Macmillan CXC Science Series

Chemistry

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This new series of textbooks for Caribbean Examinations Council [CXC] General Proficiency examinations has been developed and written by teachers with many years’ experience of preparing students for CXC (CSEC) examinations in Caribbean schools.

A textbook is used in different ways at different times.

- Readers may be starting a topic from scratch, and need to be led through a logical explanation one step at a time.
- Students with a working knowledge of a topic may need to clarify a detail, or reinforce their understanding. Or, they may simply need to believe that they do have a good grasp of the material being studied.

The specially created format is the same for all of the books in the series.

- Diagrams and pictures are placed on the page in such a way that they can be consulted as the reader wishes but interrupt the text as little as possible.
- Short-answer questions (called In-Text Questions even though they are not placed in the main body of the page) allow the student to test his or her grasp of the topic. A student who can answer an ITQ gains confidence; a student who cannot knows to go back over the topic and try again.
- The first use of any important technical term is highlighted to make subsequent revision easier.

Teachers throughout the region emphasised that inclusion of SBA material would be essential to these books. Each CXC (CSEC) science syllabus specifies the areas in which an SBA exercise is expected. Accordingly, at least one exercise is included for each area.

The books will also provide a firm foundation for more advanced study appropriate to the Caribbean Advanced Proficiency Examinations (CAPE).

Dr Mike Taylor
Series Editor
This book isn’t just words on a page. Here are some important features. Each will help you, in its own way, if you take advantage of it.

There are TWO COLUMNS.

The bigger column has the text and some really large diagrams; you can read straight down it without interruption.

The smaller column has other diagrams which the text mentions. Look at them carefully as you need them. You may find that a few seconds looking at a diagram is worth a few minutes’ reading.

The first time that an important NEW WORD occurs, it is repeated in the smaller column. If you want to check what a word means, you can find it quickly.

There are QUESTIONS called ITQs, These are ‘In-Text Questions’. When you have read the nearby big-column paragraph, try to answer the question, in your head or on paper, just as you wish. If you can, you’re on the road to understanding. If you can’t, just go back and read that bit again. Answers are at the end of each chapter, so you can tell how good your answer was.

Some possible SBA EXERCISES are included. They are printed in a separate chapter on paper with tinted edges They have outline instructions and questions to answer. Don’t copy them! Use them as models for designing your own work.

There is a detailed INDEX. Don’t be afraid to use it to find what you want.

At the end of each chapter there are some EXAMINATION-STYLE QUESTIONS. Your teacher will suggest how you can use them.
By the end of this chapter you should be able to:

- explain what is meant by the ‘states of matter’;
- describe physical differences between solids, liquids and gases;
- explain what is meant by the ‘kinetic theory’;
- describe and explain evidence for the kinetic theory;
- describe the states of matter in terms of the arrangement of particles;
- name and describe changes of state;
- explain the terms ‘melting point’ and ‘boiling point’;
- describe and explain the energy changes associated with change of state;
- explain the term ‘latent heat’;
- sketch and explain the shape of a typical heating curve for a change of state;
- explain that bond-making liberates energy, bond-breaking absorbs energy;
- describe a method for finding the melting point of a substance.

States of matter

The three states of matter

‘Matter’ is anything which has mass and takes up space. Most of the matter in our world is either solid, liquid or gas. These are called the three states of matter’. Most of the time, we can easily tell which is which:

- this book is solid;
- sea-water is liquid;
- air is a gas.

How can we tell the three states apart? We can tell because the states have different physical properties. Table 1.1 shows how the volume and shape of a substance changes when it is placed in a container.
States of matter

<table>
<thead>
<tr>
<th>Property</th>
<th>Solid</th>
<th>Liquid</th>
<th>Gas</th>
</tr>
</thead>
<tbody>
<tr>
<td>shape</td>
<td>keeps own shape</td>
<td>takes shape of container</td>
<td>takes shape of container</td>
</tr>
<tr>
<td>volume</td>
<td>keeps own volume</td>
<td>keeps own volume</td>
<td>takes volume of container</td>
</tr>
</tbody>
</table>

Table 1.1 The three states of matter in a container.

You can probably think of some materials which do not fit easily into one of these three categories.

- If you cool molasses, the liquid flows less and less easily as it gets colder. When does it become a solid?
- Glass, which looks and feels solid, is a type of liquid. A very old pane of window glass is measurably thicker at the bottom than at the top.
- ‘Potty Putty’ (figure 1.1) looks solid, and if hit with a hammer shatters into bits; but left on a bench it slowly spreads out into a puddle.

For the moment we do not need to study these substances.

Matter is made of particles

We can explain the behaviour of solids, liquids and gases using the ideas that:

- matter is made up of particles; and
- the particles are in constant motion.

Robert Brown (1773–1858), a Scottish botanist, found good evidence for this when he was using a microscope to look at pollen grains in water. He saw that the grains were always moving – but in random, zig-zag paths. Brown suggested that the water was made of particles which were bombarding the pollen grains and knocking them about. We can see the same zig-zag motion if we look at smoke particles in the air through a microscope. It is called Brownian motion.

The theory that all matter is made up of separate, moving particles is called the kinetic theory of matter.

Whether a material is solid, liquid or gas depends on how regularly the particles are fitted together and how strongly they are bonded to each other. We will look at each of the three states of matter and then look for more evidence that the kinetic theory is a good theory.

Solids

In a solid, the particles are touching each other and they fit together in a regular way. Because there are no big spaces between particles, the density of the solid is high. Solids in which the ‘fit’ is very good (the metals) have the highest densities of all. Because there is very little space between the particles, a solid cannot be compressed by squeezing. Squeezing or hammering a solid can change its shape (think of hammering a metal or a piece of glass or squeezing some putty) but its volume doesn’t change.

Because a solid normally keeps its own shape we know that the particles in a solid must be fixed in place with respect to their neighbours. There are strong bonds between them.

Once in position, there the particles stay, but they are not completely still. Each particle vibrates to and fro. If we heat the solid (add heat energy to it) the vibrations get stronger, and if we cool a solid (take heat energy out of it) the vibrations get weaker. When we say that something is ‘hot’ we mean that its particles are vibrating strongly. A substance in which the particles are only vibrating weakly is ‘cold’.

Figure 1.1 Some ‘Potty Putty’. Is this a solid or a liquid?

Figure 1.2 Brownian motion.

Figure 1.3 Particles in a solid.
Liquids
In a liquid the particles are almost as close together as in a solid, but there is no order in the way they are arranged. The particles in a liquid are slightly further apart than they are in a solid, so a liquid usually has a slightly lower density than its solid and the solid sinks in its own liquid. (Water is an exception – ice floats on water.)

The spaces between the particles in a liquid are not very large and so a liquid can only be compressed a very little. If liquids could be compressed a lot then hydraulic systems such as car brakes could not work.

As we all know, liquids do not keep their own shape – they flow. This means that the particles can move relative to each other and so they can change neighbours freely. But as it flows, a liquid keeps the same volume. Its particles stay close together. Like the particles in a solid, the particles in a liquid have strong bonds between them.

Gases
Gases have much lower densities than either liquids or solids. You can see this in action when bubbles rise through fizzy drinks! Gases can be compressed into small fractions of their ordinary volumes (think of pumping up a bicycle tyre). Gases can be compressed because the particles are far apart from each other.

Now imagine opening a bottle of perfume at one end of a room – it will not be long before you can smell it at the other end of the room. This tells us that the forces between the gas particles are weak, because if they were strong the particles would stick together. Instead, the particles are free to move anywhere in the room. If we could watch the movement of any of them, we would see something very like Brownian motion (see figure 1.2). The gas particles would move about randomly, in straight lines, only changing direction when they bumped into one another, into air particles, or bounced off the walls of the room.
Summary

Now we can summarise the properties of the three states of matter (table 1.2).

<table>
<thead>
<tr>
<th>Property</th>
<th>Solid</th>
<th>Liquid</th>
<th>Gas</th>
</tr>
</thead>
<tbody>
<tr>
<td>shape</td>
<td>keeps own shape</td>
<td>takes shape of container</td>
<td>takes shape of container</td>
</tr>
<tr>
<td>volume</td>
<td>keeps own volume</td>
<td>keeps own volume</td>
<td>takes volume of container</td>
</tr>
<tr>
<td>density</td>
<td>high to very high</td>
<td>medium to high</td>
<td>low</td>
</tr>
<tr>
<td>can be compressed</td>
<td>not at all</td>
<td>only a little</td>
<td>a great deal</td>
</tr>
<tr>
<td>force between</td>
<td>strong</td>
<td>strong</td>
<td>weak</td>
</tr>
<tr>
<td>particles</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>motion of particles</td>
<td>vibrate fixed with respect to each other</td>
<td>vibrate move within liquid volume</td>
<td>vibrate move anywhere</td>
</tr>
</tbody>
</table>

*Table 1.2* Properties of the three states of matter.

**Evidence for the kinetic theory**

We have already learned about Brownian motion. When you look through a microscope at pollen grains or smoke particles we see them moving. What we do *not* see is also important – we do *not* see the particles of water or air! This tells us that these particles are very small. Even with the most powerful microscopes, you will not be able to see them. Instead, we rely on indirect evidence to show that particles exist. We can use diffusion and osmosis as evidence.

**Diffusion**

If we pour something smelly into a dish and leave it in one corner of a room, we shall soon smell it from the opposite corner of the room. This is because the ‘smell particles’ in the gas travel in all directions, bumping into air particles as they go, until they are spread throughout the room.

If we fill a jar with water and then use a glass tube to make a layer of coloured ink at the bottom of the jar, we shall see that over a few days the colour of the ink will spread through the liquid until the whole of the liquid is a uniform colour. (Try it!) The coloured particles, like the smelly gas particles, bump about in all directions until they fill the available liquid. This process is called diffusion.

**diffusion** The movement of particles from a region of high concentration to a region of lower concentration.

*Figure 1.10* Diffusion in a liquid.
We can explain diffusion using the kinetic theory. Imagine the layer where the coloured liquid (we will use red) and the colourless water meet. This layer is AB in figure 1.11. At first there are many red particles below AB and none above it. All the rest of the particles, which not shown in the diagram, are colourless. The red particles are moving randomly – but at least one will be moving upwards.

After a little while the situation will look like figure 1.12. One red particle has moved into the top layer. Its place has been taken by a colourless water particle because some of the water particles on the layer AB were bound to have been moving downwards.

This process will go on until we reach the last stage, which is shown in figure 1.13. Now there are as many red particles above AB as there are below it, so at any moment there will be an equal number of red particles passing up and passing down through AB. The colour is uniform everywhere and will remain so.

This is exactly what happens when we do the experiment, so we believe that our explanation, based on the kinetic theory, is correct. The red particles, which were all concentrated in one part of the liquid, have now spread themselves out so that their concentration is the same everywhere.

Diffusion in gases

As we have seen already, if we pour something smelly into a dish and leave it in one corner of a room, we soon smell it from the opposite corner. This is because the smelly particles in the gas travel in all directions, bumping into air particles as they go, until they are spread throughout the room.

By making gas particles travel in a long tube we can get some information about their relative sizes. Practical activity 1.1 allows us to do this.

\[\text{Figure 1.11} \quad \text{Before diffusion begins.}\]

\[\text{Figure 1.12} \quad \text{Diffusion has begun.}\]

\[\text{Figure 1.13} \quad \text{Diffusion is now complete.}\]
Osmosis

Try cutting some chips of Irish potato, pawpaw or carrot and leaving them for an hour in a strong solution of common salt (sodium chloride). You will see that the chips shrink as if they were partly dried out. Now put them into clean, unsalted water for an hour and they will be restored. It seems as though water can be made to flow out of or into the chips by changing the conditions of their surroundings.

Now try practical activity 1.2.

Practical activity 1.1
Diffusion in gases

Concentrated hydrochloric acid gives off hydrogen chloride gas. Aqueous ammonia gives off ammonia gas. These two gases react together to give ammonium chloride, which is a white solid.

\[ \text{HCl(g)} + \text{NH}_3(\text{g}) \rightarrow \text{NH}_4\text{Cl(s)} \]

You need:
- concentrated hydrochloric acid (CARE)
- concentrated aqueous ammonia (CARE)
- glass or clear plastic tube, about 40 cm long
- cotton wool
- tweezers
- ruler
- 2 droppers

Method
1. Set up the tube horizontally.
2. Make two small balls of cotton wool of a size to fit into the tube.
3. Use a dropper to saturate one ball with concentrated hydrochloric acid.
4. Use the tweezers to put the wet cotton wool ball into one end of the tube.
5. Wash the tweezers and dry them.
6. Use the other dropper to saturate the second ball with concentrated aqueous ammonia.
7. Use the tweezers to put this wet cotton wool ball into the other end of the tube.
8. Watch what happens!

Questions
1. What did you observe happening in the tube?
2. Did this happen at once or after a short time? Why was that?
3. Is the effect you observe nearer to the source of the ammonia or the source of the hydrogen chloride?
4. What does the experiment tell you about the speeds at which the particles diffused?
5. What does this suggest about the relative sizes of the gas particles?
Practical activity 1.2

Investigating the flow of water across ‘Visking’ tubing

You need:
- two people
- thistle funnel
- length of Visking tubing
- coloured ink
- ruler
- thin string or cotton thread
- 250 cm³ beaker
- clamp and stand
- common salt

Method
1. Add ink to a strong solution of common salt to make a brightly coloured salt solution.
2. Half-fill a 250 cm³ beaker with clean water.
3. Cut a short length of Visking tubing lengthways to make a sheet about 6 cm across.
4. Hold a thistle funnel with your finger across the bottom of the stem.
5. Pour the brightly coloured solution into the funnel end of the thistle funnel, until the stem and about half of the ‘thistle’ are full. Make sure that you get rid of any air bubbles in the liquid.
6. Keep your finger over the end of the funnel whilst your friend stretches the Visking sheet over the end of the ‘thistle’ and ties it tightly in place with thread. It is important to make a watertight seal.
7. Turn the thistle funnel upside down and check that it does not leak.
8. If all is well, lower the funnel into the water in the beaker until the levels of liquid in the beaker and inside the funnel are the same.
9. Clamp the thistle funnel gently in place. It should look like this diagram.

Question
Why was it important to begin the experiment with the liquid in the beaker and the liquid in the funnel at the same level?

Observations
Make measurements of the movement of the level of the coloured liquid in the funnel over the course of an hour. If you wish you can draw a graph of the difference in level against time.

You will see that the level of the liquid in the stem of the funnel rises for a while and then becomes constant. No colour appears in the salt solution in the beaker.

Some possible results are shown in figure 1.14.
**States of matter**

**What can we deduce from our observations in practical activity 1.2?**
- Firstly, water flowed into the funnel from the beaker, because the level of liquid in the funnel increased.
- Secondly, no coloured material passed out from the funnel into the water in the beaker.

So, the Visking must allow water to pass through, without allowing other substances through. Visking is called a selectively permeable substance.

**Can we explain these observations using the kinetic theory?**
Yes. Think about the Visking sheet. At first it was being bombarded on its bottom side by water particles only. On its top side it was being bombarded by both water particles and salt particles. The water was more concentrated below the sheet than above it. In any one second, therefore, more water particles hit the bottom side of the Visking sheet than hit the top side. Just as in diffusion, more water passed up through the Visking than passed down. This net flow of water into the funnel caused the water level in the thistle funnel to rise.

You can see that osmosis is really diffusion applied to water.

**Crystal cleavage**
Many crystals can be split by placing a knife-edge on one crystal face and tapping the knife sharply with a small hammer. The crystal splits ('cleaves') neatly into two pieces. But there's more. The cleaved faces are flat; and cleavage only works if the knife-edge is set in particular directions.

Imagine a stack of small cubes lightly stuck together with weak glue (figure 1.16). Now imagine putting a sheet of stiff cardboard on the stack and tapping it with a hammer. If the card is parallel to the joins between the cubes it will cut the stack into two. If not, either the card will bend or the stack will shatter (figure 1.17). This is because the stack is made up from small units. Crystal cleavage happens for the same reason and reinforces our belief in particle theory.

**Change of state**
If you take a lump of ice and leave it in the sun, what happens?... It turns into water. If you now heat the water in a pan it will turn into water vapour. Now, cool the water vapour and you get liquid water. Finally, if you put the water in a freezer, you get ice again. Yet the only chemical substance in your experiments was water. The changes you saw were physical changes because no new substances were created. The ice, the water and the water vapour are simply different states of water. You carried out changes of state.
1 · States of matter

Some substances can change directly from solid to vapour when heated, without going through a liquid stage. When the vapour is cooled the solid forms again. This is called sublimation. One substance which sublimes when it is heated is iodine. Iodine exists as black, shiny crystals and these change directly into a beautiful purple vapour when they are heated.

![Sublimation diagram](image)

*Figure 1.18* A summary of the changes of state.

**Energy in changes of state**

Heat energy is the energy that flows to one body from a hotter body. When you boil water, heat energy passes from the flame into the water. Hot water contains more heat energy than an equal volume of cold water. What does the extra energy do to the particles?

![Energy in changes of state diagram](image)

*Figure 1.19* Iodine as a solid and a vapour, but not a liquid.

**ITQ4**

The density of water is 1.0 g dm\(^{-3}\). An experimenter found that 0.06 g of water formed 100 cm\(^3\) of steam at 100 °C. What was the density of the steam, in g cm\(^{-3}\)?

*A simple experiment will help us to understand. If we take some ice at a temperature below 0 °C and heat it until we have boiled it completely away, the change of temperature with time would be similar to figure 1.20.*

*Figure 1.20* A heating curve for water.

A similar curve would be obtained for any solid-to-gas-to-liquid change. As we heat the ice its temperature rises (part 1 of the curve). When the ice is melting its temperature stops rising and stays constant (part 2) even though we are still adding heat at the same rate. Once all the ice is melted the temperature of the liquid rises again (part 3) until it begins to boil. When the liquid boils (part 4) the temperature is constant once again.

*Figure 1.21* Bond breaking uses energy.
**What is happening to the particles?**

In the ice, the water particles can only vibrate. As they get warmer, the added energy makes their vibration stronger and stronger. Suddenly (at $T_1$) they have enough energy to break the bonds which hold them one to another. The solid starts to melt; while it is melting all the incoming energy is used up breaking bonds and none is used to raise the temperature. Only when all the solid has melted does the incoming energy start to warm the water any further.

The same thing happens to the liquid. Heating increases the motion of the particles until they have enough energy to break free from each other and become a gas. Once that starts to happen (at $T_2$) the incoming energy goes to breaking the bonds between the particles, not to making them hotter.

**Where is the heat energy?**

It may seem surprising that heat energy can be put into something without making it any hotter. However, the heat is not lost; it was used up separating the particles from each other and so it is stored in the separated particles. We call it latent heat.

Because the heat is stored in the particles, we can get it back again. When a liquid freezes, it does so at the same temperature at which it formed ($T_1$ in our graph). As the particles form bonds and turn from liquid to solid they give up their latent heat again.

**What is the difference between boiling and evaporating?**

A liquid boils when all the heat energy going into it is used up in turning the liquid into a vapour, without changing the temperature. Usually when a liquid boils we see bubbles rising through the liquid and bursting at the surface. But liquids can turn into vapour at any temperature above their melting point. For example, a wet cloth will dry in the air even though the temperature of the air is much less than 100 °C. This happens because some of the particles in the surface layer have enough energy to break free from the attraction of the particles around them, and escape. The nearer the temperature is to the boiling point, the more such energetic particles there will be. This process of turning from liquid to vapour below the boiling point is called evaporation.

**Figure 1.22** Making bonds; releasing energy!

The particles which evaporate are the ones with the most kinetic energy – they are ‘hot’. This explains why evaporation makes a liquid cool down. It is losing its hottest molecules.

**Melting point and boiling point**

The temperature at which a pure substance melts is called its melting point. The temperature at which it boils is called its boiling point. It is a very useful fact that pure substances have sharp melting points. Impure substances melt over a range of temperature. This gives us a good test for chemical purity. You can investigate this in practical activity 1.3.
Practical activity 1.3
To test the purity of a crystalline substance

You need:
- beaker
- cooking oil
- 0–110 °C thermometer
- metal wire stirrer
- capillary tubing
- small rubber band
- Bunsen burner
- tripod and gauze
- sample (e.g. stearic acid)

Method
1 Cut a length (about 6 cm) of capillary tubing. Spin one end of it in the edge of a Bunsen flame until it is sealed. Let it cool.
2 Put a little of your powdered test substance on a hard surface. Push the open end of the capillary tube into the powder. Turn the tube over and bounce its sealed end on the hard surface to shake the test material down to the bottom.
3 Fix the capillary tube to the thermometer with the rubber band, as shown in the diagram.
4 Put the thermometer and capillary into the oil and heat the oil slowly. Stir the oil constantly.
5 Watch for when the substance melts. The first indication is usually that the substance seems to shrink in the capillary tube. If the temperature interval between the start and end of melting is only about 0.5 °C, the substance is pure. If the solid melts over a range of 2 °C or more, it is impure.

Questions
1 Why is it necessary to heat the oil slowly?
2 Why is it a good idea to keep the capillary tube very close to the thermometer?

Summary
- Matter is made up of particles.
- The particles of matter are in constant random motion.
- Matter exists in three states: solid, liquid and gas.
- The states of matter are distinguished by their physical properties.
- Changes of state are accompanied by the absorption or release of heat energy.
- Heat energy absorbed or released during a change of state does not change the temperature of the substance.
- Heat energy absorbed or released during a change of state is called latent heat.
- Pure materials have sharp melting points and boiling points.
The particles of a liquid are already touching each other with hardly any space between them. The particles of a gas have large spaces between them.

(i) After 2 minutes the colour would have begun to move up into the jar containing air but the colour will have filled the vacuum jar and the colour will be uniform throughout. This is because there were no air particles to bump into the diffusing bromine particles.

(ii) After 2 days the colour will be uniform in both sets of jars.

As the height of the water column inside the thistle flask increased, its pressure on the Visking increased until the rates at which water molecules hit the Visking membrane were the same on both sides.

0.06 cm³ of water has a mass of 0.06 g
The steam had the same mass but occupied 100 cm³
The density of the steam is 0.06 g/100 cm³ = 0.0006 g cm⁻³

The steam not only scalds you, it also deposits its latent heat of vaporisation onto your skin, which makes matters worse.

1 (i) Use the particulate theory of matter to explain why gases have much lower densities than do solids.

(ii) Some very fine, dry sand was poured into a measuring cylinder until the cylinder contained 100 cm³ of the sand. 10 cm³ of water were then poured into the cylinder. The volume of the mixture was 105 cm³. Why was the final volume less than 110 cm³?

(iii) A beaker was half-filled with water and allowed to stand on the bench for some time. A very fine tube was then used to squirt a drop of ink onto the bottom of the beaker (diagram A). After some time the appearance of the liquid had changed, as shown in diagram B. Explain:
(a) what had caused the colour to spread;
(b) why the colour spread only slowly through the liquid.

2 (i) Explain what is meant by a selectively permeable membrane.

(ii) If the shell is dissolved away from a hen’s egg and the egg is placed in a concentrated solution of sodium chloride (A), the egg shrinks. If the shrunken egg is placed in pure water (B) it swells again.
(a) The shell of the egg is mostly calcium carbonate. Carbonates dissolve in acids. Suggest something found in every kitchen which might be used to dissolve the eggshell.
(b) Use the difference in concentration of water in the two liquids A and B to explain the changes seen in the volume of the egg.

3 (i) A sample of ethanol (boiling point 80 °C) was poured into a beaker and gently heated for two minutes. After this time the temperature of the ethanol was 60 °C. How is the movement of the ethanol molecules different at the higher temperature from their movement at room temperature?

(ii) The alcohol is heated more strongly. Draw a sketch graph showing how its temperature will probably change over the next four minutes. Label the axes carefully.

(iii) Why, when the liquid is heated, does its temperature not always increase? If the temperature is not increasing, what effect is the added heat energy having?
By the end of this chapter you should be able to:

- explain what is meant by an element, a mixture and a compound;
- explain what is meant by an atom and a molecule;
- describe the internal structure of an atom;
- state roughly how large atoms are;
- know the meaning of the terms:
  - relative atomic mass,
  - atomic number,
  - mass number;
- draw electronic structures of atoms up to \( Z = 20 \);
- explain what is meant by an ion;
- write and interpret symbols for atoms and simple ions.

**Concept map**

- separation methods
- elements, compounds and mixtures
- solutions
- history of atoms
- atoms have structure
  - atomic nucleus
  - size of atoms
  - electronic structure
  - atoms and ions
  - symbols for atoms and ions

**Elements, mixtures and compounds**

Throughout the ages, craftsmen and craftswomen have carried out chemical changes. Producing metals, dyeing cloth, making glass, tanning leather, cooking food; all these ancient arts are based on chemical changes. Some people were not content to use the skills they had been taught. They wanted to find out *why* things happened. They were the first real chemists.

Another important part of chemistry is separating materials. Some materials can easily be separated into simpler things.

- Red beans and grains of salt are easily sorted out.
- With the help of a magnifying glass, grains of sand can be sorted into fragments of different colours.
- Ink and water can be separated by boiling off the water and collecting the steam; this will work whether the ink is strong or dilute.

**Figure 2.1** Chemistry has been carried out for hundreds of years.
Mixtures can easily be separated into their components. Mixtures do not have a fixed composition. Moreover, the substances in a mixture keep their own properties. In a mixture of iron and sulphur, the sulphur is still yellow and the iron is still magnetic.

Solutions are a special kind of mixture. If we stir sugar into water the sugar seems to vanish. If we mix water and alcohol, we see only one liquid. The particles of the sugar have separated one from another and have vanished amongst the particles of the water. The particles of water and the particles of the alcohol have become jumbled up together. We call the water the solvent. The dissolved substance is the solute.

A solution is called a homogeneous mixture because it is apparently the same all through. In contrast, the grains of sand form a heterogeneous mixture. Different parts of the mixture can be distinguished one from another.

Elements

Some materials cannot be separated into simpler things, either mechanically or chemically. Whatever you do to a piece of iron, you cannot split it into other chemicals. Iron is an example of an element.

An element is a single substance that cannot be split into simpler substances.

Atoms and molecules

The smallest particle of an element which can exist is called an atom of that element. Some elements (like iron) can exist as single atoms. Others, such as oxygen and phosphorus, exist as groups of atoms called molecules.

Compounds

When atoms of two or more different elements unite, the result is a chemical compound.

- Compounds have different properties to the elements which make them.
- Compounds have a fixed composition, no matter how they are made.

If we heat a mixture of iron and sulphur it turns into a black solid called iron sulphide. Iron sulphide is not magnetic and it always contains one atom of sulphur for every atom of iron – it is a compound of iron and sulphur. There is no physical way in which we can get the iron and the sulphur back again from iron sulphide.
Practical activity 2.1
Separating salt and sand

You need:
- sample of salt
- sample of sand
- 2 beakers
- Bunsen burner
- tripod and gauze
- filter funnel and filter paper

Method
1. Mix approx. 5 g each of salt and sand.
2. Look at the mixture. Can you distinguish the particles of sand from the salt?
3. Add a spoonful of the sand/salt mixture to 100 cm³ water in a beaker. Stir well.
4. Fold a piece of filter paper (see figure 2.9, page 18), place it in the funnel, put a beaker under the funnel, and pour the liquid and any solid into the funnel. Collect the liquid in a beaker.
5. Hold the funnel over the sink (or another beaker) and pour a little water through it.
6. Boil the liquid that you collected in step 4. When it begins to ‘spit’ turn down the Bunsen burner and heat gently until the solid is dry.

Questions
1. What is left in the filter paper?
2. What was the purpose of the water in step 5?
3. Is the solid left in the filter funnel ‘pure’?
   What must you do to make it exactly the same as the original sand?
4. What is the white solid produced in step 6?

Extension work
Design a method for separating the following:
- a mixture of red pencils and black pencils;
- mixture of iron filings and sulphur powder;
- a mixture of ink and water (hint: ink is a solution of a dye in water).

We shall do more about separating mixtures in Chapter 12.

What is a solution?

When we stir salt in water, the salt disappears. However, the taste of the water shows that the salt is not lost – it is still there.

We say that the salt has dissolved the water. It has formed a solution. The salt is the solute. The water is the solvent. There is usually a lot more solvent than solute.

As we saw earlier in this chapter, a solution is a special type of mixture; it is homogeneous (the same all the way through). You cannot see the individual chemical particles of salt. All the solution looks exactly the same. Filtering the solution has no effect; the sodium chloride cannot be recovered except by boiling off the water.

A solvent may be a solid, a liquid or a gas. Table 2.1 shows some examples.

<table>
<thead>
<tr>
<th>Solute state</th>
<th>Solvent state</th>
<th>Example</th>
</tr>
</thead>
</table>
| gas          | liquid        | oxygen in water  
carbon dioxide in ‘fizzy’ water |
| gas          | solid         | hydrogen in metals (e.g. in palladium) |
| liquid       | liquid        | rum (alcohol in water)  
petrol (various hydrocarbons in each other) |
| solid        | liquid        | sea-water (sodium chloride in water)  
‘tincture of iodine’ (iodine in alcohol) |
| solid        | solid         | metal alloys, e.g. brass (zinc in copper)  
steel (carbon in iron) |

Table 2.1 Examples of solutions.
When the solvent is a liquid the solution is said to be:
• aqueous if the solvent is water;
• non-aqueous if the solvent is other than water.

What influences how much will dissolve?
The most important things are:
• temperature;
  • solids usually are more soluble in liquids as the temperature goes up;
  • gases are less soluble in warm water than in cold water;
  • pressure – gases dissolve more in liquids at higher pressure;
• ‘like dissolved like’;
• ionic materials dissolve in water;
• covalent materials dissolve in covalent liquids.

Saturated solutions
If we stir just a little salt into some water, all the salt dissolves. If we go on adding more salt, then eventually some solid salt is left. The solution has become saturated. The same mass of water will dissolve different amounts of different substances. For example, 100 g of water will dissolve 37 g of salt, but it will dissolve 137 g of sugar or 24 g of blue copper(II) sulphate, before it is saturated.

The mass of solute needed to form a saturated solution with 100 g of solvent is the solubility of the solute. Figure 2.6 shows how the solubility of salt and copper sulphate change with temperature.

Suppose that we have a solution which is saturated at a high temperature, and we allow it to cool. At a lower temperature the solution cannot hold all the solute, so the excess crystallises out (forms solid crystals). If this happens quickly we get a mass of small crystals, but if it is done slowly, a smaller number of larger crystals is obtained.

Similarly, if solvent is evaporated from a solution, a stage is reached when the solvent that is left cannot hold all the solute, and again crystallisation occurs. If this is done very slowly, some large crystals can be grown. You can try this yourself by leaving a saturated solution in a beaker on a windowsill, out of the sun, partly covered by a piece of filter paper.

Supersaturation
Crystals can only form if there are some tiny nuclei to start them off. A solution which holds more solute than a saturated solution is said to be supersaturated. If a piece of dust or a crystal fragment is dropped into a supersaturated solution, it crystallises at once.
Suspensions

If we shake mud with water, we do not get a homogeneous solution. The water goes cloudy, you’ll see bits floating in the water and some solid sinks to the bottom. Once the solid has settled we can pour off the cloudy water, but the water still will not become clear. The particles of mud are suspended in the water. We have made a suspension. If we filter the suspension we can remove the mud particles and get clear water again. This is because the suspended particles are so large that they cannot pass through the filter, unlike the dissolved sodium chloride in our first example.

Colloids

In between true solutions and suspensions is a class of solution called a colloid. A colloidal solution has particles small enough to pass through a filter, but large enough to scatter light. This can be demonstrated by looking at Tyndall’s cone. Figure 2.8 shows a convergent beam of light produced by a lens. In its path are two glass cells, one containing a true solution and one containing a colloid. The light is not scattered by the solute particles in the true solution, but it is scattered by the larger particles of the colloid.

There are two reasons why a substance forms a colloid.

- It may have small particles which form clusters of colloidal size. These are usually held apart from each other in the solution by electrical charges.
- The molecules of the substance itself may be of colloidal size. Starch is an example.

Just as for true solutions, the solvent need not be a liquid. Table 2.2 shows other possibilities. We say that the solute is dispersed in the other component. Colloids are common. For example, milk is tiny globules of fat dispersed in liquid, paints are solid particles dispersed in a liquid, starch solution (used in laundries) is solid starch dispersed in water, fog is tiny water droplets dispersed in the air.

Practical activity 2.2
Supersaturation

You need:

- a strong, clean glass test-tube
- a beaker
- a thermometer
- a Bunsen burner
- sodium thiosulphate (Na₂S₂O₃·5H₂O)
- cotton wool (or similar)

Method

1. Half-fill the test-tube with sodium thiosulphate crystals. Half-fill the beaker with water and put the test-tube into it.
2. Heat the water steadily. You will see the sodium thiosulphate appear to melt – in fact it is dissolving in its own water of crystallisation.
3. When it is completely dissolved tip out the water from the beaker, put the test-tube back into it, plug the test-tube loosely with a twist of cotton wool, and leave it to cool. Even at room temperature the liquid will not crystallise.
4. When the liquid is cool either drop a fragment of crystal into it or stir it with the thermometer.

Questions

1. What did you observe when you disturbed the liquid?
2. What happened to the temperature of the liquid as that change took place?
3. What explanation can you give for the temperature change?

Look in Chapter 29 (School-Based Assessment) for an experiment on plotting a solubility curve.
Separating mixtures

To separate a mixture into its parts we have to compare the physical properties of those parts. We can compare:

- the sizes of the particles of each part;
- the boiling points of each part;
- the solubility of each part in various solvents;
- any special properties, such as magnetism or sublimation.

We may need to use a variety of methods, including:

- filtration;
- evaporation;
- distillation or fractional distillation;
- gravity separation in a separating funnel;
- chromatography.

Some examples of separation methods

Mechanical separation — for solids with large particles

A mixture of sand and peas can be separated simply by picking out the peas or pouring the mixture onto a sloping board and letting the peas roll down!

Filtration — for a solid which does not dissolve in the liquid

If one component is a liquid and the other solid, pour the liquid through a filter — usually a filter paper folded in quarters and placed in a funnel (figure 2.9). The liquid drips through. Then wash the solid to get rid of traces of the liquid before you dry it.

Evaporation — for a solid which is soluble in the liquid

If one component is dissolved in the other, the solvent can be driven off by heating to leave only the solute. In this method the solvent is lost. You can try this with salt water in practical activity 2.3.
Distillation – for a liquid containing a dissolved solute

Distillation is used to separate and collect the solvent from a solution rather than collect the solute. The solution is boiled in a flask with a side-arm coming out of its neck. The vapour is condensed and collected. The apparatus is shown in figure 2.10. Always put in a tiny chip of glass or very hard stone to stop the liquid ‘bumping’ – suddenly boiling furiously. Liquid can easily be thrown up into the condenser if this happens.

Never fill a distillation flask more than one-third full.

Bubbles, too, need nuclei to form. That is why we put a ‘boiling chip’ into a flask when we boil liquid in it. Without the sharp corners on the chip the liquid might ‘superheat’ and then suddenly boil very vigorously, throwing liquid in all directions!

The type of condenser used here was invented by Justus von Liebig in about 1860. It is still called a ‘Liebig’ condenser.

Figure 2.10 Distillation apparatus.
To use the distillation apparatus, one-third fill the flask with solution and add one or two boiling chips (no more!). Heat the flask over a gauze. When the liquid boils the vapour rises up the neck of the flask and passes over into the water-cooled condenser. The vapour turns back into liquid in the condenser and is collected. The thermometer records the temperature at which the liquid is boiling, so if you distil a solution of ink the thermometer should read 100 °C whilst vapour is passing over it (figure 2.11).

Never boil the solution in a distillation flask to dryness.

**Figure 2.11** A distillation flask and thermometer.

**Fractional distillation – for a mixture of similar liquids**

Simple distillation works well when the solute itself does not vaporise. Salt or the dyes in ink are examples. But for a mixture of similar liquids, such as ethanol and water or ethanol and propanone, we add a fractionating column between the flask and the condenser (figure 2.12). The fractionating column is a tube packed with short pieces of glass tubing to give a large surface area.

When the mixture is heated, both vapours travel up the column but the vapour of the liquid with the higher boiling point condenses and drops back into the flask. When all the low-boiling liquid has distilled off the flask can be heated more strongly and the second liquid then distils over.

**Solvent extraction – for materials each soluble in a different solvent**

This method requires that the substances are soluble in different solvents and that the solvents do not mix with one another.

Salt and iodine, ground up together, cannot be separated mechanically. Shaken with water the salt dissolves completely but only a little iodine dissolves. The solution is a pale yellow colour. At this stage we could filter off the remaining iodine, wash it and dry it, but how could we get pure salt?

Iodine is very soluble in organic liquids like ethoxyethane and 1,2,3-trichloro-ethane. Neither of these organic liquids mixes with water. We can extract the iodine using a separating funnel, as described in practical activity 2.4.

**Figure 2.12** A fractionating column.
Practical activity 2.4  
Using solvent extraction

You need:  
• separating funnel  
• aqueous solution of iodine and sodium chloride  
• 2 beakers  
• ethoxyethane or 1,1,1-trichloroethane

Method

1. Make sure the tap on the funnel is closed!  
2. Fill the separating funnel one-third full of the iodine solution.  
3. Add enough of your chosen solvent to make the funnel two-thirds full.  
4. Stopper the funnel. Turn it upside down and shake it vigorously.  
5. Stop shaking the funnel, point the tap end upwards and away from you (and away from anyone else) and momentarily open the tap. This reduces the pressure in the flask back to atmospheric pressure.  
6. Turn the funnel right way up and let the two liquids settle. Note which layer contains the iodine. Ethoxyethane is less dense than water. 1,1,1-Trichloroethane is denser than water.  
7. Use the tap to produce a sample of the purple liquid free from the other layer.

Questions

1. How could you obtain a sample of solid iodine from the solution you have made?  
2. How could you purify the solid iodine further?

Chromatography – the chemists’ most powerful separating tool!

Imagine that you have a mixture of cricket balls and tennis balls. You roll them down a slope which is slightly sticky, as if it was covered with sticky tape. The heavier cricket balls will roll better and so will become separated from the tennis balls.

We can do the same with chemicals. To see the process at work, try practical activity 2.5.
Chromatography means ‘writing with colours’. It was invented by a Russian botanist called Tswett, in 1906. He found two different substances in the green matter of plants.

**Figure 2.13** A sugar factory in Martinique.

**Practical activity 2.5 Chromatography**

**You need:**
- a strip of filter paper about 10 cm long
- 1 small beaker
- 1 larger beaker
- water-soluble blue or black ink
- pin
- water

**Method**

1. Pick up a tiny drop of ink on the end of a pin and spot it onto the filter paper, as shown in the diagram. The spot needs to be about 1 cm from the bottom edge of the filter paper.
2. Put a little clean water into the small beaker.
3. Stand the filter paper in the water. The water level must be below the spot of ink.
4. Cover the beaker with the larger beaker.
5. Wait until the water has soaked up to the top of the paper.
6. Take out the paper strip and look at the colours on it.

**Observations**

What did you see?
(The colours shown here might not be the ones you observe in your experiment.)

![Diagram of chromatography process](image)

**Why does chromatography separate substances?**

Just as with the balls on the sticky slope, chromatography involves two opposing forces. The dyes are soluble in water, so as the water rises up the filter paper it pulls the dyes along with it. But there is also water bound into the fibres from which the paper is made. This water also tries to dissolve the dyes, and so holds them back. Depending on the exact nature of each dye molecule, some travel easily and others are held back more.

**Sugar from cane – an industrial separation problem!**

Sugar cane contains sucrose in the cells of the canes. Sugar is a valuable product; extracting and purifying sugar is very important in the Caribbean.

The process used to extract the sugar uses the fact that sugar is very soluble in water.
- First the cane is shredded into chips by revolving knives and then crushed in a mill.
- In the mill the crushed cane goes one way and a current of water goes the other way – this is counter current washing. The water dissolves out the sugar and the result is called ‘dilute juice’.
- The dilute juice is cloudy and slightly acidic. To make it clear it and stop the sucrose changing into other sugars, it is heated (which kills enzymes that would make the sugar change). Also lime (calcium oxide) is added (often
dissolved in sucrose solution). This removes the acidity, raising the pH from about 6 to 7.

- The liquid is heated to 100 °C and pumped to settling tanks. At this stage the liquid is 10–15% sugar.
- The clear liquid is put into large pans and steam heated to drive off water. It is concentrated until it contains about 60% sucrose.
- The syrup is then concentrated even further under vacuum. To make it crystallise, some sugar crystals are added. About 50% of the sugar crystallises out. The liquid which is left is called molasses.
- A second crystallisation gives more sugar and ‘blackstrap’ (which cannot be made to crystallise).

### Some early ideas about atoms

About 1800, when chemists first began to believe that matter was made up of atoms, they pictured them as solid objects like billiard balls, with no structure (figure 2.14). The word ‘atom’ comes from the Greek ‘atos’ – ‘something which cannot be cut’.

Atoms could not be measured or weighed, but the relative mass of an atom, on a scale where hydrogen, the lightest atom, was 1, could be found. The relative atomic mass of an atom was defined as:

\[
	ext{relative atomic mass} = \frac{\text{mass of one atom of the element}}{\text{mass of one atom of hydrogen}}
\]

So an atom which has 23 times as much mass as an atom of hydrogen has a relative atomic mass of 23.

Values of \( M_r \) for elements found on the Earth go from 1 (hydrogen) to 238 (uranium).

Nowadays we use the isotope carbon-12 as the standard, instead of hydrogen. (See page 25 for more about isotopes.) One atom of this isotope of carbon has \( M_r = 12 \), so the modern definition of \( M_r \) is:

\[
	ext{relative atomic mass} = \frac{\text{mass of one atom of the element}}{\text{mass of one-twelfth of an atom of carbon-12}}
\]

Experiments with electrical machines showed that atoms were neither solid nor uniform throughout. They had a structure and contained electrical charges. Knowing this, an atom then was pictured like a fruit pudding, with electrical charges scattered here and there throughout it (figure 2.15).

By 1900, the discovery of radioactivity showed that some atoms fire out tiny fragments from within themselves (figure 2.16). Some of these fragments have positive electrical charges and some have negative charges.

In 1906, Rutherford had noticed that the positively charged particles from a radium atom did not always travel in straight lines. In 1909, Rutherford found that the deflections were obvious when the particles passed through very thin gold foil. Rutherford then set two of his colleagues to investigate just how large the deflections were, and on June 17 1909, Hans Geiger and Ernest Marsden announced their numerical results. Geiger and Marsden expected all the fragments to pass through the gold sheet. To their astonishment, some of them bounced back. Rutherford said ‘it was almost as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it come back and hit you!’
The nuclear atom

From the experiment, Rutherford, in 1911, was able to show that atoms have a quite different structure from the one shown in figure 2.16.

Rutherford showed that atoms:
- have a very tiny nucleus at the centre;
- have all their positive charge in the nucleus;
- have almost all their mass in the nucleus;
- have only negative electrical charges (which have almost no mass) making up the rest of their volume.

How large are atoms?

An average diameter for an atom is 100 pm so it would take ten million such atoms, laid side by side, to stretch just 1 millimetre.

The nucleus of an atom is unbelievably tiny.
- If an atom were magnified to size of a cricket field, the nucleus would be the size of the cricket ball.
- If a matchbox could be filled with bare nuclei, it would weigh about 10 000 tonnes.
- A nucleus has a radius about $\frac{1}{10 000}$ that of an atom.

Atoms are made of sub-atomic particles

We now know that all atoms are made up of the same three different types of particle, each with its own properties:
- neutrons contribute mass but no charge;
- protons contribute mass and positive charge;
- electrons contribute negative charge but their mass is so small that it is usually written as zero.

Only neutrons and protons are found in the nucleus. On the same scale as the proton mass, the electron mass is about $\frac{1}{1840}$. You can think of the electrons as tiny particles of electricity making huge circles around the nucleus of the atom. There are no electrons inside the nucleus.

<table>
<thead>
<tr>
<th>Name of particle</th>
<th>Symbol</th>
<th>Relative mass</th>
<th>Charge</th>
<th>Where found</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron</td>
<td>e</td>
<td>0</td>
<td>-1</td>
<td>always outside the nucleus</td>
</tr>
<tr>
<td>neutron</td>
<td>n</td>
<td>1</td>
<td>0</td>
<td>always in the nucleus</td>
</tr>
<tr>
<td>proton</td>
<td>p</td>
<td>1</td>
<td>+1</td>
<td>always in the nucleus</td>
</tr>
</tbody>
</table>

Table 2.3 Properties of the three sub-atomic particles.

Protons

An atom must have at least one proton in its nucleus. All the atoms of the same element have the same number of protons. The number of protons in the nucleus of an atom is called the proton number or the atomic number.

The chemistry of an element is determined by the number and arrangement of its electrons. This, in turn, is set by the number of protons in its nucleus because an atom is electrically neutral and has the same number of electrons as it has protons. So the chemical behaviour and chemical identity of an element is fixed by its proton number (atomic number).

Neutrons

A nucleus made up only of protons would fly apart because positive charges repel each other. The force holding the protons together is enormously strong and it depends on there being neutrons in the nucleus, as well as protons.

The number of neutrons in the nucleus is the neutron number $N$. 

1 pm = $\frac{1}{1000000000000}$ m ($1 \times 10^{-12}$ m)
Electrons

An electron is the smallest electrical charge which can exist anywhere. Electrons have a negative charge. Although electrons are really waves (like light rays and X-rays), we can also think of them as little particles of electricity and show them in diagrams as a dot or a cross (● or ×). We may think of electrons as rushing around the nucleus in set orbits, like planets around a sun, but they are going so fast that we cannot make out just where they are at any one moment.

Mass number

Protons and neutrons together make up all the mass of an atom. The proton and the neutron each have a relative mass of 1. The sum of the number of protons and the number of neutrons in a nucleus gives the relative mass of the nucleus as a whole number.

The mass number of an element is the sum of the number of protons and the number of neutrons in one atom of that element.

Isotopes

Although all atoms of a particular element must have the same number of protons, they may contain different numbers of neutrons. So there can be atoms of the same element which have different relative molecular masses. These are called isotopes of that element. All the isotopes of the same element have the same chemical properties. We can use carbon as an example (see table 2.4).

<table>
<thead>
<tr>
<th>Proton number</th>
<th>Neutron number</th>
<th>Mass number</th>
<th>Isotope</th>
</tr>
</thead>
<tbody>
<tr>
<td>6</td>
<td>6</td>
<td>12</td>
<td>C-12</td>
</tr>
<tr>
<td>6</td>
<td>8</td>
<td>14</td>
<td>C-14</td>
</tr>
</tbody>
</table>

Table 2.4 Isotopes of carbon.

Radioactivity

If there are the ‘wrong’ number of neutrons in a nucleus, the atom changes to get the number right. The atom fires out one or more fragments from its nucleus. This type of change is called radioactivity.

We shall learn more about radioactivity and isotopes in Chapter 3.

Electron shells

The electrons are not moving at random in the space around the nucleus. They are arranged in layers, one outside the other. These layers of electrons are called electron shells.

A shell can only hold up to a maximum number of electrons. You will only be concerned with the first four shells. These can hold:

- first shell 2 electrons
- second shell 8 electrons
- third shell 8 electrons
- fourth shell 18 electrons (but only the first 2 of the 18 concern you!)

Figure 2.17 Electrons in the first shell around a helium atom.

Figure 2.18 An atom showing its nucleus and three electron shells.
Filling shells
Putting the electrons into the shells is like climbing the rungs of a ladder. You start at the nucleus and add electrons to the bottom shell until it is full. Then you begin on the second shell, and so-on. The last shell may or may not be full. Lower shells must all be full.

Shell diagrams
We can use simple diagrams to show how the electrons of an atom are fitted into the shells. Figure 2.19 shows 9 of the first 20 atoms.

Figure 2.19 Shell diagrams for some of the first 20 atoms.
Symbols for atoms

To show how atoms combine together, chemists need to draw diagrams. To do this each different atom must have a symbol. At first the symbols which were used were circles with patterns inside, but (as you can see from figure 2.20) it is not obvious what a symbol represents. Eventually, chemists agreed on a system of symbols that uses letters to represent atoms.

Figure 2.20  The original symbols for carbon, hydrogen, oxygen and sulphur.

The symbols we use now are one capital letter either on its own (as in C for carbon) or followed by a lower-case (small) letter (as in Zn for zinc). The symbols for most elements make it easy to see what the name of the element is. However, those elements that have been known about since the earliest times have symbols derived from their Latin names. So iron (called ferrum by the Romans) becomes Fe, sodium (natrium) becomes Na and lead (plumbum) becomes Pb. There is a list of elements and their symbols on page 402.

When writing formulae and equations we need to know some or all of these facts:
• which element?
• how many atoms of the element?
• what is the relative atomic mass of the atoms?
• what is the charge on each atom?

To fit all this in, imagine the symbol for an atom written at the centre of a square. The other pieces of data go at the corners of the square, as shown in figure 2.21.

Here is an example: $^{17}_3\text{Cl}^0$

This means one atom of chlorine with charge zero, mass number 35 and proton number 17.

Using the proton number together with the symbol is not really necessary, because (using the same example) all atoms with $Z = 17$ are atoms of chlorine.

We always write the letters but we only write the numbers which are important to the job in hand. Often the charge and the number of atoms are sufficient.

Atoms and ions

Atoms are electrically neutral. Removing an electron from an atom leaves the atom with a net positive charge. Adding an extra electron to an atom gives it a net negative charge. Charged atoms or groups of atoms are called ions.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Proton number</th>
<th>Number of electrons</th>
<th>Net charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl</td>
<td>17</td>
<td>17</td>
<td>0</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>17</td>
<td>18</td>
<td>-1</td>
</tr>
<tr>
<td>Fe</td>
<td>26</td>
<td>26</td>
<td>0</td>
</tr>
<tr>
<td>Fe²⁺</td>
<td>26</td>
<td>24</td>
<td>+2</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>8</td>
<td>10</td>
<td>-2</td>
</tr>
<tr>
<td>O²⁻</td>
<td>8</td>
<td>10</td>
<td>-2</td>
</tr>
<tr>
<td>Na⁺</td>
<td>11</td>
<td>11</td>
<td>+1</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>11</td>
<td>10</td>
<td>-2</td>
</tr>
</tbody>
</table>

Table 2.5  Some names and symbols.

Be warned; never use symbols as mere shorthand. ‘H’ does not mean ‘hydrogen’ – it means one atom of hydrogen or one mole of hydrogen.

Note that if the charge is 0 we do not bother to write it and if the number of atoms is 1 we do not bother to write it. So, S means ‘one atom of sulphur with a charge of zero’.

ITQ10 What is meant by each of these symbols?
(i) $^{12}\text{C}$  (iv) $\text{Na}^-$  (vii) $\text{NO}_3^-$
(ii) $\text{O}_2$  (v) $\text{H}^+$
(iii) $\text{Ca}^{2+}$  (vi) $\text{H}_2^+$

‘Net’ means ‘overall’, with all the charges, both positive and negative, taken into account.

Table 2.6  Symbols for some common atoms and ions.
Ions with a positive charge are called cations (you say ‘kat-ions’) and ions with a negative charge are called anions. You will see why when you study electrolysis in Chapter 11.

Table 2.6 (on the previous page) shows some examples of common atoms and ions.

Summary
- Any material is either an element, a compound or a mixture.
- The relative atomic mass of an atom is measured using an atom of carbon-12 as standard.
- Atoms consist of a nucleus surrounded by negatively charged particles.
  - The nucleus contains all the mass and all the positive charge of the atom.
  - Atoms are made up from protons, neutrons and electrons.
  - The electrons are arranged around the nucleus in shells.
  - Each shell has its own maximum capacity for electrons.
- The proton number of an atom is the number of protons in the nucleus of that atom. This can be called the atomic number of the element \( Z \).
- The mass number of an atom \( N \) is the sum of the number of protons and the number of neutrons in the atom.
- A neutral atom contains equal numbers of protons and electrons.
- Atoms with the same proton number but different mass numbers are called isotopes.
- Atoms can be represented by a symbol consisting of either one capital letter or one capital letter followed by one lower case letter.
- Extra symbols can be added to represent the atomic properties \( Z, N \) and electric charge.

Answers to ITQs

ITQ1 The pressure is reduced, so the gas comes out of solution.

ITQ2 Temperature rises to 80 °C, stays at 80 °C until all that liquid has distilled, then rises to 100 °C.

ITQ3 To expose as much of the sugar-bearing structure as possible.

ITQ4 Bare nuclei, without electrons, would repel each other, because they all have the same charge, and so would fly apart.

ITQ5 sum of masses = 1 + 0 = 1
sum of charges = 1 + (–1) = 0
the particle with mass 1 and charge 0 is a neutron

ITQ6 11 + 12 = 23

ITQ7 Atoms are electrically neutral. The charges on a proton and an electron are equal in size but opposite in charge.

ITQ8 3

ITQ9 There are only 12 electrons instead of 13.
There are 7 electrons in the second shell. There should be 8.

ITQ10 (i) One atom of carbon with mass number 12.
(ii) A particle consisting of two atoms of oxygen. (If you said a molecule containing two atoms of oxygen you were also correct.)
(iii) An atom of calcium, with a charge of +2. (If you said it was an ion you were also correct.)
(iv) An atom of sodium, atomic number 11, with a charge of +1. (If you said it was an ion you were also correct.)
(v) An atom of hydrogen with a mass number of 2.
(vi) A particle containing two atoms of hydrogen, with an overall charge of +1. (If you said it was an ion you were also correct.)
(vii) A particle containing one atom of nitrogen and three atoms of oxygen; with an overall charge of –1. (Again you may have called the particle an ion.)
1 (i) Describe how you could produce pure samples of iron and sulphur powders from a mixture of the two.

(ii) A student obtained these results for the solubility of a substance X at various temperatures.

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>10</th>
<th>20</th>
<th>30</th>
<th>40</th>
<th>50</th>
<th>70</th>
<th>80</th>
<th>90</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solubility (g/100 g)</td>
<td>7</td>
<td>9</td>
<td>12</td>
<td>15</td>
<td>19</td>
<td>30</td>
<td>38</td>
<td>50</td>
</tr>
</tbody>
</table>

(a) Draw a graph of these figures.
(b) Estimate the solubility of X at 60 °C.
(c) 30 g of X were dissolved in 100 cm³ of water at 80 °C. The solution was then cooled to 30 °C. What mass of X would remain in solution?

2 (i) Explain, with the aid of a diagram, what is meant by fractional distillation. When is it better to use this process rather than simple distillation?

(ii) An aqueous solution contains two solutes, A and B. A is soluble in hexane, B is not. Hexane is liquid at room temperature and does not mix with water. The density of hexane at room temperature is 0.7 g dm⁻³.

Describe, giving full practical details, how you would produce a pure sample of solid A.

3 (i) List three physical properties of sucrose (cane sugar).

(ii) Explain how sucrose is removed from raw sugar cane.

(iii) A sample of sucrose has become contaminated with sand. How would you purify the sucrose?

4 (i) Describe Rutherford’s experiment which showed that atoms have a nuclear structure.

(ii) Atoms contain electrons (e), neutrons (n) and protons (p).

(a) Draw a diagram showing where these particles are found in an atom.

(b) Draw a table listing these particles, their relative masses and their relative electrical charges.

(iii) Explain what is meant by the statement that the electron structure of a sodium atom is 2,8,1.

(iv) An atom has the symbol \(^{40}_{18}\text{Ar}\).

(a) How many neutrons does the atom contain?

(b) What is the relative atomic mass (Mr) of the atom?

(c) Explain how you arrive at your answers.

5 (i) Write a symbol representing:

(a) an atom with atomic number 2 and mass number 4;

(b) an ion of sulphur containing 18 electrons.

(ii) Draw a diagram showing the electron arrangement in an atom of magnesium.

(You will need to use the Periodic Table on page 401 to gather information for these answers.)

6 (i) Draw a shell diagram showing the electron arrangement in an atom of silicon. Silicon has atomic number Z = 14.

(ii) (a) Using the electron arrangements of silicon and of carbon, explain why the two elements are expected to be alike chemically.

(b) Is the chemistry of silicon similar to the chemistry of carbon?
Radioactive decay

In Chapter 2 we saw that atoms have nuclei containing protons and neutrons and that the nuclei are surrounded by electrons. If the balance between protons and neutrons in the nucleus is not right, the nucleus is unstable. An unstable nucleus changes by ejecting small fragments. This process is called radioactivity. The changing radioactive atoms are said to decay.

Fragments which can be thrown out include:
- neutrons – this is rare, but vital for nuclear power;
- $\alpha$-particles – bundles of two neutrons and two protons;
- $\beta$-particles – which are electrons;
- $\gamma$-rays – which are electromagnetic rays like X-rays.

When an atom emits $\alpha$- or $\beta$-particles, the number of protons in the nucleus changes, so the chemical identity of the atom changes. Losing $\gamma$-rays only reduces the energy of the nucleus.

Half-life

Each radioactive isotope decays at its own fixed rate. This rate is different for different isotopes, but the time taken for one-half of the atoms of any one isotope to decay is constant. So if 100 000 atoms decay to 50 000 atoms in a time $t$ seconds, then the 50 000 will decay to 25 000 in the same period of time. This period of time is called the half-life of that isotope. Half-lives can vary from tiny fractions of a second to hundreds or thousands of years.
Penetrating power

α- and β-particles and γ-rays have different powers of penetration, and these determine the uses to which they are put.

- α-Particles are stopped by a sheet of paper or human skin.
- β-Particles will penetrate a thin sheet of metal.
- γ-Rays may penetrate several inches of lead.

Uses of radioactivity

Carbon dating

One of the isotopes of carbon, carbon-14, is radioactive. It emits β-particles and has a half-life of 5600 years. Carbon-14 exists in small amounts in nature. When a plant absorbs carbon dioxide during photosynthesis, it absorbs this radioactive isotope of carbon also and uses it to make starch and cellulose. When the plant dies, it stops absorbing carbon dioxide. The carbon-14 within it decays with its usual half-life. By measuring how much carbon-14 is left in plant residues, and comparing this with the content of a living plant, we can calculate how long ago the plant died.

Fragments of wood have been recovered from the holes that hold the huge stones that make up Stonehenge, in England. Carbon dating of these fragments has shown that the stones were put in place 4000 years ago. Other examples of carbon dating have shown that humans first appeared in the Caribbean about 4000 years ago. Carbon dating can be used for samples aged up to 40 000 years old.

Use as a tracer

Radioactive atoms of a particular element have the same chemistry as non-radioactive atoms of that element. Because we can detect tiny numbers of radioactive atoms with a Geiger counter or a photographic film, we can follow an atom through a chemical reaction to trace the path it takes.

Figure 3.4 (on the next page) shows a pelargonium leaf which has been grown in air containing extra amounts of carbon-14 (as carbon dioxide). The leaf was then removed from the plant, placed on a photographic film, and left for several hours. The β-particles from the decaying carbon-14 affected the film just as light would have done. So we can see where the carbon-14 has been taken up and how it is being transported along the leaf veins. The white portions are from the parts of the leaf containing the C-14.
Use in radiotherapy

When an α- or β-particle, or a γ-ray, enters a human body the energy which it loses can damage dividing cells. The cells of a cancer are dividing more frequently than those of the surrounding tissue, so they are affected more by radioactivity. Some forms of cancer can be treated by subjecting them to radiation. For example, cobalt-60 decays giving penetrating γ-rays, which can be focused onto a cancer inside a human body.

Because the ray is focused onto the tumour the surrounding healthy tissue receives a smaller dose of radiation, which leaves it relatively unharmed.

As a source of energy

Large-scale sources of energy

There are some atoms whose nuclei can split into two roughly equally sized pieces. When this happens something very unexpected takes place. A tiny amount of the mass of the atomic nucleus is turned into energy. Although the mass loss is tiny, the amount of energy released is very large.

The energy appears as heat, the heat is used to boil water to make steam, then the steam is used to drive turbines which turn electrical generators. The electricity is used to power our factories and homes.
The commonest atoms which behave in this way are atoms of the isotope uranium-235. U-235 is a nuclear fuel. The process is called nuclear fission (fission means ‘splitting’). Uranium atoms do not split unless the split is triggered. A very few of these uranium atoms give off a neutron, which is the trigger for fission. Each fission gives off more than one neutron, and each of these neutrons can trigger another atom to split. So in a large stack of atoms, called a nuclear pile, the process builds up until useful amounts of energy are being produced. This is called a chain reaction.

**Small-scale sources of energy**

α-Particles lose their energy in only a few collisions with other atoms. This is why they are stopped by very thin layers of material, but it makes them very valuable as sources of energy. In a smoke detector a tiny amount of the radioactive element americium-242 is placed in a gap between two electrical contacts. The α-particles from the americium make the air in the gap electrically conducting, so that a tiny electrical current flows. If even a trace of smoke enters the detector it prevents the current from flowing and triggers a very loud alarm, giving warning of the smoke. These devices have saved many lives by alerting people to the presence of a fire in their homes.

**Figure 3.8** The build-up of a nuclear chain reaction (fission fragments are not shown).

The energy given off during radioactive decay can be used to generate electricity. Microminiature batteries generating only nanowatts of power are made from the isotope nickel-63, which emits β-particles (electrons). The electrons can be captured by a semiconductor system and used directly. Since this isotope of nickel has a half-life of over 100 years, the battery will wear out before it runs out! Such batteries can be used in devices such as heart pacemakers because the tiny amount of radioactivity which they contain cannot harm the person who has the device fitted.

Much larger batteries, generating hundreds of watts, have been used in space flights. In one design, the heat given off during the decay of plutonium-238 is captured by a sheath of thermocouples and converted to electrical energy. The plutonium gives off no γ-rays and, inside its casing, is safe to handle. However, it gets so hot that you need thick gloves to protect your hands.

---

‘Nano-’ means $10^{-9}$, so one nanowatt (1 nW) is one-thousandth of a microwatt.

NASA calls these SNAP batteries, standing for ‘Systems for Nuclear Auxiliary Power’.

**Figure 3.9** A domestic smoke detector.

**Figure 3.10** A heart pacemaker, powered by radioactivity.
Sterilisation

Large doses of $\gamma$-rays interfere so powerfully with living systems that they cause death. Surgical instruments can be sterilised by passing them repeatedly in front of a large cobalt-60 source. Foodstuffs also can be treated in this way. For example, flour can be irradiated to get rid of weevils, and soft fruit can be irradiated to kill mould spores.

Figure 3.11 These prawns have been irradiated to kill any harmful bacteria that may be present.

Summary

- Radioactive decay is the process by which a nucleus changes its balance of neutrons and protons.
- In radioactive decay, alpha particles, beta particles or gamma rays may be lost.
- Radioactivity can be used:
  - as a chemical tracer;
  - to date ancient objects;
  - to produce electricity on a large scale in nuclear piles;
  - to produce electricity on a smaller scale in cells and batteries;
  - in radiotherapy;
  - in smoke alarms;
  - to sterilise materials.

Answers to ITQs

ITQ1 In the next half-life 2000 atoms will decay. In the next half-life 1000 atoms will decay. So 3000 atoms will decay in two half-lives, which is 30 minutes.

ITQ2 A neutron can be thought of as a proton + an electron:

$$^{1}p^{+} + ^{0}e^{-} \rightarrow ^{1}n^{0}$$

The nucleus converts a neutron into a proton and an electron. The electron is lost. The nucleus now has an extra proton, so it is an atom of a different element.

ITQ3 After one half-life the sample would give 50 $\beta$-particles per minute and after another half-life the rate would fall to 25. So the sample was two half-lives old.

$$2 \times 5600 \text{ years} = 11 \text{ 200 years}$$

ITQ4 Photosynthesis.
Examination-style questions

1 (i) Give a definition of an isotope.

(ii) Which of the following particles are isotopes of each other? (In each case the symbol of an element has been replaced by X.)
   (a) \(^{35}\)X
   (b) \(^{36}\)X
   (c) \(^{39}\)X
   (d) \(^{37}\)X

(iii) State three possible uses for radioactive isotopes and explain one of those uses in as much detail as you can.

2 (i) Explain what is meant by the statement that a substance is radioactive.

(ii) The apparatus shown in the diagram contains a radioactive substance, a Geiger counter and thin sheet of lead. Without the lead in place (diagram A) the Geiger counter detected 6000 counts in 1 min. With the lead in place (diagram B) the Geiger counter detected 3000 counts in 1 min.

State and explain what types of radiation were being emitted by the source.

3 (i) What is meant by the half-life of a radioactive material?

(ii) In an experiment a Geiger counter was placed at a fixed distance from a radioactive source. The count displayed on the counter was recorded for 1 min. The counter was then reset to 0 and allowed to run for another minute. This was repeated until 10 readings had been taken (shown in the table).

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>Count</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>8000</td>
</tr>
<tr>
<td>2</td>
<td>6165</td>
</tr>
<tr>
<td>3</td>
<td>5010</td>
</tr>
<tr>
<td>4</td>
<td>4010</td>
</tr>
<tr>
<td>5</td>
<td>3090</td>
</tr>
<tr>
<td>6</td>
<td>2430</td>
</tr>
<tr>
<td>7</td>
<td>1950</td>
</tr>
<tr>
<td>8</td>
<td>1550</td>
</tr>
<tr>
<td>9</td>
<td>1260</td>
</tr>
<tr>
<td>10</td>
<td>1000</td>
</tr>
</tbody>
</table>

Plot a graph of these results and use your graph to find the half-life of the source.
By the end of this chapter you should be able to:

- write and interpret symbols for common ions and molecules;
- understand and use the subscripts and superscripts that are added to atom symbols;
- explain the meaning of the mole as a chemical unit;
- state the value of the Avogadro constant;
- construct and balance molecular chemical equations;
- construct and balance ionic chemical equations;
- use and understand state symbols in chemical equations;
- recognise spectator ions in an ionic reaction;
- carry out calculations on reacting quantities for a given chemical reaction;
- carry out calculations involving volumes of gases involved in a chemical reaction.

Quantities, moles and equations

The importance of measurements

Chemistry has been studied in one way or another for at least three thousand years. In that time, workers paid a lot of attention to observations, but very little to measurements. The art of chemistry grew but the science of chemistry was not yet born. By 1600 AD rough measurements were made. For example, people knew that if they heated a metal in air its ‘weight’ went up, but no-one tried to use measurements to form laws or make theories.

Around 1800 AD everything changed. Balances became much more reliable and people began to use their measurements of ‘weight’ and volume to try to
explain how matter is made up and how it behaves. The laws of combination, which give good evidence for the atomic theory, were worked out. The structure of atoms and chemical bonds became understood in more and more detail.

It is only in the last 70 years or so that our skills and technology have made it possible to make measurements of incredible precision and to look into the worlds smaller than the size of a single atom.

Chemical measurements
As chemists we need to make a range of measurements:
- mass and volume (from these we can work out concentration);
- numbers of particles;
- size and shape (from these, we can make 3D models of the atomic world).
In this chapter we will mainly look at mass, volume and concentration. We will not look at size and shape at all.

Chemical symbols
Definitions

| element | An element is a substance which cannot be broken down into simpler substances by chemical means. Carbon, chlorine, iron, and uranium examples of are elements. |
| atom | An atom is the smallest particle of an element which has the properties of that element. |
| compound | A compound is a substance composed of two or more elements combined in a fixed proportion by mass. Iron chloride, water, alcohol and methane are examples of compounds. |
| molecule | A molecule is the smallest particle of a compound which has the properties of that compound. |
| ion | An ion is an electrically charged atom or group of atoms. |

Chemical language
Like any other special subject, chemistry has its own language. It is based on our belief that all the atoms of a single element are alike but are different from the atoms of any other element. Because of this we can have a single symbol for ‘one atom of the element’.

For example, C means ‘one atom of carbon’ (not just ‘carbon’). If this atom of carbon then forms a bond to one atom of oxygen (O) we can write the formula of the particle formed as CO. This formula tells us that the molecule we have made has one atom of carbon combined with one atom of oxygen. We put the carbon atom on the left and the oxygen atom on the right because carbon is to the left of oxygen in the Periodic Table.

If more than one atom is involved (for example, when one atom of oxygen combines with two atoms of hydrogen, we could write either H₂O or H–O–H.
- H₂O is a molecular formula. It tells us which elements are in the compound and in what numbers.
- H–O–H is a structural formula. It gives the same information as the molecular formula but, in addition, shows which atoms are bonded together. In this case the oxygen atom is bonded to each of the hydrogen atoms, but the hydrogen atoms are not bonded to each other.

Unless there is a good reason for doing otherwise, we use molecular formulae. They are shorter!

**ITQ1**

When we write FeO what do we mean?

**ITQ2**

The structural formula for ammonia is \( \text{H} – \text{N} – \text{H} \)

(i) Write the molecular formula for ammonia.

(ii) Are the hydrogen atoms in an ammonia molecule bonded to each other?
Adding other information

To describe a chemical reaction we sometimes need two other pieces of information about the reacting particles. We may need to know about:

- any electrical charges;
- the physical states;
- the number of particles in a molecule.

If a particle has a charge, it is written above and to the right of the symbol. For example, O\(^{2-}\) means a single oxygen atom carrying a negative charge of two units. The unit we use is the same as the charge on an electron, so this atom has gained two electrons. Positive charges are written in the same way; Fe\(^{3+}\) represents a single atom of iron which carries a positive charge of three units (the atom has lost three electrons).

The chemicals we use are either solids, liquids, gases or are in aqueous solution (dissolved in water). We show the state as an abbreviation in brackets after the symbol, with no space between the two. The symbols are:

- solid (s);
- liquid (l);
- gas (g);
- aqueous solution (aq).

If the particle is a molecule, the number of atoms of a particular element in it is written to the right of the element symbol, as a subscript.

An atom or molecule which has a net charge is called an ion:

- a negative ion is an anion;
- a positive ion is a cation (pronounced ‘kat-ion’ not ‘cashun’).

Na\(^+\), Mg\(^{2+}\), S\(^2-\), NO\(_3\)^–, SO\(_4\)^{2+}\) and CH\(_3\)COO\(^–\) are examples of ions. These rules are summarised in figures 4.2 and 4.3.

**Figure 4.2** Adding numbers to a chemical symbol.

<table>
<thead>
<tr>
<th>X</th>
<th>charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>number of particles</td>
<td></td>
</tr>
</tbody>
</table>

**Cr\(^{2+}\)(s)**

*Figure 4.3* The symbol for one ion of the element chromium, with two positive charges, in the solid state.

**Law of conservation of mass**

The law of conservation of mass says that matter can neither be created nor destroyed in a chemical reaction.

Without the conservation of mass, even the best measurements would be useless. In ordinary chemical reactions atoms keep their own identity and so the law tells us that there must be the same number of atoms of each element after the reaction as there were before it.

**Concentration**

When we speak about ‘a strong solution’ or ‘a dilute solution’ we are giving only a rough idea of the strength of that solution. We can do more accurate work when we work with the exact concentration of the solution. For example, we may say that a solution of salt contains 5.85 g of salt per litre of solution. Remember that a solution is made from a solvent and a solute;

- the solvent is the substance (usually a liquid) present in the larger quantity;
- the solute is the substance which becomes dispersed in the solvent; usually there is less solute than solvent.

The mass concentration of a solution is the mass (in grams) of solute present in 1 litre of solution. Mass concentration is measured in units of g dm\(^{-3}\).
The molar concentration of a solution is the number of moles of the substance in 1 litre of solution. Molar concentration is measured in units of mol dm$^{-3}$ (see page 44).

If we know the mass concentration of a solution and its volume, we can calculate how much solute the solution contains:

$$\text{mass concentration} = \frac{\text{mass}}{\text{volume}}$$

$$\text{mass} = \text{mass concentration} \times \text{volume}$$

### Standard solutions

Some solutes are dry, crystalline solids. They can be weighed out accurately. Sodium chloride and anhydrous sodium carbonate are examples. Others cannot be weighed exactly, either because they react with the air, or they become damp as you weigh them. Sodium hydroxide is an example. In air it becomes wet and forms sodium carbonate by reacting with carbon dioxide. A solution made from sodium hydroxide pellets will not contain quite as much sodium hydroxide as the balance suggested!

A solution of accurately known concentration is a standard solution.

### Titration

The concentration of a solution can be found by reacting it with a known volume of a standard solution until the reaction is complete. This process is called titration.

The volume of standard solution is measured using a pipette.

- The pipette is filled with liquid to a level above an engraved mark on the stem of the pipette. The liquid is allowed to drain out (into a beaker to catch the waste liquid) until the meniscus reaches the engraved mark on the stem.
- The pipette is then allowed to deliver liquid until no more will drain out. This is the liquid you will use in your titration. Never blow out the last drop from a pipette; some is meant to remain when draining is complete.

The unknown solution is added from a burette, which is a long graduated tube fitted with a tap at one end.

To find an unknown concentration we need to know:

- the value of $M_r$ (relative molecular mass) for the standard solution;
- the volume of the standard solution;
- the concentration of the standard solution;
- the volume of unknown solution needed for an exact reaction;
- the equation for the reaction between the two solutions.

### Titration calculations

When carrying out a titration calculation, follow these three steps:

1. calculate the number of moles of standard substance in the known volume of liquid delivered by the pipette;
2. calculate how many moles of the test substance are needed for complete reaction;
3. calculate the concentration of the test solution from the volume of it used in the titration.
Worked example 4.1

A solution of hydrochloric acid was titrated against 25 cm$^3$ of a standard solution of sodium carbonate of concentration 0.10 mol dm$^{-3}$. 20 cm$^3$ of acid were needed for complete reaction.

1. 1.0 dm$^3$ of sodium carbonate solution contains 0.10 mol
   25 cm$^3$ of the solution contains $0.10 \times \frac{25}{1000} = 2.5 \times 10^{-3}$ mol

2. The reaction is:
   $$\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$$
   This shows that each mole of sodium carbonate reacts with two moles of hydrochloric acid.
   $2.5 \times 10^{-3}$ mol react with $5.0 \times 10^{-3}$ mol of acid

3. 20 cm$^3$ of acid contained $5.0 \times 10^{-3}$ mol
   1 cm$^3$ of acid contained $\frac{5.0 \times 10^{-3}}{20}$ mol
   1 dm$^3$ of acid contained $\frac{5.0 \times 10^{-3}}{20} \times 1000$ mol = 0.25 mol
   The concentration of the acid is 0.25 mol dm$^{-3}$

Worked example 4.2

A solution of hydrochloric acid is known to have a concentration of 0.70 mol dm$^{-3}$.
25 cm$^3$ of a solution of a base reacted exactly with 35 cm$^3$ of the acid.
The base had a concentration of 1.0 mol dm$^{-3}$.
If the reaction between the two is according to this equation:
   $$\text{HCl} + x\text{(base)} \rightarrow \text{salt} + \text{water}$$
then find the value of $x$ in the equation.

1. 35 cm$^3$ of acid contained $\frac{35}{1000} \times 0.70$ mol = 0.0245 mol

2. The base solution contained $\frac{25}{1000}$ = 0.025 mol

3. Therefore 0.0245 mol of acid reacted with 0.025 mol of base
   1 mol acid reacted with 1 mol base
   $x = 1$
   Note that although 0.0245 is not exactly equal to 0.025, the answer is 1 because the value of $x$ must be a whole number.

Indicators

Some compounds change colour according to the acidity (pH) of their surroundings. You might find that this works for coloured flowers such as bougainvillea or hibiscus.

We use compounds like this to see when a reaction is complete – the end-point of the reaction. When doing a titration, add just enough of the indicator to give an obvious colour to the solution. After each addition of the second solution, swirl the liquid in the beaker or flask to make it uniform. Where the liquids are not mixed you will see more than one colour. When you are near the end-point the colour will ‘hang’ – it will not become uniformly so quickly. Then you know to add the liquid from the burette one drop at a time, swirling the liquid after each drop, until all the colour finally changes.
Different indicators change colour at different values of pH; the chart in figure 4.5 shows some common examples.

You need:
• sodium hydroxide solution, roughly 0.1 mol dm⁻³
• hydrochloric acid, accurately 0.10 mol dm⁻³
• universal indicator liquid
• conical flask (250 cm³)
• 25 cm³ pipette (or a 25 cm³ measuring cylinder)
• burette

Method
1 Pipette 25 cm³ of sodium hydroxide into the flask. Your teacher will show you how to do this accurately. Remember: do not blow out the drop which stays in the end of the pipette!
2 Fill the burette to just below the ‘0’ mark with the hydrochloric acid.
3 Note the reading on the burette – take the reading at the bottom of the meniscus as shown in the diagram. This is \( V_1 \).
4 Add sufficient indicator to the liquid to make it coloured. One or two drops is usually enough.
5 Add the acid a little at a time to the solution in the flask. After each addition swirl the flask and note the colour.
6 When the colour gets near to the colour which represents ‘neutral’ (pH 7) add the acid one drop at a time until you judge that the pH is exactly 7.
7 Take the reading of the liquid level in the burette. This is \( V_2 \).

Using your results
1 Use \( V_1 \) and \( V_2 \) to calculate what volume of acid you added. (This is why you did not have to begin with the acid on the zero mark!)
2 The acid had a concentration of 0.10 mol dm⁻³. How many moles of acid did you add?
3 The equation for the reaction is:

\[
\text{HCl(aq) + NaOH(aq)} \rightarrow \text{NaCl(aq) + H}_2\text{O(l)}
\]

How many moles of alkali react with each mole of acid?
4 How many moles of alkali were in the volume which you added?
5 How many moles of alkali would there be in 1 cm³ of the alkali?
6 How many moles of alkali would there be in 1 dm³ of the alkali?
7 What is the concentration of the alkali (in g dm⁻³)?

You should quote your result to 2 significant figures.
Equations

Reacting particles in a chemical change may be atoms, molecules, ions or electrons. A chemical equation sets down what particles are present at the beginning (the reactants) and what particles are present at the end (the products). The equation tells us the relative numbers of each particle and their physical state. For example:

\[ \text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \]

This means ‘One atom of solid carbon reacts with one molecule of oxygen gas to make one molecule of carbon dioxide gas’.

Notice the arrow ‘\( \rightarrow \)’ between the reactants and the products. This means ‘produces’ or ‘changes into’. In older books you will find an equals sign ‘\( = \)’ instead of the arrow. Its meaning is exactly the same.

To check that the equation does not break law of conservation of mass, check that the number of atoms of each element is the same on both sides of the equation.

In our equation:

\[
\begin{array}{ccc}
\text{carbon} & 1 & \text{atom} \\
\text{oxygen} & 2 & \text{atoms}
\end{array}
\]

As these numbers do check, we can say that the equation is balanced.

**Worked example 4.3**

The following equation is not balanced.

\[ \text{H}_2\text{S(g)} + \text{SO}_2(\text{g}) \rightarrow \text{S(s)} + \text{H}_2\text{O(l)} \]

(i) Explain why it is not balanced.

(ii) Write the balanced equation for the reaction.

(iii) State the meaning of the balanced equation in words.

**Solution:**

(i) There are two atoms of sulphur on the left-hand side and only one on the right. There are two atoms of oxygen on the left-hand side and only one on the right.

(ii) \[ 2\text{H}_2\text{S(g)} + \text{SO}_2(\text{g}) \rightarrow 3\text{S(s)} + 2\text{H}_2\text{O(l)} \]

(iii) Two molecules of hydrogen sulphide gas react with one molecule of sulphur dioxide gas to give three atoms of solid sulphur and two molecules of liquid water.

**Numbers of atoms**

An equation tells us the relative numbers of atoms and molecules taking part in a reaction. But how many particles are involved? Clearly, since atoms are so tiny, the numbers involved must be immense. We count eggs in dozens, but we count chemical particles in units of \(6 \times 10^{23}\). This number of particles is called one mole (1 mol). This may seem a strange number, but there is a reason for it. It is chosen because \(6 \times 10^{23}\) hydrogen atoms make up 1 gram of hydrogen.

\(6 \times 10^{23}\) is called the Avogadro constant. It is named after Amedeo Avogadro (1776–1856), who persuaded chemists that elements like hydrogen or oxygen could exist as particles which contained more than one atom. He called the particles molecules.

More accurately, \(6.02 \times 10^{23}\) particles = 1 mol. It can be useful to know that one mole of a solid element is roughly a small handful!

The mole is the amount of substance which contains \(6.0 \times 10^{23}\) (the Avogadro constant) particles of that substance.

The symbol for the Avogadro constant is \(L\). Its unit is mol\(^{-1}\) because there are \(L\) particles per mole. All the particles in a mole must be identical. You can talk about a mole of atoms, molecules, ions or electrons.
Look again at the equation for the reaction of carbon and oxygen:
\[ C(s) + O_2(g) \rightarrow CO_2(g) \]
We can now read this as ‘One mole of solid carbon reacts with one mole of oxygen gas to make one mole of carbon dioxide gas’.

Here is an example containing multiple moles. The equation shows the reaction between hydrogen and oxygen:
\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \]
This means that ‘Two moles of hydrogen gas react with one mole of oxygen gas to produce two moles of liquid water’.

### Moles and volumes

A mole of a solid is roughly a small handful. For a gas we can be more precise. Avogadro (yes, him again!) discovered that one mole of any gas, under the same conditions of temperature and pressure, occupies the same volume.

The volume of one mole of a gas is called its molar volume, \( V_m \). We measure the volumes of gases either at 0 °C and 101 kPa (standard temperature and pressure, STP) when \( V_m = 22.4 \text{ dm}^3 \) or at 25 °C and one atmosphere (room temperature and pressure, RTP) when \( V_m = 24 \text{ dm}^3 \).

#### Worked example 4.4

96 litres of hydrogen are made to react with oxygen. What volume of oxygen will be needed? All volumes are measured at RTP.

The equation for the reaction is:
\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \]

**Solution:**

The equation shows that 2 mol of hydrogen molecules react with 1 mol of oxygen molecules.

At RTP 1 mol of a gas occupies 24 dm³.

- So the 96 dm³ of hydrogen must contain \( \frac{96}{24} = 4 \text{ mol} \)
- We need \( \frac{4}{2} = 2 \text{ mol} \) of oxygen
- 2 mol of oxygen at RTP occupies \( 2 \times 24 \text{ dm}^3 = 48 \text{ dm}^3 \)

#### Worked example 4.5

If, at RTP, 1 mol of sulphur reacts with 1 mol of oxygen, what volume of oxygen reacts with 0.5 mol of sulphur?

**Solution:**

- At RTP 1 mol of gas occupies 24 dm³.
- 0.5 mol of sulphur reacts with 0.5 mol of oxygen
- 0.5 mol oxygen occupies 12 dm³

### Moles and masses

If you cannot remember what is meant by relative atomic mass, read Chapter 2 again before you continue with this chapter.

A word of warning! One mole is \( 6 \times 10^{23} \) particles. The particles can be atoms, molecules, ions or electrons. If there is any possibility of confusion, you must state which particles you mean. For example, what is ‘1 mol of hydrogen’? Do you mean 1 mol of hydrogen gas, which contains \( 6 \times 10^{23} \) molecules and therefore \( 12 \times 10^{23} \) atoms of hydrogen, or do you mean 1 mol of hydrogen atoms, which contains only \( 6 \times 10^{23} \) atoms?
‘1 mol of hydrogen molecules’ has an exact meaning, but ‘1 mol of hydrogen’
does not, so be careful.

**Moles to masses and back again**

We have seen that the $6 \times 10^{23}$ atoms in 1 mol of hydrogen atoms weigh 1 g
(page 42, above). If you imagine that you have the same number of particles
but that each particle has 12 times the mass of a hydrogen atom, then the total
mass of those particles (still 1 mol of them) would be 12 g. If each particle has 12
times the mass of a hydrogen atom then that particle has a relative atomic mass
of 12. The particles you have imagined are atoms of carbon-12. Nowadays, we
use carbon-12 to define one mole of a substance:

One mole of a substance is that mass which contains the same
number of particles as there are atoms in 12 g of carbon-12.

- The mass of a mole of atoms in grams is equal to the relative atomic mass
  of the atoms.
- The mass of a mole of ions or molecules in grams is equal to the relative
  molecular mass of the particles.

The mass of one mole or particles is called the molar mass.

Here is a table of examples.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative atomic mass(es) ($A_r$)</th>
<th>Mass of one mole (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C atom</td>
<td>12</td>
<td>12</td>
</tr>
<tr>
<td>CO$_2$ molecule</td>
<td>$12 + (2 \times 16) = 44$</td>
<td>44</td>
</tr>
<tr>
<td>Cl$^-$ ion</td>
<td>35.5</td>
<td>35.5</td>
</tr>
<tr>
<td>SO$_4^{2-}$ ion</td>
<td>$32 + (4 \times 16) = 96$</td>
<td>96</td>
</tr>
</tbody>
</table>

**Table 4.1** Relative masses and molar masses of some particles.

For example, consider a salt solution that has a mass concentration of 5.85 g dm$^{-3}$.
What is the molar concentration of the solution?

The molar mass of sodium chloride is $(23 + 35.5) = 58.5$ g

$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}}$

so the molar concentration is $\frac{5.85}{58.5} \text{ mol dm}^{-3} = 0.10 \text{ mol dm}^{-3}$

---

**ITQ13**

Fill in the gaps in this table.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative mass ($A_r$ or $M_r$)</th>
<th>Mass of one mole (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(i) CuO molecule</td>
<td></td>
<td></td>
</tr>
<tr>
<td>(ii) He atom</td>
<td></td>
<td></td>
</tr>
<tr>
<td>(iii) NO$_3^-$ ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>(iv) MgCl$_2$ molecule</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

---

**Equations and masses**

Now we can do calculations about reactions using reacting masses.
We can solve these problems by laying out:

- the equation;
- the number of moles involved;
- the relative masses of the particles;
- the reacting masses.

Then we can do any necessary arithmetic.
Worked example 4.6

(i) What mass of oxygen is needed to react with 8 g of hydrogen?
(ii) What mass of water is produced?

Solution:
equation \[ 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \]
moles \[ \begin{array}{c} 2 \\ 1 \\ 2 \end{array} \]
relative masses \[ \begin{array}{c} 2 \times 1 = 2 \\ 2 \times 16 = 32 \\ (2 + 16) = 18 \end{array} \]
reacting masses \[ \begin{array}{c} 2 \times 2 = 4 \\ 32 \\ 2 \times 18 = 36 \end{array} \]

So, 4 g of hydrogen react with 32 g of oxygen to give 36 g of water.
Therefore, 8 g of hydrogen react with 64 g of oxygen to give 72 g of water.

Worked example 4.7

110 g of impure calcium carbonate (CaCO₃) were strongly heated. 56.0 g of calcium oxide (CaO) were produced.
The equation for the reaction is:
\[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]
What is the percentage purity of the calcium carbonate?

Solution:
The \( M_r \) values are CaCO₃ = 100 and CaO = 56
56 g of CaO are produced from 100 g of pure CaCO₃

\[ \% \text{ purity} = \frac{100}{110} \times 100 = 90.9\% \]

Mixing masses and gas volumes

Worked example 4.8

(i) What volume of hydrogen, measured at RTP, is needed to convert 8 g of copper(II) oxide (CuO) into copper metal?
(ii) What mass of water is produced?

You should lay out the data in the same way as before (see worked example 4.6) but leave out the mass data as they are not going to be of any use.

Solution:
equation \[ \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(l) \]
moles \[ \begin{array}{c} 1 \\ 1 \\ 1 \\ 1 \end{array} \]
relative masses \[ \begin{array}{c} 64 + 16 = 80 \\ 2 + 16 = 18 \end{array} \]
reacting masses \[ \begin{array}{c} 80 \text{ g} \\ 18 \text{ g} \end{array} \]

Here are the answers:
(i) 80 g of copper(II) oxide require 1 mol of hydrogen gas
so 8 g of copper(II) oxide require \( \frac{8}{80} = 0.1 \text{ mol of hydrogen gas} \)
1 mol of a gas occupies 24 dm³ at RTP; 0.1 mol occupies 2.4 dm³ at RTP
(ii) 1 mol copper(II) oxide gives 1 mol of water
0.1 mol copper(II) oxide gives 0.1 mol water
1 mol water = 18 g
0.1 mol water = 1.8 g
Calculating formulae from mass data

Empirical formulae

The empirical formula of a compound is the simplest formula which shows the ratio of atoms in that compound. For example, ethane (C$_2$H$_6$) has the empirical formula CH$_3$.

Given the percentage composition of a compound, we can calculate the ratio of the atoms of each element present.

1. Check that the percentages add up to 100!
2. Divide each percentage figure by the relative atomic mass of the element.
3. Divide each of these numerical results by the smallest of them.
4. Express your results as whole numbers. You cannot have fractions of atoms!

You can see how these steps are used in worked examples 4.9 and 4.10.

**Worked example 4.9**

A compound has the composition: C 52.17%, H 13.04%, O 34.78%

Calculate the empirical formula of the substance.

**Solution:**

1. The figures add up to 100% within 2 decimal places.
2. for C $\frac{52.17}{12} = 4.35$ for H $\frac{13.04}{1} = 13.04$ for O $\frac{34.78}{16} = 2.17$
3. for C $\frac{4.35}{2.17} = 2.00$ for H $\frac{13.04}{2.17} = 6.00$ for O $\frac{2.17}{2.17} = 1.00$
4. The empirical formula is C$_2$H$_6$O.

**Worked example 4.10**

A compound has the composition Fe 70%, O 30%

Calculate the empirical formula of the substance.

**Solution:**

1. The figures add up to 100%.
2. for Fe $\frac{70}{56} = 1.25$ for O $\frac{30}{16} = 1.875$
3. for Fe $\frac{1.25}{1.25} = 1.00$ for O $\frac{1.875}{1.25} = 1.50$
4. We cannot have 1.50 of an atom. However, if you express these numbers as halves you get $\frac{1}{2}$ and $\frac{1}{2}$. This gives a ratio of 2 : 3. So the atoms are in the ratio 2 : 3 and the empirical formula is Fe$_2$O$_3$.

Molecular formulae

To find the molecular formula from an empirical formula we need some extra information.

For example, the empirical formula of ethane is CH$_3$. This corresponds to a ‘relative particle mass’ of 15. If it is found that under laboratory conditions, 24 dm$^3$ of ethane weigh 29 g then we know from Avogadro’s law that the relative molecular mass of ethane is roughly 29. Clearly the real relative molecular mass is $2 \times 15 = 30$, showing that the molecular formula is C$_2$H$_6$. 

A sample of a gas was analysed and found to contain only 1.2 g carbon and 0.4 g hydrogen. Calculate the percentage composition of the gas and hence its empirical formula.
Building an equation

The structure of a chemical equation

A chemical equation has:
• on the left, symbols for the reacting particles of each reactant, separated by ‘plus’ signs;
• on the right, symbols for the particles of each of the products, separated by ‘plus’ signs;
• between left and right, an arrow ‘→’ (older books use an ‘equals’ sign ‘=’).

This is a general example:
\[ A + B \rightarrow C + D + E \]

A chemical equation must:
• show the relative numbers of each particle in the equation;
• use the correct formulae for every particle present;
• be balanced (see page 42 above).

A chemical equation may:
• show the physical state of the reactants and products;
• include quantitative information about the energy change in the reaction.

Setting up the equation

To set up an equation, you must know what the reactants are and what products are produced. Sometimes you may only know the main product; the equation will let you make a guess at the others!

Here are the steps you can use.

1. Write a word equation.
2. Put in the correct molecular formula for each substance.
3. Check whether the atoms of each element balance:
   - OK
   - not OK!
4. Add numbers to some molecules to balance the unbalanced element.
5. Recheck and rebalance until everything is balanced.
6. Add the state symbols.

Ionic equations

When an ionic substance dissolves its ions separate and exist independently. Ionic equations are equations which show the individual ion species. Exactly the same rules apply to ionic equations as to molecular equations, but in addition the total electrical charge on each side of the equation must be the same.

The example here represents the reaction between zinc metal and iron(III) ions, giving zinc ions and iron(II) ions. The total charge is +6 on each side. The equation balances.

You may ask ‘Where did the iron(III) ions come from?’ They can only be found combined with negative ions, such as chloride or sulphate. To see why the equation is correct we use the idea of spectator ions.

An equation written in ionic form often turns out to have some of the same ions on each side. For example, in the reaction between sodium hydroxide and hydrochloric acid we have:

\[ \text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
\[ \text{Na}^+(aq) + \text{OH}^-(aq) + \text{H}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) + \text{H}_2\text{O(l)} \]

The sodium ions and the chloride ions are there at the start of the reaction and they are there at the end. They take no part in the reaction and are called spectator ions. To make the equation clearer we can leave them out:

\[ \text{OH}^-(aq) + \text{H}^+(aq) \rightarrow \text{H}_2\text{O(l)} \]

There is no reaction other than the formation of water!
If a solid is formed as part of an ionic reaction, the solid is shown in molecular form in the equation even though the solid is in fact an ionic lattice.

For example, when aqueous sodium sulphate reacts with aqueous barium chloride, barium sulphate is precipitated – this is a test for aqueous sulphate ions.

\[
\text{Na}^+(aq) + \text{SO}_4^{2-}(aq) + \text{Ba}^{2+}(aq) + \text{Cl}^-(aq) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) + \text{BaSO}_4(s)
\]

The equation reduces to:

\[
\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s)
\]

No charges are written in the molecular formula so in counting up charges to balance the equation, the solid has zero charge.

**Worked example 4.11**

Rewrite the equation for the reaction between aqueous silver nitrate and aqueous sodium chloride:

\[
\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3
\]

Add state symbols but remove the spectator ions. Remember that this reaction provides a test for chloride ions.

The silver chloride forms a precipitate.

**Solution:**

Adding state symbols we get:

\[
\text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq)
\]

Writing this as an ionic equation:

\[
\text{Ag}^+(aq) + \text{NO}_3^-(aq) + \text{Na}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) + \text{AgCl}(s)
\]

Removing spectator ions leaves us with:

\[
\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)
\]

---

**Summary**

- Each chemical element has a symbol used to mean one atom (or one mole) of that element.
- Subscripts and superscripts can be added to a symbol to indicate mass number, atomic number, electric charge and number of particles.
- Particles are counted in units of \(6.02 \times 10^{23}\) (approx. \(6 \times 10^{23}\)). This unit is called the mole (mol).
- The Avogadro constant has the value \(6.02 \times 10^{23}\) mol\(^{-1}\).
- One mole of a substance is that mass which contains the same number of particles as there are atoms in 12 g of carbon–12.
- One mole of any gas occupies 24.0 dm\(^3\) at RTP (or 22.4 dm\(^3\) at STP).
- Since moles can refer to any particle, it is important to specify which particles are being counted if there is any ambiguity.
- The mass of one mole of particles is called the molar mass.
- Chemical equations state, in symbol form, the reactants and products of a reaction and the molar quantities in which they react.
- Chemical equations may also contain state symbols, showing the physical state of the substances.
- If chemical equations are written in ionic form, ions that take no part in the reaction (spectator ions) may be omitted.
- Any chemical equation must have the same atoms and the same total electric charge on both sides. Such an equation is **balanced**.
**ITQ1**  A molecule made by combining one atom of iron (Fe) with one atom of oxygen (O).

**ITQ2**  
(i)  \( \text{NH}_3 \) (if you wrote \( H_3N \) then count this as right – with compounds containing hydrogen there is no easy rule about the order in which we write the symbols in a formula)

(ii) No. Each hydrogen is only bonded to the nitrogen atom.

**ITQ3**  
(i)  An atom of copper with a double positive charge, solid.

(ii)  A molecule of hydrogen made up of two atoms of hydrogen.

(iii)  A molecule of hydrogen made up of two atoms of hydrogen, as a gas.

(iv)  A nitrogen atom with a triple negative charge.

**ITQ4**  
(i)  \( O_2(l) \)

(ii)  \( Cl_2(g) \)

(iii)  \( Na^+ \)

**ITQ5**  Nothing is gained or lost. 6.0 g + 1.7 g = 7.7 g

**ITQ6**  
(i)  4.0 g

(ii)  1.0 mol dm\(^{-3}\)

**ITQ7**  2.0 g

**ITQ8**  0.10 mol dm\(^{-3}\)

**ITQ9**  0.025

**ITQ10**  0.5 mol

**ITQ11**  
(i)  0.001

(ii)  2

(iii)  6

**ITQ12**  36 dm\(^3\) of \( H_2 \)  24 dm\(^3\) of \( NH_3 \)

**ITQ13**  
(i)  79.5, 79.5 g

(ii)  4, 4 g

(iii)  62, 62 g

(iv)  95, 95 g

**ITQ14**  1.0 mol carbon dioxide = 44 g  
32 g oxygen (= 1 mol) occupies 24 dm\(^3\)  
44 g occupy \( \frac{24 \times 44}{32} \) = 33 dm\(^3\)

**ITQ15**  1.2 g Mg = 0.1 mol acid contains \( 0.1 \times 0.5 = 0.05 \text{ mol} \) so the metal is in excess

**ITQ16**  6.5 g zinc = \( \frac{6.5}{65} \) = 0.1 mol  
1.0 mol zinc requires 2.0 mol hydrochloric acid and produces 1.0 mol of hydrogen  
0.1 mol zinc require 0.2 mol hydrochloric acid  
\( M_r \) for hydrochloric acid = 36.5  
\( 0.2 \times 36.5 = 7.3 \text{ g of HCl} \)  
0.2 mol of gas has a volume of \( 0.2 \times 24 = 4.8 \text{ dm}^3 \)

**ITQ17**  C 75%, H 25%; \( \text{CH}_4 \)

**ITQ18**  \( \text{N}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{NO}_2 + 2\text{H}_2\text{O} \)

**ITQ19**  
(i)  \( \text{Cu}^{2+} + \text{SO}_4^{2-} + 2\text{Na}^+ + \text{CO}_3^{2-} \rightarrow \text{CuCO}_3 + 2\text{Na}^+ + \text{SO}_4^{2-} \)

(ii)  \( \text{Cu}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{CuCO}_3(s) \)

**ITQ20**  
(i)  \( \text{Zn}(s) + \text{Fe}^{3+}(aq) + 3\text{Cl}^- (aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Fe}^{2+} + 3\text{Cl}^- (aq) \)  
\( \text{(unbalanced)} \)

or  \( \text{Zn}(s) + 2\text{Fe}^{3+}(aq) + 6\text{Cl}^- (aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Fe}^{2+} + 6\text{Cl}^- (aq) \)  
\( \text{(balanced)} \)

(ii)  \( \text{Zn}(s) + 2\text{Fe}^{3+}(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Fe}^{2+} \) (without spectator ions)
1 (i) A friend says that he will give you one cubic centimetre of pure gold. Another friend says that she will give you one mole of pure gold. Which of the two quantities of gold weighs more than the other?

(ii) Gold can react with chlorine to form gold(III) chloride.
\[2Au(s) + 3Cl_2(g) \rightarrow 2AuCl_3(s)\]
What volume of chlorine, measured at RTP, would be required to react exactly with 0.1 mole of gold?
Data: molar gas volume at RTP is 24 dm\(^3\) mol\(^{-1}\), density of gold is 19.3 g cm\(^{-3}\), \(M_r\) is 197.

2 (i) Bauxite is impure aluminium oxide. A sample of bauxite weighed 19.1 g. After purification the mass of the bauxite was 18.8 g. Calculate what percentage of the original bauxite was aluminium oxide.

(ii) A crystalline sample of sodium carbonate weighed 4.64 g. The sample was dissolved in water and titrated against aqueous hydrochloric acid. The acid had a concentration of 2.00 mol dm\(^{-3}\). For complete reaction, 20.0 cm\(^3\) of acid were needed. The reaction taking place is:
\[Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl + CO_2(g) + H_2O(l)\]
Calculate the following.
(a) How many moles of sodium carbonate were present in the sample.
(b) How many moles of sodium carbonate \(Na_2CO_3\) were in the sample?
(c) What mass of sodium carbonate was in the original sample.

(iii) The original sample had the formula \(Na_2CO_3\cdot xH_2O\). Calculate the value of \(x\).

3 The reaction between nitrogen and hydrogen in the Haber process is:
\[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)\]

(i) What volume of hydrogen (measured at RTP) is needed to react exactly with 2.8 kg of nitrogen?

(ii) What mass of ammonia would be produced assuming that reaction was complete?

(iii) What mass of ammonium nitrate fertiliser could be made from this amount of ammonia?

4 (i) Describe how you would find the concentration (in mol dm\(^{-3}\)) of a solution of potassium hydroxide, given a solution of hydrochloric acid of concentration 0.10 mol dm\(^{-3}\) and a supply of universal indicator.

(ii) In the titration, 25 cm\(^3\) of acid reacted exactly with 20 cm\(^3\) of base solution. Calculate the concentration of the base.

(iii) 25 cm\(^3\) of a dibasic acid solution containing 4.9 g dm\(^{-3}\) reacted exactly with 25 cm\(^3\) of a solution of a base of concentration 0.50 mol dm\(^{-3}\). What is the relative molecular mass of the acid?
By the end of this chapter you should be able to:

- explain what is meant by a ‘stable electronic structure’;
- explain how atoms lose or gain electrons to achieve a stable electron structure;
- describe the formation of ionic and covalent bonds;
- describe the formation of multiple covalent bonds;
- explain the energy changes involved in the formation of bonds;
- explain the properties of materials that contain ionic or covalent bonds;
- have an understanding of the meaning of the term ‘lattice’;
- explain which type of bond will be formed by a specified pair of elements;
- have an understanding of
  - the metal bond,
  - the hydrogen bond,
  - the van der Waals’ bond;
- understand what is meant by allotropy and give examples.

**Concept map**

Stable electronic structures

As a general rule, natural objects contain as little energy as possible. Hot saucepans cool down – they do not heat up by themselves. A brick may fall down from the top of a building, but it will never jump up there by itself! The kettle and the brick lose energy to reach their lowest energy state without any help from us. They become more stable.

We saw in chapter 2 that atoms have a central nucleus surrounded by electrons, and that the electrons around the nucleus are arranged in shells. An atom with a full outermost electron shell is an atom of a noble gas, such as helium, argon or neon. Noble gases exist as single atoms because the full outer shell makes each atom stable.

Atoms of other elements do not have full outer shells. To have (or at least have a share in) full outer shells they must combine with other atoms. Having a full outer shell makes the atoms more stable. Atoms can combine with other atoms of the same element, as in oxygen (O₂), or they can combine with atoms of a different element, as in sodium chloride (NaCl).
Full shells formed by electron transfer

Losing an electron

Figure 5.3 shows an atom of sodium, \( ^{23}_{11}\text{Na} \), with shells containing 2 electrons, 8 electrons and 1 electron. We write the electron structure of sodium as (2,8,1).

If a sodium atom loses one electron, the shell which is now its outer shell is full. The atom has an overall charge of +1 because the nucleus still has 11 protons but it has only 10 electrons around the nucleus. The atom is now a positive ion (Na\(^+\)).

![An atom of sodium has one electron in its outer shell. The outer shell is not full.](image)

**Figure 5.3** An atom of sodium has one electron in its outer shell. The outer shell is not full.

Losing electrons always uses energy — energy is added to pull the electron away from the protons.

A cation is a positively charged ion.

![A sodium atom loses an electron to achieve a noble gas electron structure.](image)

**Figure 5.4** A sodium atom loses an electron to achieve a noble gas electron structure.

Gaining an electron

The electron structure of chlorine is (2,8,7). Figure 5.5 shows an atom of chlorine, \( ^{35}_{17}\text{Cl} \), with shells containing 2 electrons, 8 electrons and 7 electrons. If a chlorine atom gains one electron, its outer shell is then full (holding 8 electrons). It has an overall charge of –1 because the nucleus has 17 protons with 18 electrons around it. The atom is now a negative ion (Cl\(^–\)).

![An anion is a negatively charged ion.](image)

**Figure 5.5** Gaining an electron to achieve a noble gas structure.

Anion

Gaining more than one electron needs energy to be added, to force the electron in against the repulsion of the other electrons.

Charges on stable ions

The number of electrons that need to be gained or lost to get a full outer shell depends on the structure of the original atom. An atom with six electrons in its outer shell must gain two more. An atom with two electrons in its outer shell must lose both to expose the full shell underneath them. Sometimes three electrons must be either gained or lost, but this is not common. Some examples are shown in table 5.1.

<table>
<thead>
<tr>
<th>Atom structure</th>
<th>Number of electrons gained or lost</th>
<th>Ion formed</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>......, 1</td>
<td>1 lost</td>
<td>( \text{X}^+ )</td>
<td>( \text{H} \rightarrow \text{H}^+ ) ( \text{Na} \rightarrow \text{Na}^+ )</td>
</tr>
<tr>
<td>......, 2</td>
<td>2 lost</td>
<td>( \text{X}^{2+} )</td>
<td>( \text{Mg} \rightarrow \text{Mg}^{2+} ) ( \text{Ca} \rightarrow \text{Ca}^{2+} )</td>
</tr>
<tr>
<td>......, 3</td>
<td>3 lost</td>
<td>( \text{X}^{3+} )</td>
<td>( \text{Al} \rightarrow \text{Al}^{3+} ) ( \text{Fe} \rightarrow \text{Fe}^{3+} )</td>
</tr>
<tr>
<td>......, 4</td>
<td>these atoms do not form ions</td>
<td></td>
<td></td>
</tr>
<tr>
<td>......, 5</td>
<td>3 gained</td>
<td>( \text{X}^{3–} )</td>
<td>( \text{N} \rightarrow \text{N}^{3–} ) (nitride)</td>
</tr>
<tr>
<td>......, 6</td>
<td>2 gained</td>
<td>( \text{X}^{2–} )</td>
<td>( \text{O} \rightarrow \text{O}^{2–} ) (oxide)</td>
</tr>
<tr>
<td>......, 7</td>
<td>1 gained</td>
<td>( \text{X}– )</td>
<td>( \text{Cl} \rightarrow \text{Cl}– ) (chloride)</td>
</tr>
</tbody>
</table>

**Table 5.1** Forming ions with noble gas structures.
Worked example 5.1

Aluminium atoms have the electron structure 2,8,3.

(i) There are two ways the electron structure could change to give the aluminium atom a noble gas structure. What are they? Which is more likely?

(ii) Which noble gas has the electron structure you chose in part (i)?

**Solution:**

(i) Aluminium atoms must either lose three electrons to expose their full second shell or gain five electrons to complete their third shell. Atoms never gain or lose more than three electrons, so the structure of the aluminium ion is 2,8.

(ii) From the Periodic Table you can see that neon (Ne) has the electron structure 2,8.

---

Worked example 5.2

Magnesium atoms have the structure 2,8,2.

Chlorine atoms have the structure 2,8,7.

By working out the charges on magnesium ions and on chloride ions, deduce the formula of magnesium chloride.

**Solution:**

Magnesium atoms lose two electrons to form 2,8; the ion is therefore Mg\(^{2+}\).

Chlorine atoms gain one electron to form 2,8,8; the ion is therefore Cl\(^{-}\).

The compound must be electrically neutral, so it must be MgCl\(_2\).

---

**Bonding**

**Ionic bonds**

If electrons are transferred from one atom to another, the positive and negative ions that are formed attract one another. There is an electrostatic force between them. They will move towards one another, giving out energy as they do so. To see that they will give out energy, imagine that you have to pull them apart again. This would need you to put energy in. So that energy must be given out as they move together. The energy they give out will go towards ionising more atoms so that the process can continue. (See chapter 8 'Energy in chemistry'.) Once together the ions will stay together unless they are pulled apart again. There is a bond between the ions. This bond is called an ionic bond. Ionic bonds are formed by the transfer of electrons from one atom to another. Although the charges on the ions are equal and opposite they do not ‘cancel each other out’ and more ions can join on them, forming a large stack of ions called a lattice.

An ionic crystal is a crystal formed from a lattice of ions of opposite charges.

**Covalent bonds**

Turning atoms into ions needs energy. For example, you need 2200 kJ to form a mole of magnesium ions (Mg\(^{2+}\)). But you need much more energy, as much as 11,300 kJ, to form a mole of carbon ions (C\(^{4+}\)). Because of this, carbon atoms do not ionise! Instead they ‘share’ electrons.
Look at Figure 5.7 which shows a carbon atom and a hydrogen atom (a) far apart and (b) close together.

In figure 5.7(b) the atoms are so close together that their outer shells overlap. There are now two electrons in the outer shell of the hydrogen atom, which gives it the stable helium structure. Adding three more hydrogen atoms in the same way gives the carbon atom eight electrons in its outer shell and gives it the stable structure 2,8 of neon.

The compound CH₄ (methane) is stable. Figure 5.8 shows a ‘dot-and-cross’ diagram for CH₄.

In CH₄ the carbon atom has formed four covalent bonds, one with each of the hydrogen atoms. The electrons between the two nuclei stick them together much as the jam holds the pieces of bread in a jam sandwich. A covalent bond is made up of a pair of shared electrons. The shared electrons are located between the nuclei of the atoms that they bond.

A covalent bond is shown as a line joining the two atoms

H

H—C—H

H

In a covalent bond each nucleus is attracted to the electrons between them. But the nuclei are both positively charged and so repel each other. Why do the atoms stay together?

ITQ3

Draw a dot-and-cross diagram for the molecule of ethane (C₂H₆). Use different symbols for the electrons from carbon atom 1, carbon atom 2 and the hydrogen atoms.

A tetrahedron

Carbon chains

Some elements are capable of forming short chains of atoms. Carbon can form long chains, up to several thousand atoms in length. These chains form the backbones of that molecules made by living things, and we shall study some of them in chapters 12–16. Carbon atoms can also combine in three dimensions, with each carbon atom bonding to four others, to form diamond. Figures 5.9 and 5.10 show two views of this structure. The bonds from each carbon atom point outwards towards the corners of a tetrahedron.

A covalent crystal is a crystal containing separate molecules of the substance held together by other bonds (e.g. van der Waals’ forces).

Carbon atoms can link up in a different way to form graphite, and in yet another to form fullerene. Because these forms of carbon have different structures, they also have different properties, which we shall examine later on page 60.
Multiple covalent bonds
Atoms can also share two (and even three) pairs of electrons to form multiple covalent bonds. For example, carbon forms double bonds (two shared pairs of electrons) in carbon dioxide. Each oxygen atom had six electrons in its outer shell; now it has eight. The carbon atom had four electrons in its outer shell; now it has eight (4 + 2 + 2). All the atoms now have the neon structure and the CO₂ unit is stable.

Carbon also forms double bonds with other carbon atoms. Look at figure 5.12. Each of these two carbon atoms still has two unused electrons. The carbon atoms do not yet have eight electrons in their outer shells and so the molecule C₂ is not stable. It cannot exist in the free state. If other atoms (for example, hydrogen) bond to the remaining electrons, then each carbon atom has eight electrons in its outer shell and the substance (which is ethene) is stable (figure 5.13).

Carbon can form triple bonds, –C≡C–, in a similar way. The triple bond that holds nitrogen atoms together in the N₂ molecule (N≡N) is particularly strong, which is why nitrogen is so unreactive.

Comparing ionic and covalent bonds
- Covalent bonds and ionic bonds are strong bonds.
- Ionic bonds lead to lattices of ions, not individual molecules. When a positive ion and a negative ion come together the charges are not cancelled out.
- Covalent bonds lead to either:
  (i) individual molecules, which are held one to another by weaker forces (such as van der Waals’ forces, see page 58) or
  (ii) giant lattices.

Ionic compounds
- Ionic compounds have high melting points because the ionic bonds are strong.
- Ionic compounds are soluble in water, rather than organic liquids, because ‘like dissolves like’.
- Ionic compounds form electrically conducting solutions because they contain ions – they are electrolytes. Ionic compounds also conduct electricity when molten.
- Ionic crystals are brittle because the ions are locked into place in their lattice. They cannot be drawn into wires or hammered into shape as metals can.

Covalent substances
- Covalent giant lattices have high melting points and covalent molecules are stable to heating because their covalent bonds are strong.
- Covalent molecules have low melting points because they are held one to another by weak bonds.
- Covalent compounds (and elements with covalent bonds) dissolve in organic liquids, rather than water, because ‘like dissolves like’.
- Solutions of organic materials do not conduct electricity because they don’t contain any ions.
- Covalent substances can be deformed or act as lubricants because the weak bonds holding their units together are easily broken. Substances that contain

**Melting points of some materials (°C)**

<table>
<thead>
<tr>
<th>Material</th>
<th>Melting Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>ethanol</td>
<td>-115</td>
</tr>
<tr>
<td>methane</td>
<td>-183</td>
</tr>
<tr>
<td>diamond</td>
<td>&gt;3350</td>
</tr>
<tr>
<td>sand</td>
<td>2230</td>
</tr>
<tr>
<td>lead bromide</td>
<td>918</td>
</tr>
<tr>
<td>sodium chloride</td>
<td>800</td>
</tr>
<tr>
<td>glucose</td>
<td>146</td>
</tr>
<tr>
<td>iron(III) oxide</td>
<td>1560</td>
</tr>
</tbody>
</table>

**ITQ5**
(i) Draw a dot-and-cross diagram for ethene (C₂H₄). You need to show a carbon–carbon double bond.
(ii) Draw a dot-and-cross diagram for ethyne (C₂H₂). You need to show a carbon–carbon triple bond.

*Use different symbols for the electrons from carbon atom 1, carbon atom 2 and the hydrogen atoms.*

**Figure 5.11** A dot-and-cross diagram for CO₂.

**Figure 5.12** The molecule C₂ isn’t stable.

**Figure 5.13** A ball-and-stick model of ethene H₂C=CH₂.

**ITQ6**
Here is some information about substance X.
- X dissolves in alcohol but is sparingly soluble in water.
- X melts at 122 °C.
- X produces carbon dioxide gas when mixed with aqueous sodium carbonate.

What can you deduce about X?
only covalent bonds are exceptions. Diamond and silicon dioxide are examples. A giant molecular crystal is a crystal containing atoms linked by covalent bonds without any separate molecules.

Relating structure to properties and uses

Example 1: sodium chloride, an ionic lattice
- Sodium chloride crystals melt at 800 °C.
- Aqueous sodium chloride is a good conductor of electricity.
- Sodium chloride crystals are brittle.

![A space-filling diagram of NaCl.](Figure 5.14)

![The structure of NaCl.](Figure 5.15)

Example 2: naphthalene, a covalent molecule
- Naphthalene melts at 80 °C.
- Naphthalene is insoluble in water. Solutions of naphthalene in organic liquids do not conduct electricity.

Example 3: diamond, a covalent lattice
- Diamond melts above 3500 °C.
- Diamond is insoluble in all liquids.
- Diamond is the hardest natural substance known. The lattice is held together in all directions by carbon–carbon bonds, which are very strong.

Example 4: graphite, a covalent lattice and van der Waals’ bonds
- Graphite melts above 3500 °C.
- Graphite is insoluble in any liquid.
- Graphite contains two-dimensional sheets of carbon atoms as giant lattices. The sheets are held together by van der Waals’ bonds (weak bonds). The sheets slide easily over one another.
- The bonding in each sheet is covalent but because there are many atoms joined together some electrons go free. The bonding is like that in a metal. Graphite is a good conductor of electricity.

![A space-filling diagram of graphite.](Figure 5.16)

![The structure of graphite.](Figure 5.17)
Example 5: iodine, covalent molecular

- Iodine forms black crystals.
- Iodine sublimes at 113 °C, forming a purple vapour.
- The vapour contains I$_2$ molecules.

**The metallic bond**

Metals have three important properties:

1. They have high density;
2. They are malleable (they can be hammered into different shapes);
3. They conduct electricity without any chemical change taking place.

- The first property tells us that the atoms must be packed closely together in a solid metal.
- The second property tells us that whatever the bonds that hold the atoms together are like, they can form between any atoms in the structure, and not just between pairs of atoms.
- The third property tells us that there must be mobile electrons in the structure. More experiments show that metals obey Ohm’s law, so that the electrons in the structure must be totally free from the atomic nuclei.

Taken together, these properties give us a picture of the bonding in a metal as involving a cloud of electrons floating in an array of closely packed positive ions. The electrons, although they are no longer a part of any specific atom, act as an electrostatic glue holding the metal cations together.

When a potential difference is applied across a piece of metal the electrons drift slowly from the negative connection to the positive. This flow of electrons is the electric current. Because the ions themselves do not move (they are in the
solid state and therefore are locked in place) there is no chemical change. As fast as electrons move out of the metal at the positive connection they are replaced by electrons flowing in at the negative connection.

A summary

<table>
<thead>
<tr>
<th></th>
<th>Melting point</th>
<th>Electrical conductivity</th>
<th>Solubility in water</th>
<th>Solubility in covalent liquid</th>
</tr>
</thead>
<tbody>
<tr>
<td>covalent molecular</td>
<td>low</td>
<td>low</td>
<td>low unless molecule contains an –OH group</td>
<td>high</td>
</tr>
<tr>
<td>covalent lattice</td>
<td>high</td>
<td>low</td>
<td>very low</td>
<td>very low</td>
</tr>
<tr>
<td>ionic</td>
<td>high</td>
<td>low</td>
<td>low when solid, high in solution or when molten</td>
<td>often high; exceptions are many oxides, carbonates and sulphates</td>
</tr>
<tr>
<td>metallic</td>
<td>high, except Group I elements</td>
<td>high</td>
<td>very low</td>
<td>very low</td>
</tr>
</tbody>
</table>

*Table 5.2 Some properties of covalent, ionic and metallic substances.*

**van der Waals’ forces: holding molecules together**

- Hydrocarbons with large molecules, for example \( \text{C}_{18}\text{H}_{38} \) (which is the main component of candle wax), are solids at room temperature.
- Hydrocarbons with smaller molecules, for example hexane \( \text{C}_6\text{H}_{14} \), are liquids at room temperature and pressure.
- Even smaller molecules such as methane \( \text{CH}_4 \) are gaseous at room temperature and pressure.

However, even tiny particles such as helium atoms will turn to a liquid if the temperature is low enough. In the liquid and solid states, there must be some sort of bond holding the particles together. That it is a weak bond is shown by the fact that the melting point of hexane is only \(-95^\circ \text{C}\) and that of helium is \(-270^\circ \text{C}\)!

This bond is neither a covalent bond nor an ionic bond, both of which are strong bonds. This weak bond is named after Johannes Diderik van der Waals (1837–1923), a scientist from the Netherlands who won a Nobel prize for physics in 1910. In 1869 he had been interested in the behaviour of gases under pressure, and the idea of weak bonds between the particles grew from that. Nowadays we call the bonds van der Waals’ forces.

Van der Waals’ forces exist between atoms and between molecules. Large molecules have more of these forces between them because they have more atoms, and consequently larger molecules have higher melting and boiling points. In general, we say that the higher the relative molecular mass of a compound, the higher the melting and boiling points.

**The hydrogen bond**

Water, \( \text{H}_2\text{O} \), is a fascinating compound.
- Solid water is less dense than liquid water, which is very unusual for a solid form of a molecule compared to its liquid form.
- It takes more energy to heat 1 g of water through 1 °C than for 1 g of any other substance.
- Above all, although \( M_r \) for water is only 18, the melting point of ice is 0 °C and the boiling point of water is 100 °C (these are much higher than you’d expect).
Figure 5.22 shows the boiling points of all the compounds $H_2X$, where $X$ is an element in Group VI. You can see that the boiling point of $H_2O$ ‘ought’ to be nearer $-100^\circ C$ rather than $+100^\circ C$!

Water behaves as if it has a much higher relative molecular mass than in fact it does have. This suggests that water behaves more like a polymer ($H_2O_n$) than a single molecule. But what can hold the single molecules together?

The answer is the hydrogen bond. Whenever a hydrogen atom is sandwiched in space between two oxygen atoms, a bond called the hydrogen bond is made that holds the three atoms together. This is shown for water in figure 5.23. A hydrogen bond is stronger than a van der Waals’ bond but weaker than a covalent bond or an ionic bond.

Substances which form hydrogen bonds are usually soluble in water.
- Ethanol, which you can think of as water with one hydrogen atom replaced by an ethyl group (see chapter 12), mixes with water in all proportions.
- Glucose, which is a covalent compound but has five $–OH$ groups in its molecule, is very soluble in water.

**Compounds containing more than one type of bond**

**Molecular compounds**

Molecular compounds have covalent bonds holding the atoms of each molecule together and they also have van der Waals’ forces holding the molecules one to another. We have already seen that this leads to low melting and boiling points, since the intermolecular forces (the forces between the molecules) are weak, and that such compounds do not conduct electricity.

Molecular compounds which contain hydrogen bonds have melting and boiling points which are higher than the relative molecular mass of the compound would suggest, and they are more soluble in water than other materials.

**Ionic compounds**

Some ions contain covalent bonds. For example, the oxygen atoms in a sulphate ion ($SO_4^{2-}$) are covalently bonded to the sulphur, but this has no effect on the properties of a metal sulphate. Metal sulphates have high melting points and conduct electricity when molten and when in solution.
Covalent lattices

In diamond the carbon atoms are held together by covalent bonds. Diamond is strong in every direction and is the hardest natural substance known.

Graphite is another form of carbon. In graphite the atoms are covalently bonded into flat sheets. The sheets are held together by van der Waals’ forces, which are much weaker than covalent bonds. The sheets are strong but they can slide over one another like the leaves of a book. This explains why graphite is an excellent dry lubricant.

Graphite, although it is a non-metal, is also a good conductor of electricity along the sheets, but it does not conduct electricity from one sheet to the next. The reasons for this are outside the scope of this course.

Giant molecules

Polymers like poly(ethene) (see chapter 16) are mixtures of very large molecules of different sizes. The molecules are based on long chains that can be up to 200 000 carbon atoms long. The atoms within a long chain are bonded by covalent bonds. The chains are held together by van der Waals’ forces. Poly(ethene) is not a pure material, as its molecules are of different lengths, so it does not have a sharp melting point. It begins to soften at about 100 °C, as enough van der Waals’ forces break to let some of the molecules move across each other.

Other polymers, such as ‘Bakelite’, form chains which then link up one to another (they ‘cross-link’). This makes a rigid structure that does not melt when heated.

Allotropy

When an element can exist in more than one form at room temperature and pressure, the different forms are called allotropes. The element is showing allotropy. Because the properties of a substance are partly set by its structure, the allotropes of an element usually have different appearances, melting and boiling points, and electrical properties.

Carbon has three allotropes: diamond, graphite and fullerene.

Diamond is crystalline and very hard. Large diamonds are used in jewellery, but most small diamonds are used to tip the industrial drills used, for example, to cut rock. Diamond does not conduct electricity.

Graphite is soft and slippery; it is used as a dry lubricant and to make the ‘leads’ of pencils. Some of the electrons holding the graphite sheets together are mobile, like those in a metal, so graphite is a good conductor of electricity.

Fullerene exists as separate molecules each containing 60 atoms (C_{60}). The football-like structure of a fullerene molecule is shown in figure 5.25. New uses for fullerenes are being found every day. For example, they are used in organic solar cells and some flat panel displays. When grown in tubular form (‘nanotubes’) they have 20 times the tensile strength of the best steel but only half the density of aluminium.

Other forms of carbon such as soot and charcoal are rarely pure, but mostly consist of tiny fragments of graphite.

Other elements which show allotropy are sulphur (crystalline and plastic), phosphorus (red, yellow and black) and tin (metallic and non-metallic).

The allotropy of tin is worth a second look. From its low position in the Periodic Table you would perhaps predict that tin is a metal. But we find that there is a trend in properties as we go down a Group in the Periodic Table. Carbon, at the top of Group IV, is a non-metal and lead, at the bottom of the Group, is a metal. The fact that tin has a non-metal allotrope shows that it is not at the limit of the trend down the Group.
Atoms or ions with full outer shells of electrons are stable.
- Atoms may gain, lose or share electrons to become stable.
- Ionic bonds are formed between ions of opposite charge.
- Covalent bonds are formed when atoms share electrons.
- Covalent and ionic bonds are strong bonds.
- Ionic substances form ionic lattices, not molecules.
- Covalent substances form either molecules or giant lattices.
- Other bonds are the metallic bond, the hydrogen bond, and the van der Waals’ force.
- Elements whose atoms bond together in more than one way exist as allotropes.

**Answers to ITQs**

**ITQ1** Neon, Ne.

**ITQ2** 3.

**ITQ3** The force of attraction between the nuclei and the electrons positioned between the nuclei is stronger than the force of repulsion between the two nuclei.

**ITQ4**

**ITQ5** (i)

**ITQ6** X is an organic acid. It is a compound with covalent bonds because it dissolves in alcohol. It does not contain ionic bonds because it is only sparingly soluble in water. It is an acid, because of its reaction with aqueous sodium carbonate.
ITQ7  Y is a solid at room temperature so it either has largish molecules or it is ionic. Y melts at a low temperature and boils readily, so it cannot be ionic.

However, Y must contain van der Waals’ forces between molecules but cannot be a giant lattice. Y contains molecules which are held together by strong bonds since they survive unchanged being heated to their boiling point.

ITQ8  One form of carbon (graphite) is a good conductor of electricity.

1 (i) Explain, in terms of chemical bonding, why ammonia has a much lower melting point than water.

(ii) (a) Draw a dot-and-cross diagram to show the covalent bonds in a molecule of ammonia.

(b) Use your diagram to explain why ammonia molecules can react with hydrogen ions but not with hydrogen gas.

2 The elements in period 3 of the Periodic Table are, in order: Na Mg Al Si P S Cl Ar

(i) (a) State what type of bond is formed between sodium and chlorine.

(b) Explain why these elements form this type of bond.

(ii) Sulphur and chlorine form a compound SCl₂.

(a) State which type of bond you expect to be present between sulphur atoms and chlorine atoms in this compound.

(b) Suggest two ways in which the physical properties of the compound are different from the properties of sodium chloride.

(c) Use your answer to part (ii) (a) to explain why you think the two compounds differ in these ways.

3 (i) A metal can be described as ‘a crowd of ions in an atmosphere of electrons’. Explain why this is a good description of the structure of a metal.

(ii) The diagrams show part of the electron arrangements in an atom of (a) lithium and (b) chlorine. Complete the diagrams and draw new diagrams to show how an ionic bond is formed between them.

(iii) Why does lithium metal conduct electricity well whereas solid lithium chloride does not.