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:

\[
\text{sodium} + \text{chlorine} \rightarrow \text{sodium chloride}
\]

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

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).
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:

![Diagram of sodium atom losing an electron to form 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.

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:

![Diagram of chlorine atom gaining an electron to form a chloride ion.]

The chloride ion has a charge of 1−, so it is a negative ion. Its symbol is Cl⁻.

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.

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 sodium atom
   - 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.
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:

\[
\text{magnesium atom} \quad \text{oxygen atom} \\
\begin{array}{c}
\text{Mg} \\
\text{2+8+2}
\end{array} \quad \begin{array}{c}
\text{O} \\
\text{2+6}
\end{array} \quad \text{giving} \quad \begin{array}{c}
\text{magnesium ion, } \text{Mg}^{2+} \\
\text{[2+8]^{2+}}
\end{array} \quad \begin{array}{c}
\text{oxide ion, } \text{O}^{2-} \\
\text{[2+8]^{2-}}
\end{array}
\]

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 \( \text{MgO} \). 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:

\[
\text{magnesium atom} \quad \text{2 chlorine atoms} \\
\begin{array}{c}
\text{Mg} \\
\text{2+8+2}
\end{array} \quad \begin{array}{c}
\text{Cl} \\
\text{2+8+7}
\end{array} \quad \text{giving} \quad \begin{array}{c}
\text{magnesium ion, } \text{Mg}^{2+} \\
\text{[2+8]^{2+}}
\end{array} \quad \begin{array}{c}
\text{two chloride ions, } \text{Cl}^{-} \\
\text{[2+8+8]^{-}}
\end{array}
\]

The ions form a lattice with two chloride ions for each magnesium ion. So the formula of the compound is \( \text{MgCl}_2 \). It has no overall charge.

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.

<table>
<thead>
<tr>
<th>Group</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺</td>
<td></td>
<td>Be⁺²</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>none</td>
</tr>
<tr>
<td>sodium</td>
<td>Na⁺</td>
<td>Mg⁺²</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>none</td>
</tr>
<tr>
<td>K⁺</td>
<td></td>
<td>Ca⁺²</td>
<td></td>
<td>transition elements</td>
<td></td>
<td></td>
<td></td>
<td>none</td>
</tr>
<tr>
<td>hydrogen</td>
<td>H⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

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:

<table>
<thead>
<tr>
<th>Ions in compound</th>
<th>Name of compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>K⁺ and F⁻</td>
<td>potassium fluoride</td>
</tr>
<tr>
<td>Ca⁺² and S²⁻</td>
<td>calcium sulfide</td>
</tr>
</tbody>
</table>

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

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

**Example 1**
1. Lithium fluoride.
2. The ions are Li⁺ and F⁻.
3. One Li⁺ is needed for every F⁻, to make the total charge zero.
4. The formula is LiF.

**Example 2**
1. Sodium sulfide.
2. The ions are Na⁺ and S²⁻.
3. Two Na⁺ ions are needed for every S²⁻ ion, to make the total charge zero: Na⁺ Na⁺ S²⁻.
4. 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 Zn²⁺ ions.

But most transition elements can form more than one type of ion.
For example, copper and iron can each form two:

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
<th>Example of compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu⁺</td>
<td>copper(I) ion</td>
<td>copper(I) oxide, Cu₂O</td>
</tr>
<tr>
<td>Cu²⁺</td>
<td>copper(II) ion</td>
<td>copper(II) oxide, CuO</td>
</tr>
<tr>
<td>Fe²⁺</td>
<td>iron(II) ion</td>
<td>iron(II) chloride, FeCl₂</td>
</tr>
<tr>
<td>Fe³⁺</td>
<td>iron(III) ion</td>
<td>iron(III) chloride, FeCl₃</td>
</tr>
</tbody>
</table>

The (II) in the name tells you that the ion has a charge of 2⁺.
What do the (I) and (III) show?

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.

**Example 3**
1. Sodium carbonate.
2. The ions are Na⁺ and CO₃²⁻.
3. Two Na⁺ are needed to balance the charge on one CO₃²⁻.
4. The formula is Na₂CO₃.

**Example 4**
1. Calcium nitrate.
2. The ions are Ca²⁺ and NO₃⁻.
3. Two NO₃⁻ are needed to balance the charge on one Ca²⁺.
4. The formula is Ca(NO₃)₂. Note that brackets are put round the NO₃⁻ before the ₂ is put in.
4.5 The covalent bond

Why atoms bond: a reminder
As you saw in Unit 4.3, atoms bond in order to gain a stable outer shell of electrons, like the noble gas atoms. So when sodium and chlorine react together, each sodium atom gives up an electron to a chlorine atom. But that is not the only way. Atoms can also gain stable outer shells by sharing electrons with each other.

Sharing electrons
When two non-metal atoms react together, both need to gain electrons to achieve stable outer shells. They manage this by sharing electrons.

We will look at non-metal elements in this unit, and at non-metal compounds in the next unit. Atoms can share only their outer (valence) electrons, so the diagrams will show only these.

Hydrogen
A hydrogen atom has only one shell, with one electron. The shell can hold two electrons. When two hydrogen atoms get close enough, their shells overlap and then they can share electrons. Like this:

So each has gained a full shell of two electrons, like helium atoms.

The bond between the atoms
Each hydrogen atom has a positive nucleus. Both nuclei attract the shared electrons – and this strong force of attraction holds the two atoms together. This force of attraction is called a covalent bond. A single covalent bond is formed when atoms share two electrons.

Molecules
The two bonded hydrogen atoms above form a molecule. A molecule is a group of atoms held together by covalent bonds.

Since it is made up of molecules, hydrogen is a molecular element. Its formula is $\text{H}_2$. The $2$ tells you there are 2 hydrogen atoms in each molecule.

Many other non-metals are also molecular. For example:
- iodine, $\text{I}_2$
- oxygen, $\text{O}_3$
- nitrogen, $\text{N}_2$
- chlorine, $\text{Cl}_2$
- sulfur, $\text{S}_8$
- phosphorus, $\text{P}_4$

Elements made up of molecules containing two atoms are called diatomic. So iodine and oxygen are diatomic. Can you give two other examples?
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.

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.

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.

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**.
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.

<table>
<thead>
<tr>
<th>Covalent compound</th>
<th>Description</th>
<th>Model of the molecule</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen chloride, HCl</td>
<td>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).</td>
<td><img src="image1" alt="hydrogen chloride" /></td>
</tr>
<tr>
<td>water, H₂O</td>
<td>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.</td>
<td><img src="image2" alt="water" /></td>
</tr>
<tr>
<td>methane, CH₄</td>
<td>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.</td>
<td><img src="image3" alt="methane" /></td>
</tr>
</tbody>
</table>

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°.
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.

<table>
<thead>
<tr>
<th>Covalent compound</th>
<th>Description</th>
<th>Model of the molecule</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonia, NH₃</td>
<td>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 <strong>pyramid</strong>.</td>
<td><img src="image" alt="Ammonia molecule" /></td>
</tr>
<tr>
<td>methanol, CH₃OH</td>
<td>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.</td>
<td><img src="image" alt="Methanol molecule" /></td>
</tr>
<tr>
<td>carbon dioxide, CO₂</td>
<td>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 <strong>linear</strong> molecule. All the bonds are double bonds, so we can show the molecule like this: O = C = O.</td>
<td><img src="image" alt="Carbon dioxide molecule" /></td>
</tr>
<tr>
<td>ethene, C₂H₄</td>
<td>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:</td>
<td><img src="image" alt="Ethene molecule" /></td>
</tr>
</tbody>
</table>

**Q**

1. a What is a covalent compound?  
   b Give five examples, with their formulae.
2. Draw a diagram to show the bonding in a molecule of:  
   a methane  
   b water  
3. How do the atoms gain stable outer shells, in ammonia?  
4. Draw a diagram to show the bonding in carbon dioxide.  
5. Why is the carbon dioxide molecule straight, and not bent like the water molecule?.
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:

![Sodium chloride lattice](image)

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

![Crystals of sodium chloride](image)

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.

![Water lattice with molecules](image)

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

![Crystals of ice](image)

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.)

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

These differences lead to very different properties, as you will see next.
The properties of ionic compounds

1 Ionic compounds have high melting and boiling points.
   For example:

<table>
<thead>
<tr>
<th>Compound</th>
<th>Melting point/°C</th>
<th>Boiling point/°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium chloride, NaCl</td>
<td>801</td>
<td>1413</td>
</tr>
<tr>
<td>magnesium oxide, MgO</td>
<td>2852</td>
<td>3600</td>
</tr>
</tbody>
</table>

   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.

The properties of covalent compounds

1 Molecular covalent compounds have low melting and boiling points.
   For example:

<table>
<thead>
<tr>
<th>Compound</th>
<th>Melting point/°C</th>
<th>Boiling point/°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbon monoxide, CO</td>
<td>−199</td>
<td>−191</td>
</tr>
<tr>
<td>hexane, C$<em>6$H$</em>{14}$</td>
<td>−95</td>
<td>69</td>
</tr>
</tbody>
</table>

   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.

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?

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point/°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>ice</td>
<td>0</td>
</tr>
<tr>
<td>phosphorus</td>
<td>44</td>
</tr>
<tr>
<td>sulfur</td>
<td>115</td>
</tr>
<tr>
<td>silicon dioxide (silica)</td>
<td>1710</td>
</tr>
<tr>
<td>carbon (as diamond)</td>
<td>3550</td>
</tr>
</tbody>
</table>

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, SiO₂, 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.
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.

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.

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

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.
The bonding in metals

Clues from melting points

Compare these melting points:

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

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 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.)

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.

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.
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:

   ![Diagram of a metal lattice](image)

   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:

   ![Diagram of heat transfer](image)

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.

---

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:
  hydrogen    chlorine    water
  methane    hydrogen chloride

☐ 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.
These are some properties of substances A to G.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point °C</th>
<th>Electrical conductivity</th>
<th>Solubility in water</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>–112</td>
<td>poor</td>
<td>solid</td>
</tr>
<tr>
<td>B</td>
<td>680</td>
<td>poor</td>
<td>liquid</td>
</tr>
<tr>
<td>C</td>
<td>–70</td>
<td>poor</td>
<td>liquid</td>
</tr>
<tr>
<td>D</td>
<td>1495</td>
<td>good</td>
<td>insoluble</td>
</tr>
<tr>
<td>E</td>
<td>610</td>
<td>poor</td>
<td>liquid</td>
</tr>
<tr>
<td>F</td>
<td>1610</td>
<td>poor</td>
<td>insoluble</td>
</tr>
<tr>
<td>G</td>
<td>660</td>
<td>good</td>
<td>liquid</td>
</tr>
</tbody>
</table>

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

Extended curriculum

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

\[ \text{Al} \quad \text{N} \]

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

a In which state are the two elements at room temperature (20°C)?
b Which type of structure does carbon (diamond) have: giant covalent, or molecular?
c Which type of structure would you expect to find in silicon? Give reasons.
d In which state are the two oxides, at room temperature?
e Which type of structure has carbon dioxide?
f Does silicon dioxide have the same structure as carbon dioxide? What is your evidence?

The compound zinc sulfide has a structure like this:

\[ \text{sulfide ion (S}^{2-}\text{)} \quad \text{zinc ion (Zn}^{2+}\text{)} \]

a Which does the diagram represent: a giant structure, or a molecular structure?
b Which type of bonding does zinc sulfide have?
c Look carefully at the structure. How many:
   i sulfur ions are joined to each zinc ion?
   ii zinc ions are joined to each sulfur ion?
d i From c, deduce the formula of zinc sulfide.
   ii Is this formula consistent with the charges on the two ions? Explain your answer.
e Name another metal and non-metal that will form a compound with a similar formula.

The properties of metals can be explained by the structure and bonding within the metal lattice.
a Describe the bonding in metals.
b Use the bonding to explain why metals:
   i are good conductors of electricity
   ii are malleable and flexible
5.1 The names and formulae of compounds

The names of compounds
Many compounds contain just two elements. If you know which elements they are, you can usually name the compound. Just follow these rules:

- When the compound contains a metal and a non-metal:
  - the name of the metal is given first
  - and then the name of the non-metal, but ending with -ide.
  *Examples*: sodium chloride, magnesium oxide, iron sulfide.

- When the compound is made of two non-metals:
  - if one is hydrogen, that is named first
  - otherwise the one with the lower group number comes first
  - and then the name of the other non-metal, ending with -ide.
  *Examples*: hydrogen chloride, carbon dioxide.

But some compounds have ‘everyday’ names that give you no clue about the elements in them. Water, methane, and ammonia are examples. You just have to remember their formulae!

Finding formulae from the structure of compounds
Every compound has a formula as well as a name. The formula is made up of the symbols for the elements, and often has numbers too.

The formula of a compound is related to its structure. For example:

Sodium chloride forms a giant structure with one sodium ion for every chloride ion. So its formula is NaCl.

Water is made up of molecules in which two hydrogen atoms are bonded to an oxygen atom. So its formula is H₂O.

Silicon dioxide (silica) forms a giant structure in which there are two oxygen atoms for every silicon atom. So its formula is SiO₂.

Note the difference:

- In giant structures like sodium chloride and silicon dioxide, the formula tells you the ratio of the ions or atoms in the compound.
- In a molecular compound, the formula tells you exactly how many atoms are bonded together in each molecule.

Valency
But you don’t need to draw the structure of a compound to work out its formula. You can work it out quickly if you know the valency of the elements:

The valency of an element is the number of electrons its atoms lose, gain or share, to form a compound.
Hydrogen sulfide is a very poisonous colourless gas. It smells of rotten eggs.

Look at this table. (You can check the groups in the Periodic Table on page 31.)

<table>
<thead>
<tr>
<th>Elements</th>
<th>In forming a compound, the atoms ...</th>
<th>So the valency of the element is ...</th>
<th>Examples of compounds formed (those in blue are covalent, with shared electrons)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group I</td>
<td>lose 1 electron</td>
<td>1</td>
<td>sodium chloride, NaCl</td>
</tr>
<tr>
<td>Group II</td>
<td>lose 2 electrons</td>
<td>2</td>
<td>magnesium chloride, MgCl₂</td>
</tr>
<tr>
<td>Group III</td>
<td>lose 3 electrons</td>
<td>3</td>
<td>aluminium chloride, AlCl₃</td>
</tr>
<tr>
<td>Group IV</td>
<td>share 4 electrons</td>
<td>4</td>
<td>methane, CH₄</td>
</tr>
<tr>
<td>Group V</td>
<td>gain or share 3 electrons</td>
<td>3</td>
<td>ammonia, NH₃</td>
</tr>
<tr>
<td>Group VI</td>
<td>gain or share 2 electrons</td>
<td>2</td>
<td>magnesium oxide, MgO; water, H₂O</td>
</tr>
<tr>
<td>Group VII</td>
<td>gain or share 1 electron</td>
<td>1</td>
<td>sodium chloride, NaCl; hydrogen chloride, HCl</td>
</tr>
<tr>
<td>Group 0</td>
<td>(do not form compounds)</td>
<td>–</td>
<td>none</td>
</tr>
<tr>
<td>hydrogen</td>
<td>lose or share 1 electron</td>
<td>variable</td>
<td>hydrogen bromide, HBr</td>
</tr>
<tr>
<td>transition elements</td>
<td>can lose different numbers of electrons</td>
<td>variable</td>
<td>iron (I) chloride, FeCl₃; iron (II) chloride, FeCl₂; copper (I) chloride, CuCl; copper (II) chloride, CuCl₂</td>
</tr>
</tbody>
</table>

Writing formulae using valencies
This is how to write the formula of a compound, using valencies:

1. Write down the valencies of the two elements.
2. Write down their symbols, in the same order as the elements in the name.
3. Add numbers after the symbols if you need to, to balance the valencies.

**Example 1** What is the formula of hydrogen sulfide?

1. Valencies: hydrogen, 1; sulfur (Group VI), 2
2. HS (valencies not balanced)
3. The formula is **H₂S** (2 × 1 and 2, so the valencies are now balanced)

**Example 2** What is the formula of aluminium oxide?

1. Valencies: aluminium (Group III), 3; oxygen (Group VI), 2
2. AlO (valencies not balanced)
3. The formula is **Al₂O₃** (2 × 3 and 3 × 2, so the valencies are now balanced)

Writing formulae by balancing charges
In an ionic compound, the total charge is zero. So you can also work out the formula of an ionic compound by balancing the charges on its ions. To find out how to do this, turn to Unit 4.4.