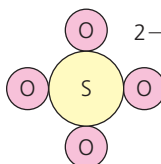
OH^- , the hydroxide ion



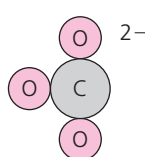
NO_3^- , the nitrate ion



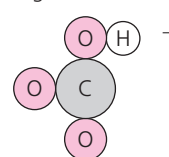
SO_4^{2-} , the sulfate ion



CO_3^{2-} , the carbonate ion



HCO_3^- , the hydrogen carbonate ion



Example 3

- 1 Sodium carbonate.
- 2 The ions are Na^+ and CO_3^{2-} .
- 3 Two Na^+ are needed to balance the charge on one CO_3^{2-} .
- 4 The formula is Na_2CO_3 .

Example 4

- 1 Calcium nitrate.
- 2 The ions are Ca^{2+} and NO_3^- .
- 3 Two NO_3^- are needed to balance the charge on one Ca^{2+} .
- 4 The formula is $\text{Ca}(\text{NO}_3)_2$. Note that brackets are put round the NO_3 , before the $_2$ is put in.

Q

1 Explain why a calcium ion has a charge of $2+$.

2 Why is the charge on an aluminium ion $3+$?

3 Write down the symbols for the ions in:

- a potassium chloride b calcium sulfide
c lithium sulfide d magnesium fluoride

4 Now work out the formula for each compound in 3.

5 Work out the formula for each compound:

- a copper(II) chloride b iron(III) oxide

6 Write a name for each compound:

CuCl , FeS , $\text{Mg}(\text{NO}_3)_2$, NH_4NO_3 , $\text{Ca}(\text{HCO}_3)_2$

7 Work out the formula for:

- a sodium sulfate b potassium hydroxide c silver nitrate

54

Chlorine

A chlorine atom needs a share in one more electron, to obtain a stable outer shell of eight electrons. So two chlorine atoms bond covalently like this:



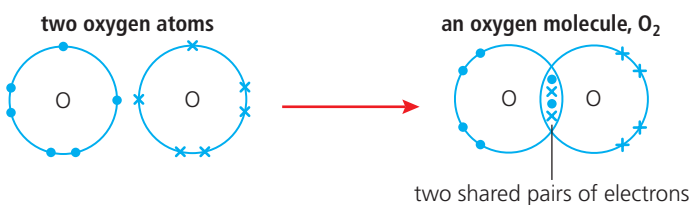
Since only one pair of electrons is shared, the bond between the atoms is called a **single covalent bond**, or just a **single bond**. You can show it in a short way by a single line, like this: Cl–Cl.



▲ A model of the chlorine molecule.

Oxygen

An oxygen atom has six outer electrons, so needs a share in *two* more. So two oxygen atoms share two electrons each, giving molecules with the formula O₂. Each atom now has a stable outer shell of eight electrons:



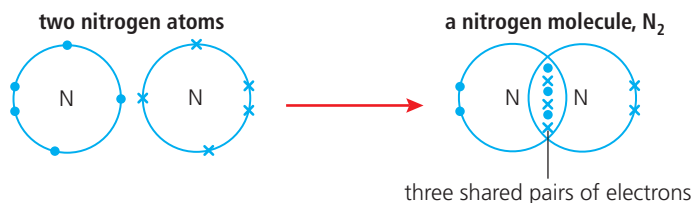
Since the oxygen atoms share two pairs of electrons, the bond between them is called a **double bond**. You can show it like this: O=O.



▲ A model of the oxygen molecule.

Nitrogen

A nitrogen atom has five outer electrons, so needs a share in *three* more. So two nitrogen atoms share three electrons each, giving molecules with the formula N₂. Each atom now has a stable outer shell of eight electrons:



Since the nitrogen atoms share three pairs of electrons, the bond between them is called a **triple bond**. You can show it like this: N≡N.



▲ A model of the nitrogen molecule.

Q

- 1 a Name the bond between atoms that share electrons.
b What holds the bonded atoms together?
- 2 What is a *molecule*?
- 3 Give five examples of molecular elements.
- 4 Draw a diagram to show the bonding in:
a hydrogen b chlorine
- 5 Now explain why the bond in a nitrogen molecule is called a *triple bond*.

4.6 Covalent compounds

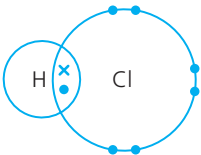
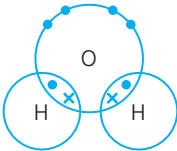
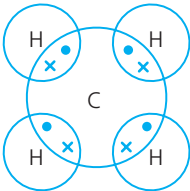
Covalent compounds

In the last unit you saw that many non-metal elements exist as molecules. A huge number of *compounds* also exist as molecules.

In a molecular compound, atoms of *different* elements share electrons. The compounds are called **covalent compounds**. Here are three examples.

Most are molecular ...

Most non-metal elements and their compounds exist as molecules.

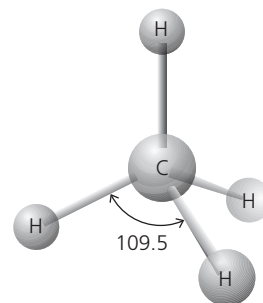
Covalent compound	Description	Model of the molecule
hydrogen chloride, HCl  a molecule of hydrogen chloride	The chlorine atom shares one electron with the hydrogen atom. Both now have a stable arrangement of electrons in their outer shells: 2 for hydrogen (like the helium atom) and 8 for chlorine (like the other noble gas atoms).	
water, H₂O  a molecule of water	The oxygen atom shares electrons with the two hydrogen atoms. All now have a stable arrangement of electrons in their outer shells: 2 for hydrogen and 8 for oxygen.	
methane, CH₄  a molecule of methane	The carbon atom shares electrons with four hydrogen atoms. All now have a stable arrangement of electrons in their outer shells: 2 for hydrogen and 8 for carbon.	

The shapes of the molecules

Look at the models of the methane molecule, above and on the right. The molecule is **tetrahedral** in shape, because the four pairs of electrons around carbon repel each other, and move as far apart as possible.

Now look at the model of the water molecule above. The hydrogen atoms are closer together than in methane. This is because the two non-bonding pairs of atoms repel more strongly than the bonding pairs. So they push these closer together.

The angle between the hydrogen atoms in water is 104.5°.

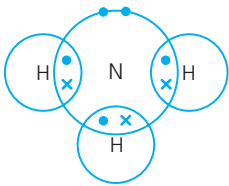
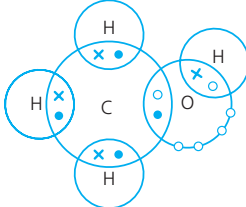
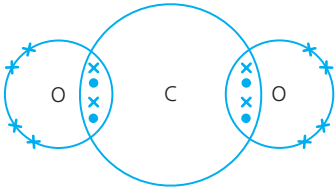
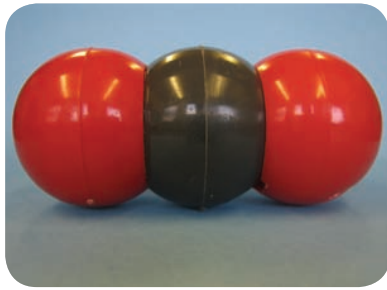
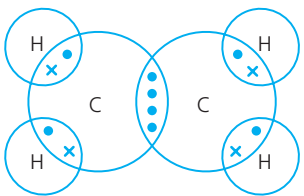


▲ The methane molecule: the same angle between all the H atoms.

More examples of covalent compounds

This table shows three more examples of covalent compounds. Each time:

- the atoms share electrons, to gain stable outer shells
- repulsion between pairs of electrons dictates the shape of the molecule.

Covalent compound	Description	Model of the molecule
ammonia, NH_3  a molecule of ammonia	Each nitrogen atom shares electrons with three hydrogen atoms. So all three atoms now have a stable arrangement of electrons in their outer shells: 2 for hydrogen and 8 for nitrogen. The molecule is shaped like a pyramid.	
methanol, CH_3OH  a molecule of methanol	The carbon atom shares electrons with three hydrogen atoms and one oxygen atom. Look at the shape of the molecule: a little like methane, but changed by the presence of the oxygen atom.	
carbon dioxide, CO_2  a molecule of carbon dioxide	The carbon atom shares all four of its electrons: two with each oxygen atom. So all three atoms gain stable shells. The two sets of bonding electrons repel each other. They move as far apart as they can, giving a linear molecule. All the bonds are double bonds, so we can show the molecule like this: $\text{O}=\text{C}=\text{O}$.	
ethene, C_2H_4  a molecule of ethene	Look how each carbon atom shares its four electrons this time. It shares two with two hydrogen atoms, and two with another carbon atom, giving a carbon-carbon double bond. So the molecule is usually drawn like this: $\begin{array}{c} \text{H} & & \text{H} \\ & \diagdown & / \\ & \text{C} = \text{C} \\ & / & \diagdown \\ \text{H} & & \text{H} \end{array}$	

Q

- a** What is a *covalent compound*?

b Give five examples, with their formulae.
- Draw a diagram to show the bonding in a molecule of:

a methane **b** water
- How do the atoms gain stable outer shells, in ammonia?
- Draw a diagram to show the bonding in carbon dioxide.
- Why is the carbon dioxide molecule straight, and not bent like the water molecule?

4.7

Comparing ionic and covalent compounds

Remember

Metals and non-metals react together to form **ionic compounds**.

Non-metals react together to form **covalent compounds**.

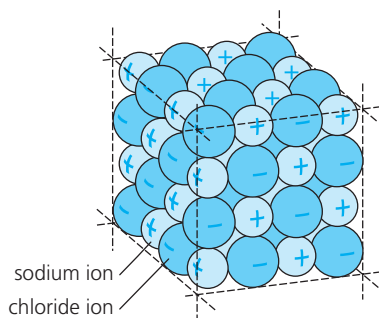
The covalent compounds you have met so far exist as **molecules**.

Comparing the structures of the solids

In Chapter 1, you met the idea that solids are a **regular lattice** of particles.

In ionic compounds, these particles are **ions**. In the covalent compounds you have met so far, they are **molecules**. Let's compare their lattices.

A solid ionic compound Sodium chloride is a typical ionic compound:



In sodium chloride, the ions are held in a regular lattice like this. They are held by strong ionic bonds.

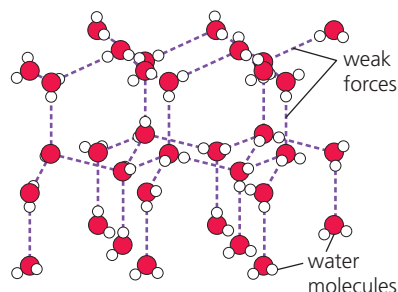


The lattice grows in all directions, giving a crystal of sodium chloride. This one is magnified 35 times.



The crystals look white and shiny. We add them to food, as **salt**, to bring out its taste.

A solid molecular covalent compound Water is a molecular covalent compound. When you cool it below 0°C it becomes a solid: ice.



In ice, the water molecules are held in a regular lattice like this. But the forces between them are weak.



The lattice grows in all directions, giving a crystal of ice. These grew in an ice-tray in a freezer.



We use ice to keep drinks cool, and food fresh. (The reactions that cause food to decay are slower in the cold.)

So both types of compounds have a regular lattice structure in the solid state, and form crystals. But they differ in two key ways:

- In ionic solids the particles (ions) are charged, and the forces between them are strong.
- In molecular covalent solids the particles (molecules) are not charged, and the forces between them are weak.

These differences lead to very different properties, as you will see next.

About crystals

- A regular arrangement of particles in a lattice always leads to crystals.
- The particles can be atoms, ions, or molecules.

The properties of ionic compounds

1 Ionic compounds have high melting and boiling points.

For example:

Compound	Melting point/°C	Boiling point/°C
sodium chloride, NaCl	801	1413
magnesium oxide, MgO	2852	3600

This is because the ionic bonds are very strong. It takes a lot of heat energy to break up the lattice. So ionic compounds are solid at room temperature.

Note that magnesium oxide has a far higher melting and boiling point than sodium chloride does. This is because its ions have double the charge (Mg^{2+} and O^{2-} compared with Na^+ and Cl^-), so its ionic bonds are stronger.

2 Ionic compounds are usually soluble in water.

The water molecules are able to separate the ions from each other. The ions then move apart, surrounded by water molecules.

3 Ionic compounds conduct electricity, when melted or dissolved in water.

A solid ionic compound will not conduct electricity. But when it melts, or dissolves in water, the ions become free to move. Since they are charged, they can then conduct electricity.



▲ Magnesium oxide is used to line furnaces in steel works, because of its high melting point, 2852 °C. (By contrast, iron melts at 1538 °C.)

The properties of covalent compounds

1 Molecular covalent compounds have low melting and boiling points.

For example:

Compound	Melting point/°C	Boiling point/°C
carbon monoxide, CO	-199	-191
hexane, C_6H_{14}	-95	69

This is because the attraction between the molecules is low. So it does not take much energy to break up the lattice and separate them from each other. That explains why many molecular compounds are liquids or gases at room temperature – and why many of the liquids are **volatile** (evaporate easily).

2 Covalent compounds tend to be insoluble in water.

But they do dissolve in some solvents, for example tetrachloromethane.

3 Covalent compounds do not conduct electricity.

There are no charged particles, so they cannot conduct.



▲ The covalent compound carbon monoxide is formed when petrol burns in the limited supply of air in a car engine. And it is poisonous.

Q

- 1 The particles in solids usually form a *regular lattice*. Explain what that means, in your own words.
- 2 Which type of particles make up the lattice, in:
a ionic compounds? b molecular compounds?
- 3 Solid sodium chloride will not conduct electricity, but a solution of sodium chloride will conduct. Explain this.

- 4 A compound melts at 20 °C.
a What kind of structure do you think it has? Why do you think so?
b Will it conduct electricity at 25 °C? Give a reason.
- 5 Describe the arrangement of the molecules in ice. How will the arrangement change as the ice warms up?

4.8 Giant covalent structures

Not all covalent solids are molecular

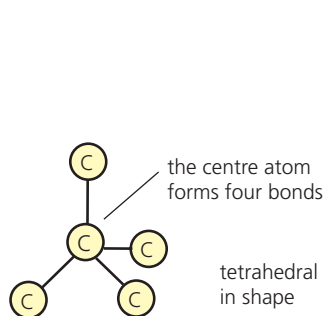
In all the solids in this table, the atoms are held together by covalent bonds. But compare their melting points. What do you notice?

Substance	Melting point/°C
ice	0
phosphorus	44
sulfur	115
silicon dioxide (silica)	1710
carbon (as diamond)	3550

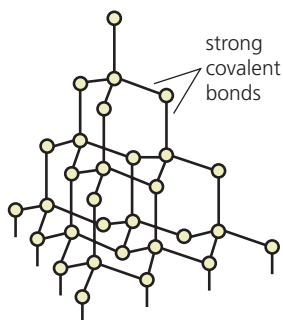
The first three substances are molecular solids. Their molecules are held in a lattice by weak forces – so the solids melt easily, at low temperatures. But diamond and silica are different. Their melting points show that *they* are not molecular solids with weak lattices. In fact they exist as **giant covalent structures**, or **macromolecules**.

Diamond – a giant covalent structure

Diamond is made of carbon atoms, held in a strong lattice:



A carbon atom forms covalent bonds to *four* others, as shown above. Each outer atom then bonds to three more, and so on.



Eventually billions of carbon atoms are bonded together, in a giant covalent structure. This shows just a very tiny part of it.

Diamond has these properties:

- 1 It is very hard, because each atom is held in place by four strong covalent bonds. In fact it is the hardest substance on Earth.
- 2 For the same reason it has a very high melting point, 3550°C.
- 3 It can't conduct electricity because there are no ions or free electrons to carry the charge.

Silica is similar to diamond

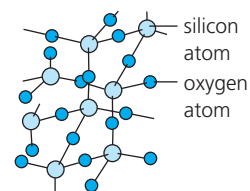
Silica, SiO_2 , occurs naturally as **quartz**, the main mineral in **sand**. Like diamond, it forms a giant covalent structure, as shown on the right. Each silicon atom bonds covalently to four oxygen atoms. And each oxygen atom bonds covalently to two silicon atoms. The result is a very hard substance with a melting point of 1710°C.



▲ Diamond: so hard that it is used to edge wheels for cutting stone.



The result is a single crystal of diamond. This one has been cut, shaped, and polished, to make it sparkle.

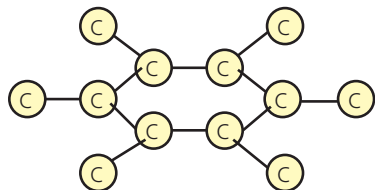


▲ Silicon dioxide is made up of oxygen atoms ● and silicon atoms ●. Billions of them bond together like this, to give a giant structure.

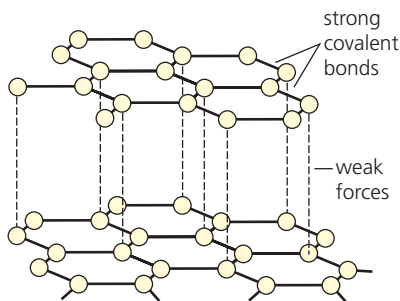
Graphite – a very different giant structure

Like diamond, graphite is made only of carbon atoms. So diamond and graphite are **allotropes** of carbon – two forms of the same element.

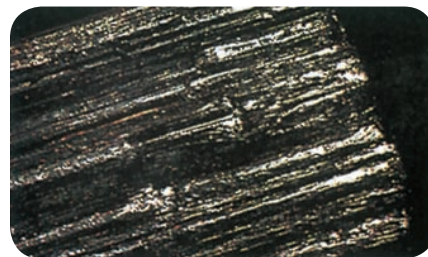
Diamond is the hardest solid on Earth. But graphite is one of the softest! This difference is a result of their very different structures:



In graphite, each carbon atom forms covalent bonds to *three* others. This gives rings of *six* atoms.



The rings form flat sheets that lie on top of each other, held together by weak forces.



Under a microscope, you can see the layered structure of graphite quite clearly.

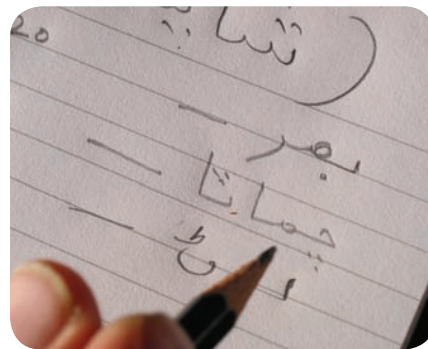
Graphite has these properties:

- 1 Unlike diamond, it is soft and slippery. That is because the sheets can slide over each other easily.
- 2 Unlike diamond, it is a good conductor of electricity. That is because each carbon atom has four outer electrons, but forms only three bonds. So the fourth electron is free to move through the graphite, carrying charge.

Making use of these giant structures

Different properties lead to different uses, as this table shows.

Substance	Properties	Uses
diamond	hardest known substance	in tools for drilling and cutting
	does not conduct	
graphite	sparkles when cut	for jewellery
	soft and slippery	as a lubricant for engines and locks
	soft and dark in colour	for pencil 'lead' (mixed with clay)
silica	conducts electricity	for electrodes, and connecting brushes in generators
	hard, can scratch things	in sandpaper
	hard, lets light through	for making glass and lenses
	high melting point	in bricks for lining furnaces



▲ Pencil 'lead' is a mixture of graphite and clay.

Q

- 1 The covalent compound ethanol melts at -114°C . Is it a molecular compound, or a giant structure? Explain.
- 2 Diamond and graphite are *allotropes* of carbon. What does that mean?
- 3 Why is diamond so hard?
- 4 Why do diamond and graphite have such very different properties? Draw diagrams to help you explain.
- 5 a Explain why silica has a high melting point.
b See if you can suggest a reason why its melting point is lower than diamond's.

4.9 The bonding in metals

Clues from melting points

Compare these melting points:

Structure	Examples	Melting point / °C
molecular	carbon dioxide water	−56 0
giant ionic	sodium chloride magnesium oxide	801 2852
giant covalent	diamond silica	3550 1610
metal	iron copper	1535 1083

The table shows clearly that:

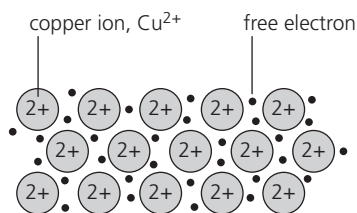
- **molecular substances have low melting points.** That is because the forces between molecules in the lattice are weak.
- **giant structures such as sodium chloride and diamond have much higher melting points.** That is because the bonds between ions or atoms within giant structures are very strong.

Now look at the metals. They too have high melting points – much higher than for carbon dioxide or water. This gives us a clue that they too might be giant structures. And so they are, as you'll see below.

The structure of metals

In metals, the atoms are packed tightly together in a regular lattice. The tight packing allows outer electrons to separate from their atoms. The result is a lattice of ions in a 'sea' of electrons that are free to move.

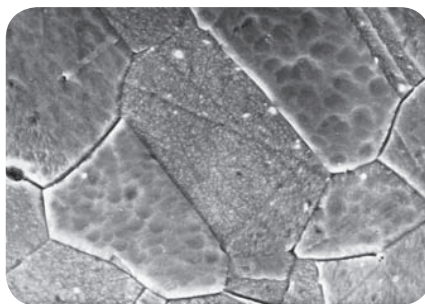
Look at copper:



The copper ions are held together by their attraction to the free electrons between them. The strong forces of attraction are called **metallic bonds**.

The metallic bond is the attraction between metal ions and free electrons.

It is the same with all metals. The ions sit in a lattice, held together by their strong attraction to the free electrons. And because the ions are in a regular pattern, metals are crystalline.



The regular arrangement of ions results in **crystals** of copper. This shows the crystals in a piece of copper, magnified 1000 times. (They are all at different angles.)



▲ Equipment for measuring melting points in the school lab. It can heat substances up to 300 °C – so no good for sodium chloride!



The copper crystals are called **grains**. A lump of copper like this one consists of millions of grains joined together. You need a microscope to see them.

Delocalised electrons

The electrons that move freely in the metal lattice are not tied to any one ion. So they are called **delocalised**.

Explaining some key properties of metals

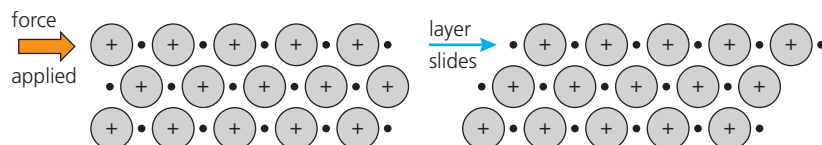
In Unit 3.5 you read about the properties of metals. We can now explain some of those properties. Look at these examples.

1 Metals usually have high melting points.

That is because it takes a lot of heat energy to break up the lattice, with its strong metallic bonds. Copper melts at 1083°C , and nickel at 1455°C . (But there are exceptions. Sodium melts at only 98°C , for example. And mercury melts at -39°C , so it is a liquid at room temperature.)

2 Metals are malleable and ductile.

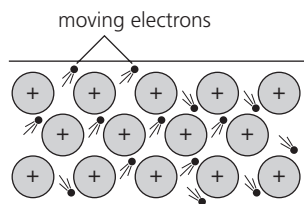
Malleable means they can be bent and pressed into shape. *Ductile* means they can be drawn out into wires. This is because the layers can slide over each other. This diagram represents any metal lattice:



The layers can slide without breaking the metallic bond, because the electrons are free to move too.

3 Metals are good conductors of heat.

That is because the free electrons take in heat energy, which makes them move faster. They quickly transfer the heat through the metal structure:



4 Metals are good conductors of electricity.

That is because the free electrons can move through the lattice carrying charge, when a voltage is applied across the metal.

Silver is the best conductor of all the metals.

Copper is next – but it is used much more than silver because it is cheaper.



▲ Metals: malleable, ductile, and sometimes very glamorous – like this silver bracelet.



▲ What uses of metals can you see in this scene?

Q

- 1 Describe in your own words the structure of a metal.
- 2 What is a *metallic bond*?
- 3 What does *malleable* mean?
- 4 Explain why metals can be drawn out into wires without breaking.
- 5 a Explain why metals can conduct electricity.
b Would you expect molten metals to conduct? Give a reason.

- 6 Because metals are malleable, we use some of them to make saucepans. Give two other examples of uses of metals that depend on:
a their malleability b their ductility
c their ability to conduct electricity
- 7 Mercury forms ions with a charge of $2+$. It goes solid (freezes) at -39°C . Try drawing a diagram to show the structure of solid mercury.

Checkup on Chapter 4

Revision checklist

Core curriculum

Make sure you can ...

- ☐ explain the difference between:
 - an *element* and a *compound*
 - a *compound* and a *mixture*
- ☐ say what the signs of a chemical change are
- ☐ explain why:
 - atoms of Group 0 elements do not form bonds
 - atoms of other elements do form bonds
- ☐ explain the difference between an *ionic bond* and a *covalent bond*
- ☐ draw a diagram to show how an ionic bond forms between atoms of sodium and chlorine
- ☐ explain what a *molecule* is
- ☐ say that non-metal atoms form covalent bonds with each other (except for the noble gas atoms)
- ☐ draw diagrams to show the covalent bonding in:

<i>hydrogen</i>	<i>chlorine</i>	<i>water</i>
<i>methane</i>	<i>hydrogen chloride</i>	
- ☐ give three ways in which ionic and molecular compounds differ in their properties, and explain these differences
- ☐ describe the giant covalent structures of graphite and diamond, and sketch them
- ☐ explain how their structures lead to different uses for diamond and graphite, with examples

Extended curriculum

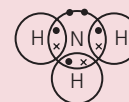
Make sure you can also ...

- ☐ show how ionic bonds form between atoms of other metals and non-metals
- ☐ describe the lattice structure of ionic compounds
- ☐ work out the formulae of ionic compounds, from the charges on the ions
- ☐ draw diagrams to show the covalent bonding in nitrogen, oxygen, ammonia, methanol, carbon dioxide, and ethene
- ☐ describe metallic bonding, and draw a sketch for it
- ☐ explain how the structure and bonding in metals enables them to be malleable, ductile, and good conductors of heat and electricity
- ☐ describe the structure of silicon dioxide
- ☐ explain why silicon dioxide and diamond have similar properties
- ☐ give examples of uses for silicon dioxide

Questions

Core curriculum

- 1 This question is about the ionic bond formed between the metal lithium (proton number 3) and the non-metal fluorine (proton number 9).
 - a How many electrons does a lithium atom have? Draw a diagram to show its electron structure.
 - b How does a metal atom obtain a stable outer shell of electrons?
 - c Draw the structure of a lithium ion, and write the symbol for it, showing its charge.
 - d How many electrons does a fluorine atom have? Draw a diagram to show its electron structure.
 - e How does a non-metal atom become an ion?
 - f Draw the structure of a fluoride ion, and write a symbol for it, showing its charge.
 - g Draw a diagram to show what happens when a lithium atom reacts with a fluorine atom.
 - h Write a word equation for the reaction between lithium and fluorine.



- 2 This diagram represents a molecule of a certain gas.
 - a Name the gas, and give its formula.
 - b What do the symbols • and × represent?
 - c Which type of bonding holds the atoms together?
 - d Name another compound with this type of bonding.
- 3 Hydrogen bromide is a compound of the two elements hydrogen and bromine. It melts at -87°C and boils at -67°C . It has the same type of bonding as hydrogen chloride.
 - a Is hydrogen bromide a solid, a liquid, or a gas at room temperature (20°C)?
 - b Is hydrogen bromide molecular, or does it have a giant structure? What is your evidence?
 - c i Which type of bond is formed between the hydrogen and bromine atoms, in hydrogen bromide?
 - ii Draw a diagram of the bonding between the atoms, showing only the outer electrons.
 - d Write a formula for hydrogen bromide.
 - e i Name two other compounds with bonding similar to that in hydrogen bromide.
 - ii Write formulae for these two compounds.

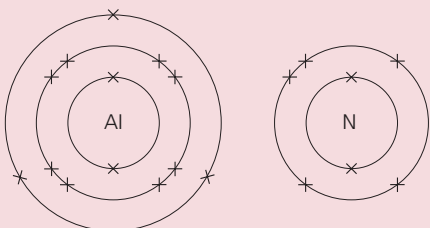
- 4 These are some properties of substances A to G.

Substance	Melting point / °C	Electrical conductivity		Solubility in water
		solid	liquid	
A	-112	poor	poor	insoluble
B	680	poor	good	soluble
C	-70	poor	poor	insoluble
D	1495	good	good	insoluble
E	610	poor	good	soluble
F	1610	poor	poor	insoluble
G	660	good	good	insoluble

- Which of the seven substances are metals? Give reasons for your choice.
- Which of the substances are ionic compounds? Give reasons for your choice.
- Two of the substances have very low melting points, compared with the rest. Explain why these could *not* be ionic compounds.
- Two of the substances are molecular. Which two are they?
- Which substance is a giant covalent structure?
 - What other name is used to describe this type of structure? (Hint: starts with *m*.)
- Name the type of bonding found in:
 - B
 - C
 - E
 - F

Extended curriculum

- 5 Aluminium and nitrogen react to form an ionic compound called aluminium nitride. These show the electron arrangement for the two elements:

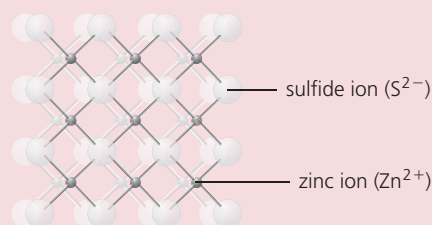


- Answer these questions for an aluminium atom.
 - Does it gain or lose electrons, to form an ion?
 - How many electrons are transferred?
 - Is the ion formed positive, or negative?
 - What charge does the ion have?
- Now repeat **a**, but for a nitrogen atom.
- Give the electron distribution for the ions formed by the two atoms. (2 + ...)
 - What do you notice about these distributions? Explain it.
- Name another non-metal that will form an ionic compound with aluminium, in the same way as nitrogen does.

- 6 Silicon lies directly below carbon in Group IV of the Periodic Table. Here is some data for silicon, carbon (in the form of diamond), and their oxides.

Substance	Symbol or formula	Melting point / °C	Boiling point / °C
carbon	C	3730	4530
silicon	Si	1410	2400
carbon dioxide	CO ₂	(turns to gas at -78°C)	
silicon dioxide	SiO ₂	1610	2230

- In which state are the two *elements* at room temperature (20 °C)?
 - Which type of structure does carbon (diamond) have: giant covalent, or molecular?
 - Which type of structure would you expect to find in silicon? Give reasons.
 - In which state are the two oxides, at room temperature?
 - Which type of structure has carbon dioxide?
 - Does silicon dioxide have the same structure as carbon dioxide? What is your evidence?
- 7 The compound zinc sulfide has a structure like this:



- Which does the diagram represent: a giant structure, or a molecular structure?
 - Which type of bonding does zinc sulfide have?
 - Look carefully at the structure. How many:
 - sulfur ions are joined to each zinc ion?
 - zinc ions are joined to each sulfur ion?
 - From **c**, deduce the formula of zinc sulfide.
 - Is this formula consistent with the charges on the two ions? Explain your answer.
 - Name another metal and non-metal that will form a compound with a similar formula.
- 8 The properties of metals can be explained by the structure and bonding within the metal lattice.
- Describe the bonding in metals.
 - Use the bonding to explain why metals:
 - are good conductors of electricity
 - are malleable and flexible