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The Free High School Science Texts: Textbooks for High School Students Studying the Sciences Chemistry Grades 10 - 12

Version 0
November 9, 2008

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Part II

Matter and Materials

Chapter 1

Classification of Matter - Grade 10

All the objects that we see in the world around us, are made of **matter**. Matter makes up the air we breathe, the ground we walk on, the food we eat and the animals and plants that live around us. Even our own human bodies are made of matter!

Different objects can be made of different types of matter, or **materials**. For example, a cupboard (an *object*) is made of wood, nails and hinges (the *materials*). The **properties** of the materials will affect the properties of the object. In the example of the cupboard, the strength of the wood and metals make the cupboard strong and durable. In the same way, the raincoats that you wear during bad weather, are made of a material that is waterproof. The electrical wires in your home are made of metal because metals are a type of material that is able to conduct electricity. It is very important to understand the properties of materials, so that we can use them in our homes, in industry and in other applications. In this chapter, we will be looking at different types of materials and their properties.

The diagram below shows one way in which matter can be classified (grouped) according to its different properties. As you read further in this chapter, you will see that there are also other ways of classifying materials, for example according to whether they are good electrical conductors.

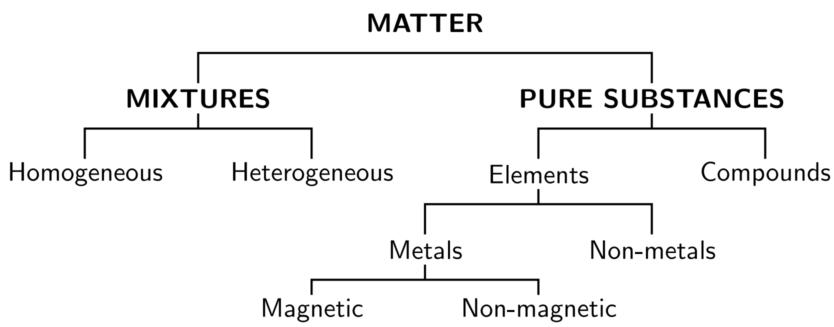


Figure 1.1: The classification of matter

1.1 Mixtures

We see mixtures all the time in our everyday lives. A stew, for example, is a mixture of different foods such as meat and vegetables; sea water is a mixture of water, salt and other substances, and air is a mixture of gases such as carbon dioxide, oxygen and nitrogen.

Definition: Mixture
A **mixture** is a combination of more than one substance, where these substances are not bonded to each other.

In a mixture, the substances that make up the mixture:

- *are not in a fixed ratio*

  Imagine, for example, that you have a 250 ml beaker of water. It doesn't matter whether you add 20 g, 40 g, 100 g or any other mass of sand to the water; it will still be called a mixture of sand and water.

- *keep their physical properties*

  In the example we used of the sand and water, neither of these substances has changed in any way when they are mixed together. Even though the sand is in water, it still has the same properties as when it was out of the water.

- *can be separated by mechanical means*

  To separate something by 'mechanical means', means that there is no chemical process involved. In our sand and water example, it is possible to separate the mixture by simply pouring the water through a filter. Something *physical* is done to the mixture, rather than something *chemical*.

Some other examples of mixtures include blood (a mixture of blood cells, platelets and plasma), steel (a mixture of iron and other materials) and the gold that is used to make jewellery. The gold in jewellery is not pure gold but is a mixture of metals. The *carat* of the gold gives an idea of how much gold is in the item.

We can group mixtures further by dividing them into those that are heterogeneous and those that are homogeneous.

1.1.1 Heterogeneous mixtures

A **heterogeneous** mixture does not have a definite composition. Think of a pizza, that is a mixture of cheese, tomato, mushrooms and peppers. Each slice will probably be slightly different from the next because the toppings like the mushrooms and peppers are not evenly distributed. Another example would be granite, a type of rock. Granite is made up of lots of different mineral substances including quartz and feldspar. But these minerals are not spread evenly through the rock and so some parts of the rock may have more quartz than others. Another example is a mixture of oil and water. Although you may add one substance to the other, they will stay separate in the mixture. We say that these heterogeneous mixtures are *non-uniform*, in other words they are not exactly the same throughout.

Definition: Heterogeneous mixture
A heterogeneous mixture is one that is non-uniform, and where the different components of the mixture can be seen.

1.1.2 Homogeneous mixtures

A **homogeneous** mixture has a definite composition, and specific properties. In a homogeneous mixture, the different parts cannot be seen. A solution of salt dissolved in water is an example of a homogeneous mixture. When the salt dissolves, it will spread evenly through the water so that all parts of the solution are the same, and you can no longer see the salt as being separate from the water. Think also of a powdered drink that you mix with water. Provided you give the container a good shake after you have added the powder to the water, the drink will have the same sweet taste for anyone who drinks it, it won't matter whether they take a sip from the top or from the bottom. The air we breathe is another example of a homogeneous mixture since it is made up of different gases which are in a constant ratio, and which can't be distinguished from each other.

Definition: Homogeneous mixture
A homogeneous mixture is one that is uniform, and where the different components of the mixture cannot be seen.

An **alloy** is a homogeneous mixture of two or more elements, at least one of which is a metal, where the resulting material has metallic properties. Alloys are usually made to improve on the properties of the elements that make them up. Steel for example, is much stronger than iron, which is its main component.

1.1.3 Separating mixtures

Sometimes it is important to be able to separate a mixture. There are lots of different ways to do this. These are some examples:

- *Filtration*

  A piece of filter paper in a funnel can be used to separate a mixture of sand and water.

- *Heating / evaporation*

  Sometimes, heating a solution causes the water to evaporate, leaving the other part of the mixture behind. You can try this using a salt solution.

- *Centrifugation*

  This is a laboratory process which uses the centrifugal force of spinning objects to separate out the heavier substances from a mixture. This process is used to separate the cells and plasma in blood. When the test tubes that hold the blood are spun round in the machine, the heavier cells sink to the bottom of the test tube. Can you think of a reason why it might be important to have a way of separating blood in this way?

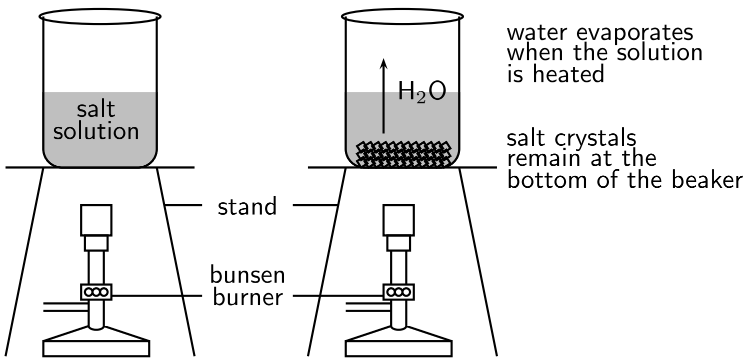
- *Dialysis*

  This is an interesting way of separating a mixture because it can be used in some important applications. Dialysis works using a process called *diffusion*. Diffusion takes place when one substance in a mixture moves from an area where it has a high concentration to an area where its concentration is lower. This movement takes place across a semi-permeable membrane. A semi-permeable membrane is a barrier that lets some things move across it, but not others. This process is very important for people whose kidneys are not functioning properly, an illness called *renal failure*.

Normally, healthy kidneys remove waste products from the blood. When a person has renal failure, their kidneys cannot do this any more, and this can be life-threatening. Using dialysis, the blood of the patient flows on one side of a semi-permeable membrane. On the other side there will be a fluid that has no waste products but lots of other important substances such as potassium ions (K^+) that the person will need. Waste products from the blood diffuse from where their concentration is high (i.e. in the person's blood) into the 'clean' fluid on the other side of the membrane. The potassium ions will move in the opposite direction from the fluid into the blood. Through this process, waste products are taken out of the blood so that the person stays healthy.

Activity :: Investigation : The separation of a salt solution
Aim:
To demonstrate that a homogeneous salt solution can be separated using physical methods.
Apparatus:
glass beaker, salt, water, retort stand, bunsen burner.
Method:

1. Pour a small amount of water (about 20 ml) into a beaker.
2. Measure a teaspoon of salt and pour this into the water.
3. Stir until the salt dissolves completely. This is now called a *salt solution*. This salt solution is a homogeneous mixture.
4. Place the beaker on a retort stand over a bunsen burner and heat gently. You should increase the heat until the water almost boils.
5. Watch the beaker until all the water has evaporated. What do you see in the beaker?



Results:
The water evaporates from the beaker and tiny grains of salt remain at the bottom.
Conclusion:
The sodium chloride solution, which was a homogeneous mixture of salt and water, has been separated using heating and evaporation.

Activity :: Discussion : Separating mixtures
Work in groups of 3-4
Imagine that you have been given a container which holds a mixture of sand, iron filings (small pieces of iron metal), salt and small stones of different sizes. Is this a homogeneous or a heterogeneous mixture? In your group, discuss how you would go about separating this mixture into the four materials that it contains.

?

Exercise: Mixtures

1. Which of the following subtances are *mixtures*?
   (a) tap water
   (b) brass (an alloy of copper and zinc)
   (c) concrete
   (d) aluminium
   (e) Coca cola
   (f) distilled water
2. In each of the examples above, say whether the mixture is homogeneous or heterogeneous

1.2 Pure Substances: Elements and Compounds

Any material that is not a mixture, is called a **pure substance**. Pure substances include **elements** and **compounds**. It is much more difficult to break down pure substances into their parts, and complex chemical methods are needed to do this.

1.2.1 Elements

An **element** is a chemical substance that can't be divided or changed into other chemical substances by any ordinary chemical means. The smallest unit of an element is the **atom**.

Definition: Element
An element is a substance that cannot be broken down into other substances through chemical means.

There are 109 known elements. Most of these are natural, but some are man-made. The elements we know are represented in the **Periodic Table of the Elements**, where each element is abbreviated to a **chemical symbol**. Examples of elements are magnesium (Mg), hydrogen (H), oxygen (O) and carbon (C). On the Periodic Table you will notice that some of the abbreviations do not seem to match the elements they represent. The element iron, for example, has the chemical formula Fe. This is because the elements were originally given Latin names. Iron has the abbreviation Fe because its Latin name is 'ferrum'. In the same way, sodium's Latin name is 'natrium' (Na) and gold's is 'aurum' (Au).

1.2.2 Compounds

A **compound** is a chemical substance that forms when two or more elements combine in a fixed ratio. Water (H_2O), for example, is a compound that is made up of two hydrogen atoms for every one oxygen atom. Sodium chloride (NaCl) is a compound made up of one sodium atom for every chlorine atom. An important characteristic of a compound is that it has a **chemical formula**, which describes the ratio in which the atoms of each element in the compound occur.

Definition: Compound
A substance made up of two or more elements that are joined together in a fixed ratio.

Diagram 1.2 might help you to understand the difference between the terms *element*, *mixture* and *compound*. Iron (Fe) and sulfur (S) are two elements. When they are added together, they

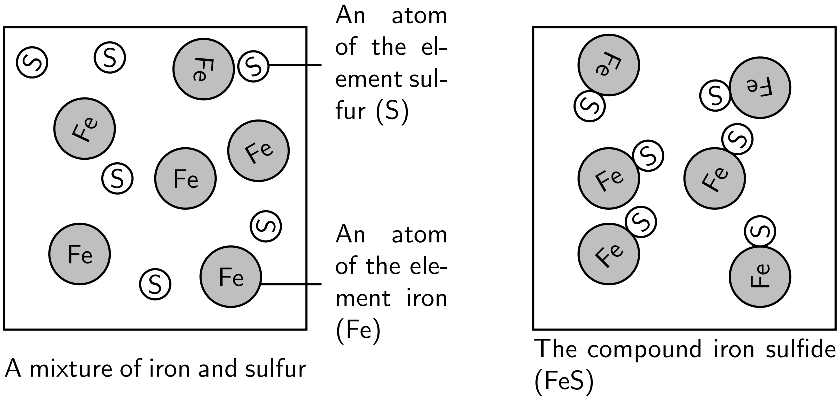


Figure 1.2: Understanding the difference between a mixture and a compound

form a *mixture* or iron and sulfur. The iron and sulfur are not joined together. However, if the mixture is heated, a new *compound* is formed, which is called iron sulfide (FeS). In this compound, the iron and sulfur are joined to each other in a ratio of 1:1. In other words, one atom of iron is joined to one atom of sulfur in the compound iron sulfide.

Exercise: Elements, mixtures and compounds

1. In the following table, tick whether each of the substances listed is a *mixture* or a *pure substance*. If it is a mixture, also say whether it is a homogeneous or heterogeneous mixture.

Substance	**Mixture or pure**	**Homogeneous or heterogeneous mixture**
fizzy colddrink		
steel		
oxygen		
iron filings		
smoke		
limestone ($CaCO_3$)		

2. In each of the following cases, say whether the substance is an element, a mixture or a compound.
   (a) Cu
   (b) iron and sulfur
   (c) Al
   (d) H_2SO_4
   (e) SO_3

1.3 Giving names and formulae to substances

It is easy to describe elements and mixtures. But how are compounds named? In the example of iron sulfide that was used earlier, which element is named first, and which 'ending' is given to the compound name (in this case, the ending is -ide)?

The following are some guidelines for naming compounds:

1. The compound name will always include the **names of the elements** that are part of it.
   - A compound of **iron** (Fe) and *sulfur* (S) is **iron** *sulf*ide (FeS)
   - A compound of **potassium** (K) and *bromine* (S) is **potassium** *brom*ide (KBr)
   - A compound of **sodium** (Na) and *chlorine* (Cl) is **sodium** *chlor*ide (NaCl)
2. In a compound, the element that is to the left and lower down on the Periodic Table, is used *first* when naming the compound. In the example of NaCl, sodium is a group 1 element on the left hand side of the table, while chlorine is in group 7 on the right of the table. Sodium therefore comes first in the compound name. The same is true for FeS and KBr.
3. The **symbols** of the elements can be used to represent compounds e.g. FeS, NaCl and KBr. These are called **chemical formulae**. In these three examples, the ratio of the elements in each compound is 1:1. So, for FeS, there is one atom of iron for every atom of sulfur in the compound.
4. A compound may contain **compound ions**. Some of the more common compound ions and their names are shown below.

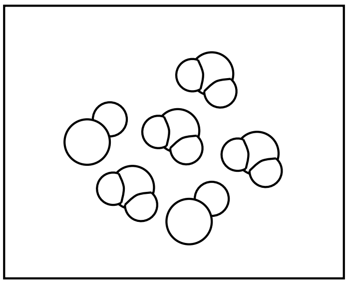
Name of compound ion	formula
Carbonate	$CO_3{}^{2-}$
sulphate	$SO_4{}^{2-}$
Hydroxide	OH^-
Ammonium	$NH_4{}^+$
Nitrate	$NO_3{}^-$
Hydrogen carbonate	$HCO_3{}^-$
Phosphate	$PO_4{}^{3-}$
Chlorate	$ClO_3{}^-$
Cyanide	CN^-
Chromate	$CrO_4{}^{2-}$
Permanganate	$MnO_4{}^-$

5. When there are only two elements in the compound, the compound is often given a **suffix** (ending) of -ide. You would have seen this in some of the examples we have used so far. When a non-metal is combined with oxygen to form a negative ion (anion) which then combines with a positive ion (cation) from hydrogen or a metal, then the suffix of the name will be ...ate or ...ite. NO_3^- for example, is a negative ion, which may combine with a cation such as hydrogen (HNO_3) or a metal like potassium (KNO_3). The NO_3^- anion has the name nitr**ate**. SO_3 in a formula is sulph**ite**, e.g. sodium sulphite (Na_2SO_3). SO_4 is sulph**ate** and PO_4 is phosph**ate**.
6. **Prefixes** can be used to describe the ratio of the elements that are in the compound. You should know the following prefixes: 'mono' (one), 'di' (two) and 'tri' (three).
   - CO (carbon monoxide) - There is one atom of oxygen for every one atom of carbon
   - NO_2 (nitrogen dioxide) - There are two atoms of oxygen for every one atom of nitrogen
   - SO_3 (sulfur trioxide) - There are three atoms of oxygen for every one atom of sulfur

Important:
When numbers are written as 'subscripts' in compounds (i.e. they are written below the element symbol), this tells us how many atoms of that element there are in relation to other elements in the compound. For example in nitrogen dioxide (NO_2) there are two oxygen atoms for every one atom of nitrogen. In sulfur trioxide (SO_3), there are three oxygen atoms for every one atom of sulfur in the compound. Later, when we start looking at chemical equations, you will notice that sometimes there are numbers *before* the compound name. For example, $2H_2O$ means that there are two molecules of water, and that in each molecule there are two hydrogen atoms for every one oxygen atom.

Exercise: Naming compounds

1. The formula for calcium carbonate is $CaCO_3$.
   (a) Is calcium carbonate a mixture or a compound? Give a reason for your answer.
   (b) What is the ratio of Ca:C:O atoms in the formula?
2. Give the name of each of the following substances.
   (a) KBr
   (b) HCl
   (c) $KMnO_4$
   (d) NO_2
   (e) NH_4OH
   (f) Na_2SO_4
3. Give the chemical formula for each of the following compounds.
   (a) potassium nitrate
   (b) sodium iodide
   (c) barium sulphate
   (d) nitrogen dioxide
   (e) sodium monosulphate
4. Refer to the diagram below, showing sodium chloride and water, and then answer the questions that follow.



   (a) What is the chemical formula for water?
   (b) What is the chemical formula for sodium chloride?
   (c) Label the water and sodium chloride in the diagram.
   (d) Which of the following statements most accurately describes the picture?
      i. The picture shows a mixture of an element and a compound
      ii. The picture shows a mixture of two compounds
      iii. The picture shows two compounds that have been chemically bonded to each other
5. What is the formula of this molecule?

$$\begin{array}{ccccccccc} & & H & & H & & & & \\ & & | & & | & & & & \\ H & — & C & — & C & — & O & — & H \\ & & | & & | & & & & \\ & & H & & H & & & & \end{array}$$

   A C_6H_2O
   B C_2H_6O
   C 2C6HO
   D $_2CH_6O$

1.4 Metals, Semi-metals and Non-metals

The elements in the Periodic Table can also be divided according to whether they are **metals**, **semi-metals** or **non-metals**. On the right hand side of the Periodic Table is a dark 'zigzag' line. This line separates all the elements that are metals from those that are non-metals. Metals are found on the left of the line, and non-metals are those on the right. Metals, semi-metals and non-metals all have their own specific properties.

1.4.1 Metals

Examples of metals include copper (Cu), zinc (Zn), gold (Au) and silver (Ag). On the Periodic Table, the metals are on the left of the zig-zag line. There are a large number of elements that are metals. The following are some of the properties of metals:

- *Thermal conductors*

  Metals are good conductors of heat and are therefore used in cooking utensils such as pots and pans.

- *Electrical conductors*

  Metals are good conductors of electricity, and are therefore used in electrical conducting wires.

- *Shiny metallic lustre*

  Metals have a characteristic shiny appearance and are often used to make jewellery.

- *Malleable*

  This means that they can be bent into shape without breaking.

- *Ductile*

  Metals can stretched into thin wires such as copper, which can then be used to conduct electricity.

- *Melting point*

  Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged.

You can see how the properties of metals make them very useful in certain applications.

Activity :: Group Work : Looking at metals

1. Collect a number of metal items from your home or school. Some examples are listed below:
   - hammer
   - electrical wiring
   - cooking pots
   - jewellery
   - burglar bars
   - coins
2. In groups of 3-4, combine your collection of metal objects.
3. What is the function of each of these objects?
4. Discuss why you think metal was used to make each object. You should consider the properties of metals when you answer this question.

1.4.2 Non-metals

In contrast to metals, non-metals are poor thermal conductors, good electrical insulators (meaning that they do *not* conduct electrical charge) and are neither malleable nor ductile. The non-metals are found on the right hand side of the Periodic Table, and include elements such as sulfur (S), phosphorus (P), nitrogen (N) and oxygen (O).

1.4.3 Semi-metals

Semi-metals have mostly non-metallic properties. One of their distinguishing characteristics is that their conductivity increases as their temperature increases. This is the opposite of what happens in metals. The semi-metals include elements such as silicon (Si) and germanium (Ge). Notice where these elements are positioned in the Periodic Table.

1.5 Electrical conductors, semi-conductors and insulators

An **electrical conductor** is a substance that allows an electrical current to pass through it. Many electrical conductors are metals, but non-metals can also be good conductors. *Copper* is one of the best electrical conductors, and this is why it is used to make conducting wire. In reality, *silver* actually has an even higher electrical conductivity than copper, but because silver is so expensive, it is not practical to use it for electrical wiring because such large amounts are needed. In the overhead power lines that we see above us, *aluminium* is used. The aluminium usually surrounds a steel core which adds tensile strength to the metal so that it doesn't break when it is stretched across distances. Occasionally gold is used to make wire, not because it is a particularly good conductor, but because it is very resistant to surface corrosion. *Corrosion* is when a material starts to deteriorate at the surface because of its reactions with the surroundings, for example oxygen and water in the air.

An **insulator** is a non-conducting material that does not carry any charge. Examples of insulators would be plastic and wood. Do you understand now why electrical wires are normally covered with plastic insulation? **Semi-conductors** behave like insulators when they are cold, and like conductors when they are hot. The elements silicon and germanium are examples of semiconductors.

Definition: Conductors and insulators
A conductor allows the easy movement or flow of something such as heat or electrical charge through it. Insulators are the opposite to conductors because they *inhibit* or reduce the flow of heat, electrical charge, sound etc through them.

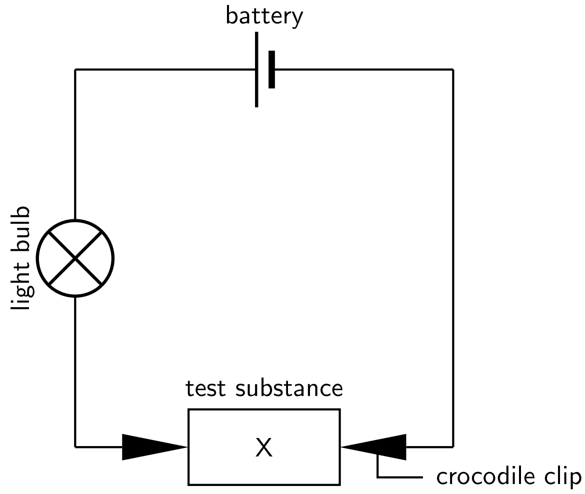
Activity :: Experiment : Electrical conductivity

Aim:

To investigate the electrical conductivity of a number of substances

Apparatus:

- two or three cells
- light bulb
- crocodile clips
- wire leads
- a selection of test substances (e.g. a piece of plastic, aluminium can, metal pencil sharpener, metal magnet, wood, chalk).



Method:

1. Set up the circuit as shown above, so that the test substance is held between the two crocodile clips. The wire leads should be connected to the cells and the light bulb should also be connected into the circuit.
2. Place the test substances one by one between the crocodile clips and see what happens to the light bulb.

Results:

Record your results in the table below:

Test substance	Metal/non-metal	Does bulb glow?	Conductor or insulator

Conclusions:

In the substances that were tested, the metals were able to conduct electricity and the non-metals were not. Metals are good electrical conductors and non-metals are not.

1.6 Thermal Conductors and Insulators

A **thermal conductor** is a material that allows energy in the form of heat, to be transferred within the material, without any movement of the material itself. An easy way to understand this concept is through a simple demonstration.

Activity :: Demonstration : Thermal conductivity
Aim:
To demonstrate the ability of different substances to conduct heat.
Apparatus:

You will need two cups (made from the same material e.g. plastic); a metal spoon and a plastic spoon.

Method:

- Pour boiling water into the two cups so that they are about half full.
- At the same time, place a metal spoon into one cup and a plastic spoon in the other.
- Note which spoon heats up more quickly

Results:

The metal spoon heats up more quickly than the plastic spoon. In other words, the metal conducts heat well, but the plastic does not.

Conclusion:

Metal is a good thermal conductor, while plastic is a poor thermal conductor. This explains why cooking pots are metal, but their handles are often plastic or wooden. The pot itself must be metal so that heat from the cooking surface can heat up the pot to cook the food inside it, but the handle is made from a poor thermal conductor so that the heat does not burn the hand of the person who is cooking.

An **insulator** is a material that does not allow a transfer of electricity or energy. Materials that are poor thermal conductors can also be described as being good insulators.

Interesting Fact

Water is a better thermal conductor than air and conducts heat away from the body about 20 times more efficiently than air. A person who is not wearing a wetsuit, will lose heat very quickly to the water around them and can be vulnerable to hypothermia. Wetsuits help to preserve body heat by trapping a layer of water against the skin. This water is then warmed by body heat and acts as an insulator. Wetsuits are made out of closed-cell, foam neoprene. Neoprene is a synthetic rubber that contains small bubbles of nitrogen gas when made for use as wetsuit material. Nitrogen gas has very low thermal conductivity, so it does not allow heat from the body (or the water trapped between the body and the wetsuit) to be lost to the water outside of the wetsuit. In this way a person in a wetsuit is able to keep their body temperature much higher than they would otherwise.

Activity :: Investigation : A closer look at thermal conductivity

Look at the table below, which shows the thermal conductivity of a number of different materials, and then answer the questions that follow. The higher the number in the second column, the better the material is at conducting heat (i.e. it is a good thermal conductor). Remember that a material that conducts heat efficiently, will also lose heat more quickly than an insulating material.

Material	Thermal Conductivity (W/m/K)
Silver	429
Stainless steel	16
Standard glass	1.05
Concrete	0.9 - 2
Red brick	0.69
Water	0.58
Snow	0.5 - 0.25
Wood	0.04 - 0.12
Polystyrene	0.03
Air	0.024

Use this information to answer the following questions:

1. Name two materials that are good thermal conductors.
2. Name two materials that are good insulators.
3. Explain why:
   (a) cooler boxes are often made of polystyrene
   (b) homes that are made from wood need less internal heating during the winter months.
   (c) igloos (homes made from snow) are so good at maintaining warm temperatures, even in freezing conditions.

It is a known fact that well-insulated buildings need less energy for heating than do buildings that have no insulation. Two building materials that are being used more and more worldwide, are **mineral wool** and **polystyrene**. Mineral wool is a good insulator because it holds air still in the matrix of the wool so that heat is not lost. Since air is a poor conductor and a good insulator, this helps to keep energy within the building. Polystyrene is also a good insulator and is able to keep cool things cool and hot things hot! It has the added advantage of being resistant to moisture, mould and mildew.

Remember that concepts such as conductivity and insulation are not only relevant in the building, industrial and home environments. Think for example of the layer of blubber or fat that we find in animals. In very cold environments, fat and blubber not only provide protection, but also act as an insulator to help the animal to keep its body temperature at the right level. This is known as *thermoregulation*.

1.7 Magnetic and Non-magnetic Materials

We have now looked at a number of ways in which matter can be grouped, such as into metals, semi-metals and non-metals; electrical conductors and insulators, and thermal conductors and insulators. One way in which we can further group metals, is to divide them into those that are **magnetic** and those that are **non-magnetic.**

Definition: Magnetism
Magnetism is one of the phenomena by which materials exert attractive or repulsive forces on other materials.

A metal is said to be **ferromagnetic** if it can be magnetised (i.e. made into a magnet). If you hold a magnet very close to a metal object, it may happen that its own electrical field will be induced and the object becomes magnetic. Some metals keep their magnetism for longer than others. Look at iron and steel for example. Iron loses its magnetism quite quickly if it is taken away from the magnet. Steel on the other hand will stay magnetic for a longer time. Steel is often used to make permanent magnets that can be used for a variety of purposes.

Magnets are used to sort the metals in a scrap yard, in compasses to find direction, in the magnetic strips of video tapes and ATM cards where information must be stored, in computers and TV's, as well as in generators and electric motors.

Activity :: Investigation : Magnetism

You can test whether an object is magnetic or not by holding another magnet close to it. If the object is attracted to the magnet, then it too is magnetic.

Find some objects in your classroom or your home and test whether they are magnetic or not. Then complete the table below:

Object	**Magnetic or non-magnetic**

Activity :: Group Discussion : Properties of materials

In groups of 4-5, discuss how our knowledge of the properties of materials has allowed society to:

- develop advanced computer technology
- provide homes with electricity
- find ways to conserve energy

1.8 Summary

- All the objects and substances that we see in the world are made of **matter**.
- This matter can be classified according to whether it is a **mixture** or a **pure substance**.
- A **mixture** is a combination of one or more substances that are not chemically bonded to each other. Examples of mixtures are air (a mixture of different gases) and blood (a mixture of cells, platelets and plasma).
- The main **characteristics** of mixtures are that the substances that make them up are not in a fixed ratio, they keep their individual properties and they can be separated from each other using mechanical means.

- A **heterogeneous mixture** is non-uniform and the different parts of the mixture can be seen. An example would be a mixture of sand and salt.
- A **homogeneous mixture** is uniform, and the different components of the mixture can't be seen. An example would be a salt solution. A salt solution is a mixture of salt and water. The salt dissolves in the water, meaning that you can't see the individual salt particles. They are interspersed between the water molecules. Another example is a metal **alloy** such as steel.
- Mixtures can be **separated** using a number of methods such as filtration, heating, evaporation, centrifugation and dialysis.
- Pure substances can be further divided into **elements** and **compounds**.
- An **element** is a substance that can't be broken down into simpler substances through chemical means.
- All the elements are recorded in the **Periodic Table of the Elements**. Each element has its own chemical symbol. Examples are iron (Fe), sulfur (S), calcium (Ca), magnesium (Mg) and fluorine (F).
- A **compound** is a substance that is made up of two or more elements that are chemically bonded to each other in a fixed ratio. Examples of compounds are sodium chloride (NaCl), iron sulfide (FeS), calcium carbonate ($CaCO_3$) and water (H_2O).
- When **naming compounds** and writing their **chemical formula**, it is important to know the elements that are in the compound, how many atoms of each of these elements will combine in the compound and where the elements are in the Periodic Table. A number of rules can then be followed to name the compound.
- Another way of classifying matter is into **metals** (e.g. iron, gold, copper), **semi-metals** (e.g. silicon and germanium) and **non-metals** (e.g. sulfur, phosphorus and nitrogen).
- **Metals** are good electrical and thermal conductors, they have a shiny lustre, they are malleable and ductile, and they have a high melting point. These properties make metals very useful in electrical wires, cooking utensils, jewellery and many other applications.
- A further way of classifying matter is into **electrical conductors**, **semi-conductors** and **insulators**.
- An **electrical conductor** allows an electrical current to pass through it. Most metals are good electrical conductors.
- An **electrical insulator** is not able to carry an electrical current. Examples are plastic, wood, cotton material and ceramic.
- Materials may also be classified as **thermal conductors** or **thermal insulators** depending on whether or not they are able to conduct heat.
- Materials may also be either **magnetic** or **non-magnetic**.

Exercise: Summary

1. For each of the following **multiple choice** questions, choose *one* correct answer from the list provided.
   A Which of the following can be classified as a mixture:
      i. sugar
      ii. table salt
      iii. air
      iv. Iron
   B An element can be defined as:

   i. A substance that cannot be separated into two or more substances by ordinary chemical (or physical) means
   ii. A substance with constant composition
   iii. A substance that contains two or more substances, in definite proportion by weight
   iv. A uniform substance

2. Classify each of the following substances as an *element*, a *compound*, a *solution*(homogeneous mixture), or a *heterogeneous mixture*: salt, pure water, soil, salt water, pure air, carbon dioxide, gold and bronze
3. Look at the table below. In the first column (A) is a list of substances. In the second column (B) is a description of the group that each of these substances belongs in. Match up the *substance* in Column A with the *description* in Column B.

Column A	**Column B**
iron	a compound containing 2 elements
H_2S	a heterogeneous mixture
sugar solution	a metal alloy
sand and stones	an element
steel	a homogeneous mixture

4. You are given a test tube that contains a mixture of iron filings and sulfur. You are asked to weigh the amount of iron in the sample.
   a Suggest one method that you could use to separate the iron filings from the sulfur.
   b What property of metals allows you to do this?
5. Given the following descriptions, write the chemical formula for each of the following substances:
   a silver metal
   b a compound that contains only potassium and bromine
   c a gas that contains the elements carbon and oxygen in a ratio of 1:2
6. Give the names of each of the following compounds:
   a NaBr
   b $BaSO_4$
   c SO_2
7. For each of the following materials, say what properties of the material make it important in carrying out its particular function.
   a **tar** on roads
   b **iron** burglar bars
   c **plastic** furniture
   d **metal** jewellery
   e **clay** for building
   f **cotton** clothing

Chapter 2

What are the objects around us made of? - Grade 10

2.1 Introduction: The atom as the building block of matter

We have now seen that different materials have different properties. Some materials are metals and some are non-metals; some are electrical or thermal conductors, while others are not. Depending on the properties of these materials, they can be used in lots of useful applications. But what is it exactly that makes up these materials? In other words, if we were to break down a material into the parts that make it up, what would we find? And how is it that a material's microscopic structure is able to give it all these different properties?

The answer lies in the smallest building block of matter: the **atom**. It is the *type* of atoms, and the way in which they are *arranged* in a material, that affects the properties of that substance.

It is not often that substances are found in atomic form. Normally, atoms are bonded to other atoms to form **compounds** or **molecules**. It is only in the *noble gases* (e.g. helium, neon and argon) that atoms are found individually and are not bonded to other atoms. We will look at the reasons for this in a later chapter.

2.2 Molecules

Definition: Molecule
A molecule is a group of two or more atoms that are attracted to each other by relatively strong forces or bonds

Almost everything around us is made up of molecules. *Water* is made up of molecules, each of which has two hydrogen atoms joined to one oxygen atom. *Oxygen* is a molecule that is made up of two oxygen atoms that are joined to one another. Even the food that we eat is made up of molecules that contain atoms of elements such as carbon, hydrogen and oxygen that are joined to one another in different ways. All of these are known as **small molecules** because there are only a few atoms in each molecule. **Giant molecules** are those where there may be millions of atoms per molecule. Examples of giant molecules are *diamonds*, which are made up of millions of carbon atoms bonded to each other, and *metals*, which are made up of millions of metal atoms bonded to each other.

2.2.1 Representing molecules

The structure of a molecule can be shown in many different ways. Sometimes it is easiest to show what a molecule looks like by using different types of **diagrams**, but at other times, we may decide to simply represent a molecule using its **chemical formula** or its written name.

1. **Using formulae to show the structure of a molecule**

   A **chemical formula** is an abbreviated (shortened) way of describing a molecule, or some other chemical substance. In chapter 1, we saw how chemical compounds can be represented using element symbols from the Periodic Table. A chemical formula can also tell us the *number* of atoms of each element that are in a molecule, and their *ratio* in that molecule.

   For example, the chemical formula for a molecule of carbon dioxide is:

   $$CO_2$$

   The formula above is called the **molecular formula** of that compound. The formula tells us that in one molecule of carbon dioxide, there is one atom of carbon and two atoms of oxygen. The ratio of carbon atoms to oxygen atoms is 1:2.

   Definition: Molecular formula
   A concise way of expressing information about the atoms that make up a particular chemical compound. The molecular formula gives the exact number of each type of atom in the molecule.

   A molecule of glucose has the molecular formula:

   $$C_6H_{12}O_6$$

   In each glucose molecule, there are six carbon atoms, twelve hydrogen atoms and six oxygen atoms. The ratio of carbon:hydrogen:oxygen is 6:12:6. We can simplify this ratio to write 1:2:1, or if we were to use the element symbols, the formula would be written as CH_2O. This is called the **empirical formula** of the molecule.

   Definition: Empirical formula
   This is a way of expressing the *relative* number of each type of atom in a chemical compound. In most cases, the empirical formula does not show the exact number of atoms, but rather the simplest *ratio* of the atoms in the compound.

   The empirical formula is useful when we want to write the formula for a *giant molecule*. Since giant molecules may consist of millions of atoms, it is impossible to say exactly how many atoms are in each molecule. It makes sense then to represent these molecules using their empirical formula. So, in the case of a metal such as copper, we would simply write Cu, or if we were to represent a molecule of sodium chloride, we would simply write NaCl.

   Chemical formulae therefore tell us something about the *types* of atoms that are in a molecule and the *ratio* in which these atoms occur in the molecule, but they don't give us any idea of what the molecule actually looks like, in other words its *shape*. Another useful way of representing molecules is to use diagrams.

   Another type of formula that can be used to describe a molecule is its **structural formula**. A structural formula uses a graphical representation to show a molecule's structure (figure 2.1).

2. **Using diagrams to show the structure of a molecule**

   Diagrams of molecules are very useful because they give us an idea of the *space* that is occupied by the molecule, and they also help us to picture how the atoms are arranged in the molecule. There are two types of diagrams that are commonly used:

(a) C_4H_{10} (b) C_2H_5

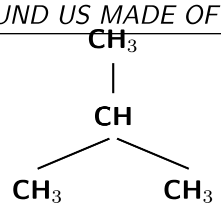


Figure 2.1: Diagram showing (a) the molecular, (b) the empirical and (c) the structural formula of isobutane

- *Ball and stick models*

  This is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them. The centres of the atoms (the balls) are connected by straight lines which represent the bonds between them. A simplified example is shown in figure 2.2.

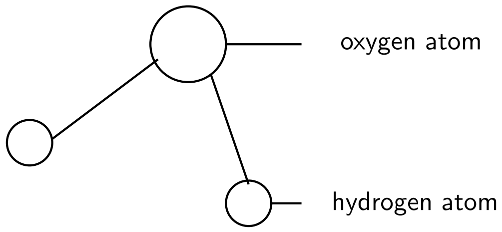


Figure 2.2: A ball and stick model of a water molecule

- *Space-filling model*

  This is also a 3-dimensional molecular model. The atoms are represented by multi-coloured spheres. Space-filling models of water and ammonia are shown in figures 2.3 and 2.4.

Figures 2.3 and 2.4 are some examples of **simple molecules** that are represented in different ways.

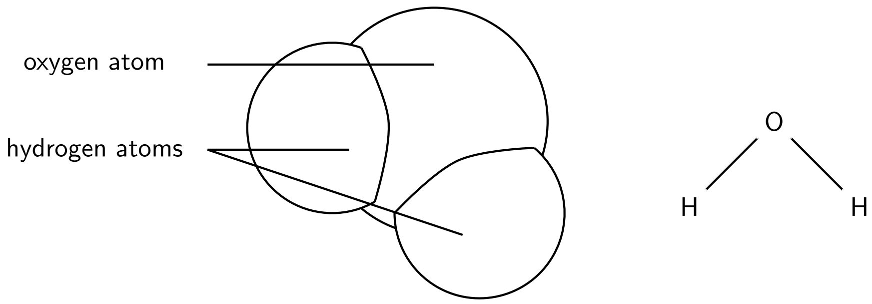


Figure 2.3: A space-filling model and structural formula of a water molecule. Each molecule is made up of two hydrogen atoms that are attached to one oxygen atom. This is a simple molecule.

Figure 2.5 shows the bonds between the carbon atoms in diamond, which is a **giant molecule**. Each carbon atom is joined to four others, and this pattern repeats itself until a complex *lattice* structure is formed. Each black ball in the diagram represents a carbon atom, and each line represents the bond between two carbon atoms.

Interesting Fact

Diamonds are most often thought of in terms of their use in the jewellery industry. However, about 80% of mined diamonds are unsuitable for use as gemstones and are therefore used in industry because of their strength and hardness. These

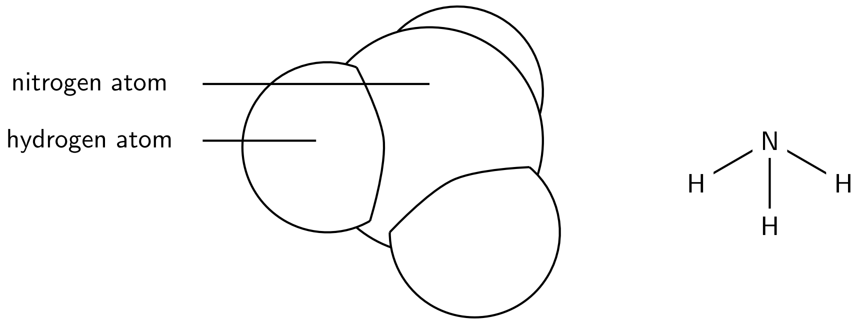


Figure 2.4: A space-filling model and structural formula of a molecule of ammonia. Each molecule is made up of one nitrogen atom and three hydrogen atoms. This is a simple molecule.

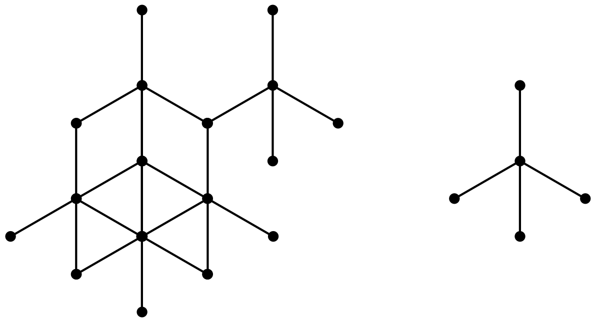


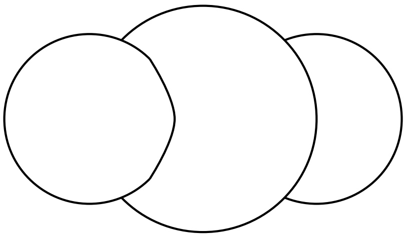
Figure 2.5: Diagrams showing the microscopic structure of diamond. The diagram on the left shows part of a diamond lattice, made up of numerous carbon atoms. The diagram on the right shows how each carbon atom in the lattice is joined to four others. This forms the basis of the lattice structure. Diamond is a giant molecule.

properties of diamonds are due to the strong covalent bonds betwene the carbon atoms in diamond. The most common uses for diamonds in industry are in cutting, drilling, grinding, and polishing.

?

Exercise: Atoms and molecules

1. In each of the following, say whether the chemical substance is made up of single atoms, simple molecules or giant molecules.
   (a) ammonia gas (NH_3)
   (b) zinc metal (Zn)
   (c) graphite (C)
   (d) nitric acid (HNO_3)
   (e) neon gas (Ne_2)
2. Refer to the diagram below and then answer the questions that follow:



O — C — O

(a) Identify the molecule.
(b) Write the molecular formula for the molecule.
(c) Is the molecule a simple or giant molecule?

3. Represent each of the following molecules using its *chemical formula*, *structural formula* and *ball and stick model*.
   (a) H_2
   (b) NH_3
   (c) sulfur dioxide

2.3 Intramolecular and intermolecular forces

When atoms join to form molecules, they are held together by **chemical bonds**. The type of bond, and the strength of the bond, depends on the atoms that are involved. These bonds are called **intramolecular forces** because they are bonding forces *inside* a molecule ('intra' means 'within' or 'inside'). Sometimes we simply call these intramolecular forces chemical bonds.

Definition: Intramolecular force
The force between the atoms of a molecule, which holds them together.

Examples of the types of chemical bonds that can exist between atoms inside a molecule are shown below. These will be looked at in more detail in chapter 4.

- *Covalent bond*

  Covalent bonds exist between non-metal atoms e.g. There are covalent bonds between the carbon and oxygen atoms in a molecule of carbon dioxide.

- *Ionic bond*

  Ionic bonds occur between non-metal and metal atoms e.g. There are ionic bonds between the sodium and chlorine atoms in a molecule of sodium chloride.

- *Metallic bond*

  Metallic bonds join metal atoms e.g. There are metallic bonds between copper atoms in a piece of copper metal.

Intermolecular forces are those bonds that hold *molecules* together. A glass of water for example, contains many molecules of water. These molecules are held together by intermolecular forces. The strength of the intermolecular forces is important because they affect properties such as *melting point* and *boiling point*. For example, the stronger the intermolecular forces, the higher the melting point and boiling point for that substance. The strength of the intermolecular forces increases as the size of the molecule increases.

Definition: Intermolecular force
A force between molecules, which holds them together.

Diagram 2.6 may help you to understand the difference between intramolecular forces and intermolecular forces.

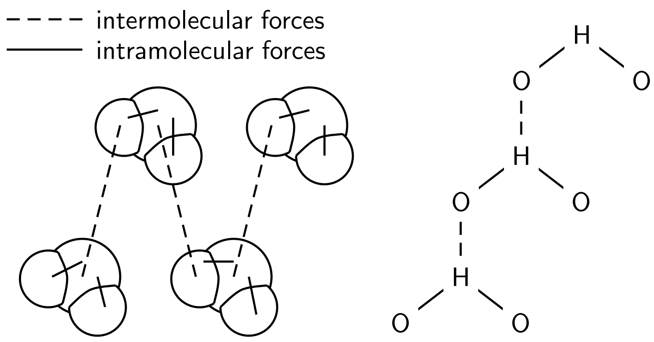


Figure 2.6: Two representations showing the intermolecular and intramolecular forces in water: space-filling model and structural formula.

It should be clearer now that there are two types of forces that hold matter together. In the case of water, there are intramolecular forces that hold the two hydrogen atoms to the oxygen atom *in each molecule of water*. There are also intramolecular forces *between each of these water molecules*. As mentioned earlier, these forces are very important because they affect many of the *properties of matter* such as boiling point, melting point and a number of other properties. Before we go on to look at some of these examples, it is important that we first take a look at the **Kinetic Theory of Matter**.

Exercise: Intramolecular and intermolecular forces

1. Using ammonia gas as an example...
   (a) Explain what is meant by an *intramolecular* force or *chemical bond*.
   (b) Explain what is meant by an *intermolecular* force.
2. Draw a diagram showing three molecules of carbon dioxide. On the diagram, show where the intramolecular and intermolecular forces are.
3. Why is it important to understand the types of forces that exist between atoms and between molecules? Try to use some practical examples in your answer.

2.4 The Kinetic Theory of Matter

The **kinetic theory of matter** is used to explain why matter exists in different *phases* (i.e. solid, liquid and gas), and how matter can change from one phase to the next. The kinetic theory of matter also helps us to understand other properties of matter. It is important to realise that what we will go on to describe is only a *theory*. It cannot be proved beyond doubt, but the fact that it helps us to explain our observations of changes in phase, and other properties of matter, suggests that it probably is more than just a theory.

Broadly, the Kinetic Theory of Matter says that:

- Matter is made up of **particles** that are constantly moving.
- All particles have **energy**, but the energy varies depending on whether the substance is a solid, liquid or gas. Solid particles have the least energy and gas particles have the most amount of energy.
- The **temperature** of a substance is a measure of the *average kinetic energy* of the particles.
- A change in **phase** may occur when the energy of the particles is changed.
- There are **spaces** between the particles of matter.
- There are **attractive forces** between particles and these become stronger as the particles move closer together. These attractive forces will either be intramolecular forces (if the particles are atoms) or intermolecular forces (if the particles are molecules). When the particles are extremely close, repulsive forces start to act.

Table 2.1 summarises the characteristics of the particles that are in each phase of matter.

Table 2.1: Table summarising the general features of solids, liquids and gases.

Property of matter	**Gas**	**Liquid**	**Gas**
Particles	Atoms or molecules	Atoms or molecules	Atoms or molecules
Energy and movement of particles	Particles have high energy and are constantly moving	Particles have less energy than in the gas phase	Low energy - particles vibrate around a fixed point
Spaces between particles	Large spaces because of high energy	Smaller spaces than in gases	Very little space between particles. Particles are tightly packed together
Attractive forces between particles	Weak forces because of the large distance between particles	Stronger forces than in gas. Liquids can be poured.	Very strong forces. Solids have a fixed volume.
Changes in phase	In general a gas becomes a liquid or solid when it is cooled. Particles have less energy and therefore move closer together so that the attractive forces become stronger, and the gas becomes a liquid or a solid	A liquid becomes a gas if its temperature is increased. It becomes a solid if its temperature decreases.	Solids become liquids or gases if their temperature is increased.

Let's look at an example that involves the three phases of water: ice (solid), water (liquid) and water vapour (gas).

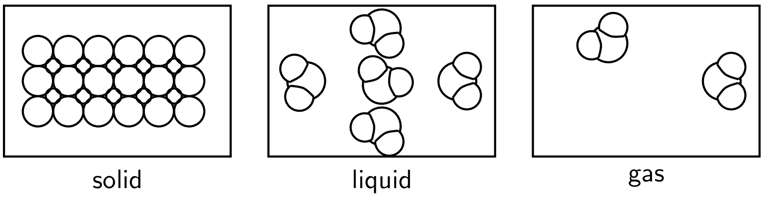


Figure 2.7: The three phases of matter

In a solid (e.g. ice), the water molecules have very little energy and can't move away from each other. The molecules are held close together in a regular pattern called a *lattice*. If the ice is heated, the energy of the molecules increases. This means that some of the water molecules are able to overcome the intermolecular forces that are holding them together, and the molecules move further apart to form *liquid water*. This is why liquid water is able to flow, because the molecules are more free to move than they were in the solid lattice. If the molecules are heated further, the liquid water will become water vapour, which is a gas. Gas particles have lots of energy and are far away from each other. That is why it is difficult to keep a gas in a specific area! The attractive forces between the particles are very weak and they are only loosely held together. Figure 2.8 shows the changes in phase that may occur in matter, and the names that describe these processes.

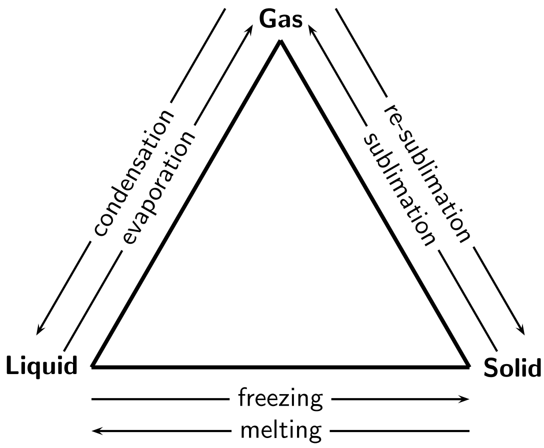


Figure 2.8: Changes in phase

2.5 The Properties of Matter

Let us now look at what we have learned about chemical bonds, intermolecular forces and the kinetic theory of matter, and see whether this can help us to understand some of the macroscopic properties of materials.

1. **Melting point**

Definition: Melting point
The temperature at which a *solid* changes its phase or state to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**.

In order for a solid to melt, the energy of the particles must increase enough to overcome the bonds that are holding the particles together. It makes sense then that a solid which is held together by strong bonds will have a *higher* melting point than one where the bonds are weak, because more energy (heat) is needed to break the bonds. In the examples we have looked at, metals, ionic solids and some atomic lattices (e.g. diamond) have high melting points, whereas the melting points for molecular solids and other atomic lattices (e.g. graphite) are much lower. Generally, the intermolecular forces between molecular solids are *weaker* than those between ionic and metallic solids.

2. **Boiling point**

Definition: Boiling point
The temperature at which a *liquid* changes its phase to become a *gas*.

When the temperature of a liquid increases, the average kinetic energy of the particles also increases, and they are able to overcome the bonding forces that are holding them in the liquid. When boiling point is reached, *evaporation* takes place and some particles in the liquid become a gas. In other words, the energy of the particles is too great for them to be held in a liquid anymore. The stronger the bonds within a liquid, the higher the boiling point needs to be in order to break these bonds. Metallic and ionic compounds have high boiling points while the boiling point for molecular liquids is lower.

The data in table 2.2 below may help you to understand some of the concepts we have explained. Not all of the substances in the table are solids at room temperature, so for now, let's just focus on the *boiling points* for each of these substances. Of the substances listed, ethanol has the weakest intermolecular forces, and sodium chloride and mercury have the strongest. What do you notice?

Substance	**Melting point (^0C)**	**Boiling point (^0C)**
Ethanol (C_2H_6O)	-114,3	78,4
Water	0	100
Mercury	-38,83	356,73
Sodium chloride	801	1465

Table 2.2: The melting and boiling points for a number of substances

You will have seen that substances such as ethanol, with relatively weak intermolecular forces, have the lowest boiling point, while substances with stronger intermolecular forces such as sodium chloride and mercury, must be heated much more if the particles are to have enough energy to overcome the forces that are holding them together in the liquid or solid phase.

?

Exercise: Forces and boiling point

The table below gives the molecular formula and the boiling point for a number of organic compounds called *alkanes*. Refer to the table and then answer the questions that follow.

Organic compound	**Molecular formula**	**Boiling point (^{0}C)**
Methane	CH_2	-161.6
Ethane	C_2H_6	-88.6
Propane	C_3H_8	-45
Butane	C_4H_{10}	-0.5
Pentane	C_5H_{12}	36.1
Hexane	C_6H_{14}	69
Heptane	C_7H_{16}	98.42
Octane	C_8H_{18}	125.52

Data from: http://www.wikipedia.com

(a) Draw a graph to show the relationship between the number of carbon atoms in each alkane, and its boiling point (Number of carbon atoms will go on the x-axis and boiling point on the y-axis).

(b) Describe what you see.

(c) Suggest a reason for what you have observed.

(d) Why was it enough for us to use 'number of carbon atoms' as a measure of the molecular weight of the molecules?

3. **Density and viscosity**

Density is a measure of the mass of a substance per unit volume. The density of a solid is generally higher than that of a liquid because the particles are hold much more closely together and therefore there are more particles packed together in a particular volume. In other words, there is a greater mass of the substance in a particular volume. In general, density increases as the strength of the intermolecular forces increases. **Viscosity** is a measure of how resistant a liquid is to changing its form. Viscosity is also sometimes described as the 'thickness' of a fluid. Think for example of syrup and how slowly it pours from one container into another. Now compare this to how easy it is to pour water. The viscosity of syrup is greater than the viscosity of water. Once again, the stronger the intermolecular forces in the liquid, the greater its viscosity.

It should be clear now that we can explain a lot of the **macroscopic properties** of matter (i.e. the characteristics we can see or observe) by understanding their **microscopic structure** and the way in which the atoms and molecules that make up matter are held together.

Activity :: Investigation : Determining the density of liquids:

Density is a very important property because it helps us to identify different materials. Every material, depending on the elements that make it up, and the arrangement of its atoms, will have a different density.

The equation for density is:

$$\text{Density} = \text{Mass}/\text{Volume}$$

Discussion questions:

To calculate the density of liquids and solids, we need to be able to first determine their mass and volume. As a group, think about the following questions:

- How would you determine the mass of a liquid?
- How would you determine the volume of an irregular solid?

Apparatus:

Laboratory mass balance, 10 ml and 100 ml graduated cylinders, thread, distilled water, two different liquids.

Method:

Determine the density of the distilled water and two liquids as follows:

1. Measure and record the mass of a 10 ml graduated cyclinder.
2. Pour an amount of distilled water into the cylinder.
3. Measure and record the combined mass of the water and cylinder.
4. Record the volume of distilled water in the cylinder
5. Empty, clean and dry the graduated cylinder.
6. Repeat the above steps for the other two liquids you have.
7. Complete the table below.

Liquid	**Mass (g)**	**Volume (ml)**	**Density (g/ml)**
Distilled water			
Liquid 1			
Liquid 2			

Activity :: Investigation : Determining the density of irregular solids:

Apparatus:

Use the same materials and equipment as before (for the liquids). Also find a number of solids that have an irregular shape.

Method:

Determine the density of irregular solids as follows:

1. Measure and record the mass of one of the irregular solids.
2. Tie a piece of thread around the solid.
3. Pour some water into a 100 ml graduated cylinder and record the volume.
4. Gently lower the solid into the water, keeping hold of the thread. Record the combined volume of the solid and the water.
5. Dtermine the volume of the solid by subtracting the combined volume from the original volume of the water only.
6. Repeat these steps for the second object.
7. Complete the table below.

Solid	**Mass (g)**	**Volume (ml)**	**Density (g/ml)**
Solid 1			
Solid 2			
Solid 3			

2.6 Summary

- The smallest unit of matter is the **atom**. Atoms can combine to form **molecules**.
- A **molecule** is a group of two or more atoms that are attracted to each other by chemical bonds.
- A **small molecule** consists of a few atoms per molecule. A **giant molecule** consists of millions of atoms per molecule, for example metals and diamonds.
- The structure of a molecule can be represented in a number of ways.
- The **chemical formula** of a molecule is an abbreviated way of showing a molecule, using the symbols for the elements in the molecule. There are two types of chemical formulae: molecular and empirical formula.
- The **molecular formula** of a molecule gives the exact number of atoms of each element that are in the molecule.
- The **empirical formula** of a molecule gives the relative number of atoms of each element in the molecule.
- Molecules can also be represented using **diagrams**.
- A **ball and stick** diagram is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them.
- A **space-filling model** is also a 3-dimensional molecular model. The atoms are represented by multi-coloured spheres.
- In a molecule, atoms are held together by **chemical bonds** or **intramolecular forces**. Covalent bonds, ionic bonds and metallic bonds are examples of chemical bonds.
- A **covalent bond** exists between non-metal atoms. An **ionic bond** exists between non-metal and metal atoms, and a **metallic bond** exists between metal atoms.
- **Intermolecular forces** are the bonds that hold *molecules* together.
- The **kinetic theory of matter** attempts to explain the behaviour of matter in different phases.
- The theory says that all matter is composed of **particles** which have a certain amount of **energy** which allows them to **move** at different speeds depending on the temperature (energy). There are **spaces** between the particles, and also **attractive forces** between particles when they come close together.

- Understanding chemical bonds, intermolecular forces and the kinetic theory of matter, can help to explain many of the **macroscopic properties** of matter.
- **Melting point** is the temperature at which a *solid* changes its phase to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the melting point will be.
- **Boiling point** is the temperature at which a liquid changes phase to become a gas. The stronger the chemical bonds and intermolecular forces in a substance, the higher the boiling point will be.
- **Density** is a measure of the mass of a substance per unit volume.
- **Viscosity** is a measure of how resistant a liquid is to changing its form.

?

Exercise: Summary exercise

1. Give one word or term for each of the following descriptions.
   (a) The property that determines how easily a liquid flows.
   (b) The change in phase from liquid to gas.
   (c) A composition of two or more atoms that act as a unit.
   (d) Chemical formula that gives the relative number of atoms of each element that are in a molecule.
2. For each of the following questions, choose the one correct answer from the list provided.
   A Ammonia, an ingredient in household cleaners, can be broken down to form one part nitrogen (N) and three parts hydrogen (H). This means that ammonia...
      i. is a colourless gas
      ii. is not a compound
      iii. cannot be an element
      iv. has the formula N_3H
   B If one substance A has a melting point that is *lower* than the melting point of substance B, this suggests that...
      i. A will be a liquid at room temperature.
      ii. The chemical bonds in substance A are weaker than those in substance B.
      iii. The chemical bonds in substance A are stronger than those in substance B.
      iv. B will be a gas at room temperature.
3. Boiling point is an important concept to understand.
   a Define 'boiling point'.
   b What change in phase takes place when a liquid reaches its boiling point?
   c What is the boiling point of water?
   d Use the kinetic theory of matter and your knowledge of intermolecular forces, to explain why water changes phase at this temperature.
4. Refer to the table below which gives the melting and boiling points of a number of elements, and then answer the questions that follow. (*Data from http://www.chemicalelements.com*)

Element	**Melting point**	**Boiling point (^{0}C)**
copper	1083	2567
magnesium	650	1107
oxygen	-218.4	-183
carbon	3500	4827
helium	-272	-268.6
sulfur	112.8	444.6

a What state of matter (i.e. solid, liquid or gas) will each of these elements be in at room temperature?

b Which of these elements has the strongest forces between its atoms? Give a reason for your answer.

c Which of these elements has the weakest forces between its atoms? Give a reason for your answer.

Chapter 3

The Atom - Grade 10

We have now looked at many examples of the types of matter and materials that exist around us, and we have investigated some of the ways that materials are classified. But what is it that makes up these materials? And what makes one material different from another? In order to understand this, we need to take a closer look at the building block of matter, the **atom**. Atoms are the basis of all the structures and organisms in the universe. The planets, the sun, grass and trees, the air we breathe, and people are all made up of different combinations of atoms.

3.1 Models of the Atom

It is important to realise that a lot of what we know about the structure of atoms has been developed over a long period of time. This is often how scientific knowledge develops, with one person building on the ideas of someone else. We are going to look at how our modern understanding of the atom has evolved over time.

The idea of atoms was invented by two Greek philosophers, Democritus and Leucippus in the fifth century BC. The Greek word $\alpha\tau o\mu o\nu$ (atom) means *indivisible* because they believed that atoms could not be broken into smaller pieces.

Nowadays, we know that atoms are made up of a *positively charged* **nucleus** in the centre surrounded by *negatively charged* **electrons**. However, in the past, before the structure of the atom was properly understood, scientists came up with lots of different *models* or *pictures* to describe what atoms look like.

Definition: Model
A model is a representation of a system in the real world. Models help us to understand systems and their properties. For example, an *atomic model* represents what the structure of an atom *could* look like, based on what we know about how atoms behave. It is not necessarily a true picture of the exact structure of an atom.

3.1.1 The Plum Pudding Model

After the electron was discovered by J.J. Thomson in 1897, people realised that atoms were made up of even smaller particles than they had previously thought. However, the atomic nucleus had not been discovered yet, and so the 'plum pudding model' was put forward in 1904. In this model, the atom is made up of negative electrons that float in a soup of positive charge, much like plums in a pudding or raisins in a fruit cake (figure 3.1). In 1906, Thomson was awarded the Nobel Prize for his work in this field. However, even with the Plum Pudding Model, there was still no understanding of how these electrons in the atom were arranged.

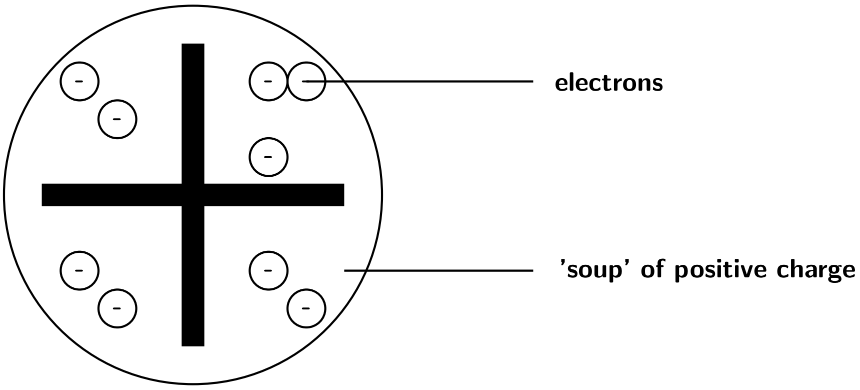


Figure 3.1: A schematic diagram to show what the atom looked like according to the Plum Pudding model

The discovery of **radiation** was the next step along the path to building an accurate picture of atomic structure. In the early twentieth century, Marie Curie and her husband discovered that some elements (the *radioactive* elements) emit particles, which are able to pass through matter in a similar way to X-rays (read more about this in chapter 7). It was Ernest Rutherford who, in 1911, used this discovery to revise the model of the atom.

3.1.2 Rutherford's model of the atom

Radioactive elements emit different types of particles. Some of these are positively charged alpha (α) particles. Rutherford carried out a series of experiments where he bombarded sheets of gold foil with these particles, to try to get a better understanding of where the positive charge in the atom was. A simplified diagram of his experiment is shown in figure 3.2.

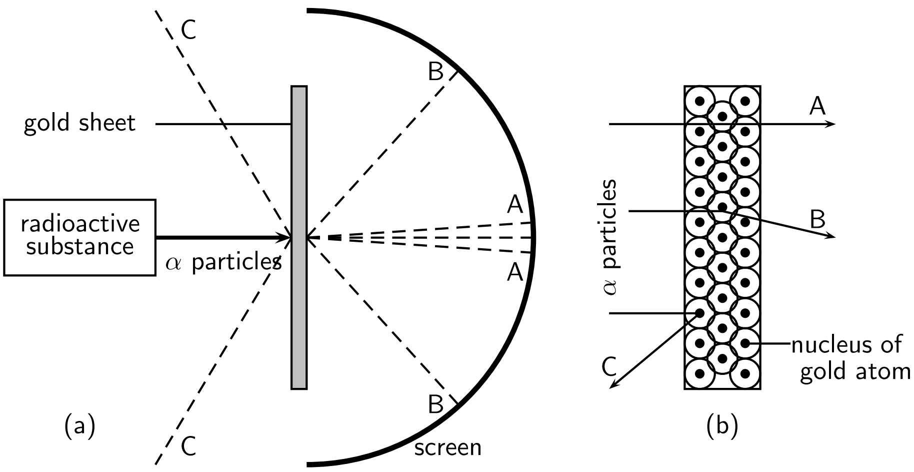


Figure 3.2: Rutherford's gold foil experiment. Figure (a) shows the path of the α particles after they hit the gold sheet. Figure (b) shows the arrangement of atoms in the gold sheets, and the path of the α particles in relation to this.

Rutherford set up his experiment so that a beam of alpha particles was directed at the gold sheets. Behind the gold sheets, was a screen made of zinc sulfide. This screen allowed Rutherford to see where the alpha particles were landing. Rutherford knew that the *electrons* in the gold atoms would not really affect the path of the alpha particles, because the mass of an electron is so much smaller than that of a proton. He reasoned that the positively charged *protons* would be the ones to *repel* the positively charged alpha particles and alter their path.

What he discovered was that most of the alpha particles passed through the foil undisturbed, and could be detected on the screen directly behind the foil (A). Some of the particles ended up being slightly deflected onto other parts of the screen (B). But what was even more interesting was that some of the particles were deflected straight back in the direction from where they had come (C)! These were the particles that had been repelled by the positive protons in the gold atoms. If the Plum Pudding model of the atom were true, then Rutherford would have expected much more repulsion since the positive charge, according to that model, is distributed throughout the atom. But this was not the case. The fact that most particles passed straight through suggested that the positive charge was concentrated in one part of the atom only.

Rutherford's work led to a change in ideas around the atom. His new model described the atom as a tiny, dense, positively charged core called a nucleus, surrounded by lighter, negatively charged electrons. Another way of thinking about this model was that the atom was seen to be like a mini solar system where the electrons orbit the nucleus like planets orbiting around the sun. A simplified picture of this is shown in figure 3.3.

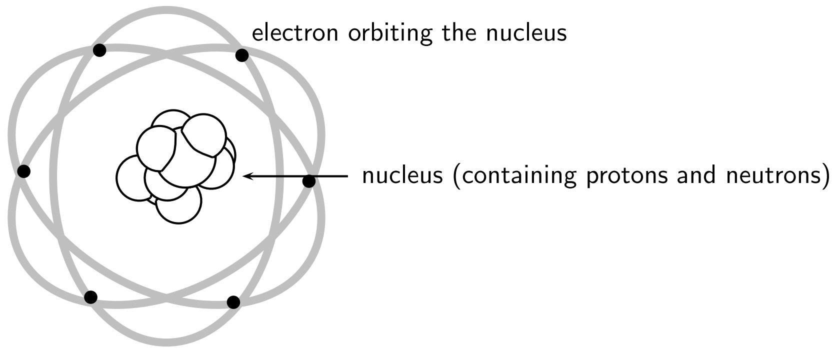


Figure 3.3: Rutherford's model of the atom

3.1.3 The Bohr Model

There were, however, some problems with this model: for example it could not explain the very interesting observation that atoms only emit light at certain wavelengths or frequencies. Niels Bohr solved this problem by proposing that the electrons could only orbit the nucleus in certain special orbits at different energy levels around the nucleus. The exact energies of the orbitals in each energy level depends on the type of atom. Helium for example, has different energy levels to Carbon. If an electron jumps down from a higher energy level to a lower energy level, then light is emitted from the atom. The energy of the light emitted is the same as the gap in the energy between the two energy levels. You can read more about this in section 3.6. The distance between the nucleus and the electron in the lowest energy level of a hydrogen atom is known as the **Bohr radius**.

Interesting Fact

Light has the properties of both a particle *and* a wave! Einstein discovered that light comes in energy packets which are called **photons**. When an electron in an atom changes energy levels, a photon of light is emitted. This photon has the same energy as the difference between the two electron energy levels.

3.2 How big is an atom?

It is difficult sometimes to imagine the size of an atom, or its mass, because we cannot see them, and also because we are not used to working with such small measurements.

3.2.1 How heavy is an atom?

It is possible to determine the mass of a single atom in kilograms. But to do this, you would need very modern *mass spectrometers*, and the values you would get would be very clumsy and difficult to use. The mass of a carbon atom, for example, is about 1.99 x 10^{-26}kg, while the mass of an atom of hydrogen is about 1.67 x 10^{-27}kg. Looking at these very small numbers makes it difficult to compare how much bigger the mass of one atom is when compared to another.

To make the situation simpler, scientists use a different unit of mass when they are describing the mass of an atom. This unit is called the **atomic mass unit** (amu). We can abbreviate (shorten) this unit to just 'u'. If we give carbon an atomic mass of 12 u, then the mass of an atom of hydrogen will be 1 u. You can check this by dividing the mass of a carbon atom in kilograms (see above) by the mass of a hydrogen atom in kilograms (you will need to use a calculator for this!). If you do this calculation, you will see that the mass of a carbon atom is twelve times greater than the mass of a hydrogen atom. When we use atomic mass units instead of kilograms, it becomes easier to see this. Atomic mass units are therefore not giving us the *actual* mass of an atom, but rather its mass *relative* to the mass of other atoms in the Periodic Table. The atomic masses of some elements are shown in table 3.1 below.

Table 3.1: The atomic mass of a number of elements

Element	Atomic mass (u)
Nitrogen (N)	14
Bromine (Br)	80
Magnesium (Mg)	24
Potassium (K)	39
Calcium (Ca)	40
Oxygen (O)	16

The actual value of 1 atomic mass unit is 1.67 x 10^{-24}g or 1.67 x 10^{-27}kg. This is a very tiny mass!

3.2.2 How big is an atom?

pm stands for *picometres*. 1 pm $= 10^{-12}$ m

Atomic diameter also varies depending on the element. On average, the diameter of an atom ranges from 100 pm (Helium) to 670 pm (Caesium). Using different units, 100 pm = 1 Angstrom, and 1 Angstrom = 10^{-10} m. That is the same as saying that 1 Angstrom = 0,0000000010 m or that 100 pm = 0,0000000010 m! In other words, the diameter of an atom ranges from 0.0000000010 m to 0.0000000067 m. This is very small indeed.

3.3 Atomic structure

As a result of the models that we discussed in section 3.1, scientists now have a good idea of what an atom looks like. This knowledge is important because it helps us to understand things like why materials have different properties and why some materials bond with others. Let us now take a closer look at the microscopic structure of the atom.

So far, we have discussed that atoms are made up of a positively charged **nucleus** surrounded by one or more negatively charged **electrons**. These electrons orbit the nucleus.

3.3.1 The Electron

The electron is a very light particle. It has a mass of 9.11 x 10^{-31} kg. Scientists believe that the electron can be treated as a *point particle* or *elementary particle* meaning that it can't be broken down into anything smaller. The electron also carries one unit of **negative** electric charge which is the same as 1.6 x 10^{-19} C (Coulombs).

3.3.2 The Nucleus

Unlike the electron, the nucleus *can* be broken up into smaller building blocks called **protons** and **neutrons**. Together, the protons and neutrons are called **nucleons**.

The Proton

Each proton carries one unit of **positive** electric charge. Since we know that atoms are *electrically neutral*, i.e. do not carry any extra charge, then the number of protons in an atom has to be the same as the number of electrons to balance out the positive and negative charge to zero. The total positive charge of a nucleus is equal to the number of protons in the nucleus. The proton is much heavier than the electron (10 000 times heavier!) and has a mass of 1.6726 x 10^{-27} kg. When we talk about the atomic mass of an atom, we are mostly referring to the combined mass of the protons and neutrons, i.e. the nucleons.

The Neutron

The neutron is electrically neutral i.e. it carries no charge at all. Like the proton, it is much heavier than the electron and its mass is 1.6749 x 10^{-27} kg (slightly heavier than the proton).

Rutherford predicted (in 1920) that another kind of particle must be present in the nucleus along with the proton. He predicted this because if there were only positively charged protons in the nucleus, then it should break into bits because of the repulsive forces between the like-charged protons! Also, if protons were the only particles in the nucleus, then a helium nucleus (atomic number 2) would have two protons and therefore only twice the mass of hydrogen. However, it is actually *four* times heavier than hydrogen. This suggested that there must be something else inside the nucleus as well as the protons. To make sure that the atom stays electrically neutral, this particle would have to be neutral itself. In 1932 James Chadwick discovered the neutron and measured its mass.

	proton	neutron	electron
Mass (kg)	1.6726 x 10^{-27}	1.6749 x 10^{-27}	9.11 x 10^{-31}
Units of charge	+1	0	-1
Charge (C)	1.6 x 10^{-19}	0	-1.6 x 10^{-19}

Table 3.2: Summary of the particles inside the atom

Unlike the electron which is thought to be a *point particle* and unable to be broken up into smaller pieces, the proton and neutron **can** be divided. Protons and neutrons are built up of smaller particles called *quarks*. The proton and neutron are made up of 3 quarks each.

3.4 Atomic number and atomic mass number

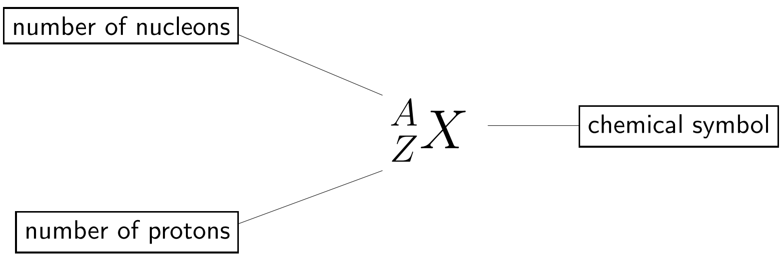
The chemical properties of an element are determined by the charge of its nucleus, i.e. by the *number of protons*. This number is called the **atomic number** and is denoted by the letter **Z**.

Definition: Atomic number (Z)
The number of protons in an atom

The mass of an atom depends on how many nucleons its nucleus contains. The number of nucleons, i.e. the total number of protons *plus* neutrons, is called the **atomic mass number** and is denoted by the letter **A**.

Definition: Atomic mass number (A)
The number of protons and neutrons in the nucleus of an atom

Standard notation shows the chemical symbol, the atomic mass number and the atomic number of an element as follows:



For example, the iron nucleus which has 26 protons and 30 neutrons, is denoted as

$$^{56}_{26}\mathrm{Fe}\ ,$$

where the total nuclear charge is $Z = 26$ and the mass number $A = 56$. The number of neutrons is simply the difference $N = A - Z$.

Important:
Don't confuse the notation we have used above, with the way this information appears on the Periodic Table. On the Periodic Table, the atomic number usually appears in the top lefthand corner of the block or immediately above the element's symbol. The number below the element's symbol is its **relative atomic mass**. This is not exactly the same as the atomic mass number. This will be explained in section 3.5. The example of iron is used again below.

26
Fe
55.85

You will notice in the example of iron that the atomic mass number is more or less the same as its atomic mass. Generally, an atom that contains *n* protons and neutrons (i.e. Z = *n*), will have a mass approximately equal to *n* u. The reason is that a C-12 atom has 6 protons, 6 neutrons and 6 electrons, with the protons and neutrons having about the same mass and the electron mass being negligible in comparison.

Exercise: The structure of the atom

1. Explain the meaning of each of the following terms:
   (a) nucleus
   (b) electron
   (c) atomic mass
2. Complete the following table: (Note: You will see that the atomic masses on the Periodic Table are not *whole numbers*. This will be explained later. For now, you can round off to the nearest whole number.)

Element	**Atomic mass**	**Atomic number**	**Number of protons**	**Number of electrons**	**Number of neutrons**
Mg	24	12			
O			8		
		17			
Ni				28	
	40				20
Zn					
					0
C	12			6	

3. Use standard notation to represent the following elements:
   (a) potassium
   (b) copper
   (c) chlorine
4. For the element $^{35}_{17}$Cl, give the number of ...
   (a) protons
   (b) neutrons
   (c) electrons

   ... in the atom.

5. Which of the following atoms has 7 electrons?
   (a) ${}^{5}_{2}$He
   (b) ${}^{13}_{6}$C
   (c) ${}^{7}_{3}$Li
   (d) ${}^{15}_{7}$N
6. In each of the following cases, give the number or the element symbol represented by 'X'.
   (a) ${}^{40}_{18}$X
   (b) ${}^{x}_{20}$Ca
   (c) ${}^{31}_{x}$P
7. Complete the following table:

	A	Z	N
${}^{235}_{92}$U			
${}^{238}_{92}$U			

   In these two different forms of Uranium...
   (a) What is the *same*?
   (b) What is *different*?

   Uranium can occur in different forms, called *isotopes*. You will learn more about isotopes in section 3.5.

3.5 Isotopes

3.5.1 What is an isotope?

If a few neutrons are added to or removed from a nucleus, the chemical properties of the atom will stay the same because its charge is still the same. Therefore, the chemical properties of an element depend on the number of protons inside the atom. This means that such an atom should remain in the same place in the Periodic table. For example, no matter how many neutrons we add or subtract from a nucleus with 6 protons, that element will *always* be called carbon and have the element symbol C (see the Table of Elements). Atoms which have the same number of protons, but a different number of neutrons, are called **isotopes**.

Definition: Isotope
The **isotope** of a particular element, is made up of atoms which have the same number of protons as the atoms in the orginal element, but a different number of neutrons.

The different isotopes of an element have the same atomic number Z but different mass numbers A because they have a different number of neutrons N. The chemical properties of the different isotopes of an element are the same, but they might vary in how stable their nucleus is. Note that if an element is written for example as C-12, the '12' is the atomic mass of that atom. So, Cl-35 has an atomic mass of 35 u, while Cl-37 has an atomic mass of 37 u.

Interesting Fact

In Greek, "same place" reads as ὶσος τὸπος (isos topos). This is why atoms which have the same number of protons, but different numbers of neutrons, are called *isotopes*. They are in the same place on the Periodic Table!

The following worked examples will help you to understand the concept of an isotope better.

Worked Example 1: Isotopes

Question: For the element $^{234}_{92}\mathrm{U}$ (uranium), use standard notation to describe:

1. the isotope with 2 fewer neutrons
2. the isotope with 4 more neutrons

Answer

Step 1 : Go over the definition of isotope

We know that isotopes of any element have the *same* number of protons (same atomic number) in each atom which means that they have the same chemical symbol. However, they have a different number of neutrons, and therefore a different mass number.

Step 2 : Rewrite the notation for the isotopes

Therefore, any isotope of uranium will have the symbol:

$$\mathrm{U}$$

Also, since the number of protons in uranium isotopes is always the same, we can write down the atomic number:

$$_{92}\mathrm{U}$$

Now, if the isotope we want has 2 fewer neutrons than $^{234}_{92}\mathrm{U}$, then we take the original mass number and subtract 2, which gives:

$$^{232}_{92}\mathrm{U}$$

Following the steps above, we can write the isotope with 4 more neutrons as:

$$^{238}_{92}\mathrm{U}$$

Worked Example 2: Isotopes

Question: Which of the following are isotopes of $^{40}_{20}\mathrm{Ca}$?

- $^{40}_{19}\mathrm{K}$
- $^{42}_{20}\mathrm{Ca}$
- $^{40}_{18}\mathrm{Ar}$

Answer

Step 1 : Go over the definition of isotope:

We know that isotopes have the same atomic number but different mass numbers.

Step 2 : Determine which of the elements listed fits the definition of an isotope.

You need to look for the element that has the same atomic number but a different atomic mass number. The only element is $^{12}_{20}$Ca. What is different is that there are 2 more neutrons than in the original element.

Worked Example 3: Isotopes

Question: For the sulfur isotope $^{33}_{16}S$, give the number of...

1. protons
2. nucleons
3. electrons
4. neutrons

Answer

Step 1 : Determine the number of protons by looking at the atomic number, Z.

Z = 16, therefore the number of protons is 16 (answer to (a)).

Step 2 : Determine the number of nucleons by looking at the atomic mass number, A.

A = 33, therefore the number of nucleons is 33 (answer to (b)).

Step 3 : Determine the number of electrons

The atom is neutral, and therefore the number of electrons is the same as the number of protons. The number of electrons is 16 (answer to (c)).

Step 4 : Calculate the number of neutrons

$$N = A - Z = 33 - 16 = 17$$

The number of neutrons is 17 (answer to (d)).

Exercise: Isotopes

1. Atom A has 5 protons and 5 neutrons, and atom B has 6 protons and 5 neutrons. These atoms are...
   (a) allotropes
   (b) isotopes
   (c) isomers
   (d) atoms of different elements
2. For the sulfur isotopes, $^{32}_{16}S$ and $^{34}_{16}S$, give the number of...
   (a) protons
   (b) nucleons
   (c) electrons
   (d) neutrons
3. Which of the following are isotopes of Cl^{35}?
   (a) $^{17}_{35}Cl$
   (b) $^{35}_{17}Cl$
   (c) $^{37}_{17}Cl$
4. Which of the following are isotopes of U-235? (X represents an element symbol)
   (a) $^{238}_{92}X$
   (b) $^{238}_{90}X$
   (c) $^{235}_{92}X$

3.5.2 Relative atomic mass

It is important to realise that the atomic mass of isotopes of the same element will be different because they have a different number of nucleons. Chlorine, for example, has two common isotopes which are chlorine-35 and chlorine-37. Chlorine-35 has an atomic mass of 35 u, while chlorine-37 has an atomic mass of 37 u. In the world around us, both of these isotopes occur naturally. It doesn't make sense to say that the element chlorine has an atomic mass of 35 u, or that it has an atomic mass of 37 u. Neither of these are absolutely true since the mass varies depending on the form in which the element occurs. We need to look at how much more common one is than the other in order to calculate the **relative atomic mass** for the element chlorine. This is then the number that will appear on the Periodic Table.

Definition: Relative atomic mass
Relative atomic mass is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units.

Worked Example 4: The relative atomic mass of an isotopic element

Question: The element chlorine has two isotopes, chlorine-35 and chlorine-37. The abundance of these isotopes when they occur naturally is 75% chlorine-35 and 25% chlorine-37. Calculate the *average* relative atomic mass for chlorine.

Answer
Step 1 : Calculate the mass contribution of chlorine-35 to the average relative atomic mass

Contribution of Cl-35 = (75/100 x 35) = 26.25 u

Step 2 : Calculate the contribution of chlorine-37 to the average relative atomic mass

Contribution of Cl-37 = (25/100 x 37) = 9.25 u

Step 3 : Add the two values to arrive at the average relative atomic mass of chlorine

Relative atomic mass of chlorine = 26.25 u + 9.25 u = 35.5 u.

If you look on the periodic table, the average relative atomic mass for chlorine is 35,5 u. You will notice that for many elements, the relative atomic mass that is shown is not a whole number. You should now understand that this number is the *average* relative atomic mass for those elements that have naturally occurring isotopes.

Exercise: Isotopes
You are given a sample that contains carbon-12 and carbon-14.

1. Complete the table below:

Isotope	Z	A	Protons	Neutrons	Electrons
Carbon-12					
Carbon-14					
Chlorine-35					
Chlorine-37					

2. If the sample you have contains 90% carbon-12 and 10% carbon-14, calculate the relative atomic mass of an atom in that sample.
3. In another sample, you have 22.5% Cl-37 and 77.5% Cl-35. Calculate the relative atomic mass of an atom in that sample.

Activity :: Group Discussion : The changing nature of scientific knowledge

Scientific knowledge is not static: it changes and evolves over time as scientists build on the ideas of others to come up with revised (and often improved) theories and ideas. In this chapter for example, we saw how peoples' understanding of atomic structure changed as more information was gathered about the atom. There are many more examples like this one in the field of science. Think for example, about our knowledge of the solar system and the origin of the universe, or about the particle and wave nature of light.

Often, these changes in scientific thinking can be very controversial because they disturb what people have come to know and accept. It is important that we realise that what we know *now* about science may also change. An important part of being a scientist is to be a *critical thinker*. This means that you need to question information that you are given and decide whether it is accurate and whether it can be accepted as true. At the same time, you need to learn to be open to new ideas and not to become stuck in what you believe is right... there might just be something new waiting around the corner that you have not thought about!

In groups of 4-5, discuss the following questions:

- Think about some other examples where scientific knowledge has changed because of new ideas and discoveries:
  - What were these new ideas?
  - Were they controversial? If so, why?
  - What role (if any) did *technology* play in developing these new ideas?
  - How have these ideas affected the way we understand the world?
- Many people come up with their own ideas about how the world works. The same is true in science. So how do we, and other scientists, know what to believe and what not to? How do we know when new ideas are 'good' science or 'bad' science? In your groups, discuss some of the things that would need to be done to check whether a new idea or theory was worth listening to, or whether it was not.
- Present your ideas to the rest of the class.

3.6 Energy quantisation and electron configuration

3.6.1 The energy of electrons

You will remember from our earlier discussions, that an atom is made up of a central nucleus, which contains protons and neutrons, and that this nucleus is surrounded by electrons. Although these electrons all have the same charge and the same mass, each electron in an atom has a different amount of *energy*. Electrons that have the *lowest* energy are found closest to the nucleus where the attractive force of the positively charged nucleus is the greatest. Those electrons that have *higher* energy, and which are able to overcome the attractive force of the nucleus, are found further away.

3.6.2 Energy quantisation and line emission spectra

If the energy of an atom is increased (for example when a substance is heated), the energy of the electrons inside the atom can be increased (when an electron has a higher energy than normal it is said to be "excited"). For the excited electron to go back to its original energy (called the ground state), it needs to release energy. It releases energy by emitting light. If one heats up different elements, one will see that for each element, light is emitted only at certain frequencies (or wavelengths). Instead of a smooth continuum of frequencies, we see lines (called emission lines) at particular frequencies. These frequencies correspond to the energy of the emitted light. If electrons could be excited to any energy and lose any amount of energy, there would be a continuous spread of light frequencies emitted. However, the sharp lines we see mean that there are only certain particular energies that an electron can be excited to, or can lose, for each element.

You can think of this like going up a flight of steps: you can't lift your foot by *any* amount to go from the ground to the first step. If you lift your foot too low you'll bump into the step and be stuck on the ground level. You have to lift your foot just the right amount (the height of the step) to go to the next step, and so on. The same goes for electrons and the amount of energy they can have. This is called **quantisation of energy** because there are only certain quantities of energy that an electron can have in an atom. Like steps, we can think of these quantities as **energy levels** in the atom. The energy of the light released when an electron drops down from a higher energy level to a lower energy level is the same as the difference in energy between the two levels.

3.6.3 Electron configuration

Electrons are arranged in energy levels around the nucleus of an atom. Electrons that are in the energy level that is closest to the nucleus, will have the lowest energy and those further away will have a higher energy. Each energy level can only hold a certain number of electrons, and an electron will only be found in the second energy level once the first energy level is full. The same rule applies for the higher energy levels. You will need to learn the following rules:

- The 1st energy level can hold a maximum of 2 electrons
- The 2nd energy level can hold a maximum of 8 electrons
- The 3rd energy level can hold a maximum of 8 electrons
- If the number of electrons in the atom is greater than 18, they will need to move to the 4th energy level.

In the following examples, the energy levels are shown as concentric circles around the central nucleus.

1. **Lithium**

   Lithium (Li) has an atomic number of 3, meaning that in a neutral atom, the number of electrons will also be 3. The first two electrons are found in the first energy level, while the third electron is found in the second energy level (figure 3.11).

2. **Fluorine**

   Fluorine (F) has an atomic number of 9, meaning that a neutral atom also has 9 electrons. The first 2 electrons are found in the first energy level, while the other 7 are found in the second energy level (figure 3.12).

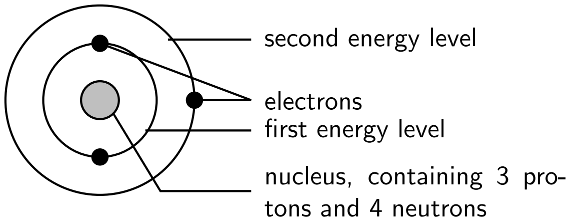


Figure 3.4: The arrangement of electrons in a lithium atom.

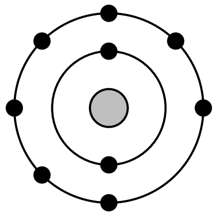


Figure 3.5: The arrangement of electrons in a fluorine atom.

3. **Argon**

   Argon has an atomic number of 18, meaning that a neutral atom also has 18 electrons. The first 2 electrons are found in the first energy level, the next 8 are found in the second energy level, and the last 8 are found in the third energy level (figure 3.6).

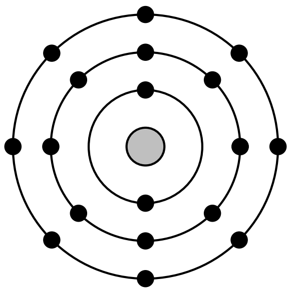
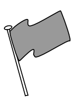


Figure 3.6: The arrangement of electrons in an argon atom.

But the situation is slightly more complicated than this. Within each energy level, the electrons move in **orbitals**. An orbital defines the spaces or regions where electrons move.



Definition: Atomic orbital
An atomic orbital is the region in which an electron may be found around a single atom.

There are different orbital shapes, but we will be dealing with only two. These are the 's' and 'p' orbitals (there are also 'd' and 'f' orbitals). The first energy level contains only one 's' orbital, the second energy level contains one 's' orbital and three 'p' orbitals and the third energy level also contains one 's' orbital and three 'p' orbitals. Within each energy level, the 's' orbital is at a lower energy than the 'p' orbitals. This arrangement is shown in figure 3.7.

When we want to show how electrons are arranged in an atom, we need to remember the following principles:

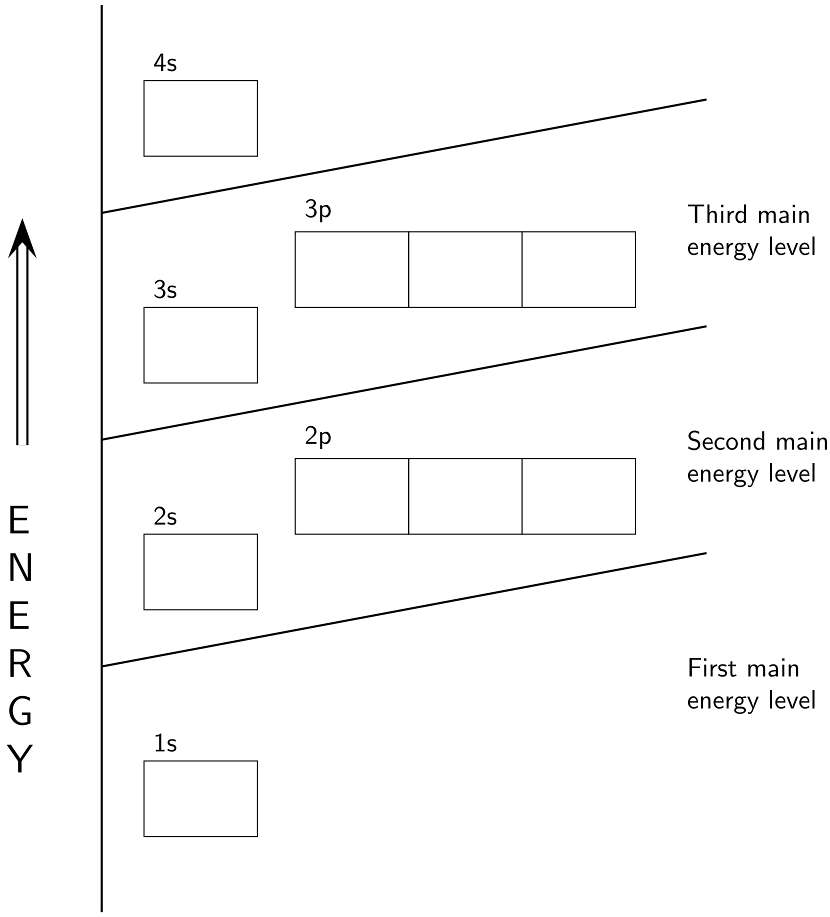


Figure 3.7: The positions of the first ten orbits of an atom on an energy diagram. Note that each block is able to hold two electrons.

- Each orbital can only hold **two electrons**. Electrons that occur together in an orbital are called an **electron pair**. These electrons spin in opposite directions around their own axes.
- An electron will always try to enter an orbital with the lowest possible energy.
- An electron will occupy an orbital on its own, rather than share an orbital with another electron. An electron would also rather occupy a lower energy orbital *with* another electron, before occupying a higher energy orbital. In other words, within one energy level, electrons will fill an 's' orbital before starting to fill 'p' orbitals.

The way that electrons are arranged in an atom is called its **electron configuration**.

Definition: Electron configuration
Electron configuration is the arrangement of electrons in an atom, molecule, or other physical structure.

An element's electron configuration can be represented using **Aufbau diagrams** or energy level diagrams. An Aufbau diagram uses arrows to represent electrons. You can use the following steps to help you to draw an Aufbau diagram:

1. Determine the number of electrons that the atom has.
2. Fill the 's' orbital in the first energy level (the 1s orbital) with the first two electrons.
3. Fill the 's' orbital in the second energy level (the 2s orbital) with the second two electrons.
4. Put one electron in each of the three 'p' orbitals in the second energy level (the 2p orbitals), and then if there are still electrons remaining, go back and place a second electron in each of the 2p orbitals to complete the electron pairs.
5. Carry on in this way through each of the successive energy levels until all the electrons have been drawn.

Important:
When there are two electrons in an orbital, the electrons are called an **electron pair**. If the orbital only has one electron, this electron is said to be an **unpaired electron**. Electron pairs are shown with arrows in opposite directions. This is because when two electrons occupy the same orbital, they spin in opposite directions on their axes.

An Aufbau diagram for the element Lithium is shown in figure 3.8.

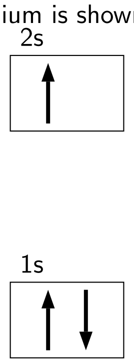


Figure 3.8: The electron configuration of Lithium, shown on an Aufbau diagram

A special type of notation is used to show an atom's electron configuration. The notation describes the energy levels, orbitals and the number of electrons in each. For example, the electron configuration of lithium is $1s^2$ $2s^1$. The number and letter describe the energy level and orbital, and the number above the orbital shows how many electrons are in that orbital.

Aufbau diagrams for the elements fluorine and argon are shown in figures 3.9 and 3.10 respectively. Using standard notation, the electron configuration of fluorine is $1s^2$ $2s^2$ $2p^5$ and the electron configuration of argon is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$.

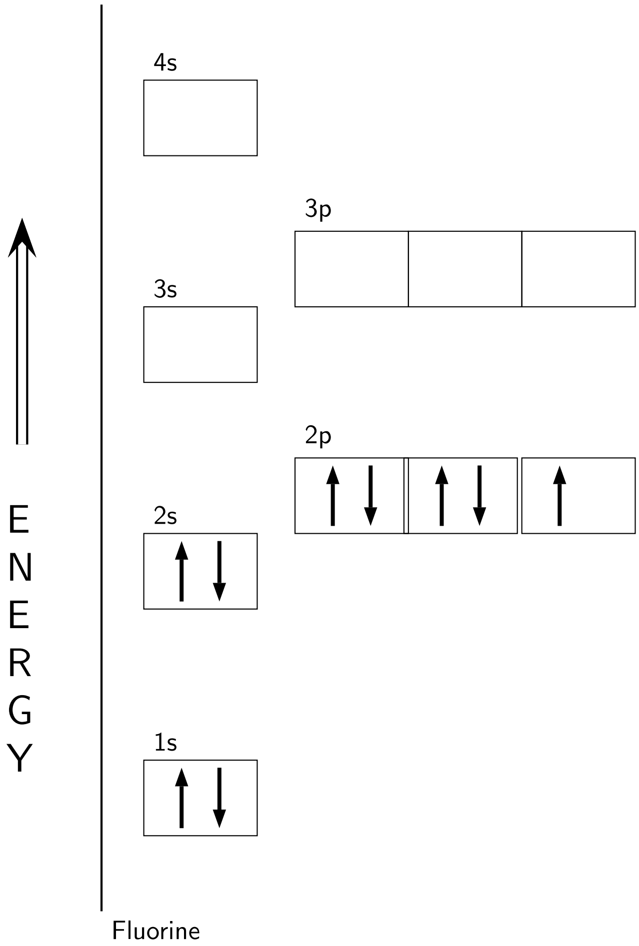


Figure 3.9: An Aufbau diagram showing the electron configuration of fluorine

3.6.4 Core and valence electrons

Electrons in the outermost energy level of an atom are called **valence electrons**. The electrons that are in the energy shells closer to the nucleus are called **core electrons**. Core electrons are all the electrons in an atom, excluding the valence electrons. An element that has its valence energy level full is *more stable* and *less likely to react* than other elements with a valence energy level that is not full.

Definition: Valence electrons
The electrons in the outer energy level of an atom

Definition: Core electrons
All the electrons in an atom, excluding the valence electrons

3.6.5 The importance of understanding electron configuration

By this stage, you may well be wondering why it is important for you to understand how electrons are arranged around the nucleus of an atom. Remember that during chemical reactions, when atoms come into contact with one another, it is the *electrons* of these atoms that will interact first. More specifically, it is the **valence electrons** of the atoms that will determine how they

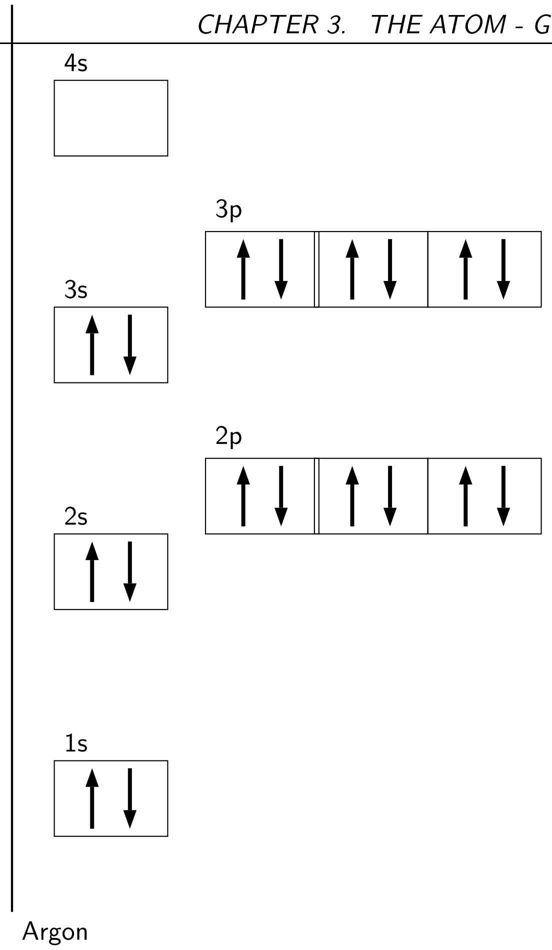


Figure 3.10: An Aufbau diagram showing the electron configuration of argon

react with one another.

To take this a step further, an atom is at its most stable (and therefore *unreactive*) when all its orbitals are full. On the other hand, an atom is least stable (and therefore most *reactive*) when its valence electron orbitals are not full. This will make more sense when we go on to look at chemical bonding in a later chapter. To put it simply, the valence electrons are largely responsible for an element's chemical behaviour, and elements that have the same number of valence electrons often have similar chemical properties.

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Exercise: Energy diagrams and electrons

1. Draw Aufbau diagrams to show the electron configuration of each of the following elements:

    (a) magnesium

    (b) potassium

    (c) sulfur

    (d) neon

    (e) nitrogen

2. Use the Aufbau diagrams you drew to help you complete the following table:

Element	No. of energy levels	No. of core electrons	No. of valence electrons	Electron configuration (standard notation)
Mg				
K				
S				
Ne				
N				

3. Rank the elements used above in order of *increasing reactivity*. Give reasons for the order you give.

Activity :: Group work : Building a model of an atom

Earlier in this chapter, we talked about different 'models' of the atom. In science, one of the uses of models is that they can help us to understand the structure of something that we can't see. In the case of the atom, models help us to build a picture in our heads of what the atom looks like.

Models are often simplified. The small toy cars that you may have played with as a child are models. They give you a good idea of what a real car looks like, but they are much smaller and much simpler. A model cannot always be absolutely accurate and it is important that we realise this so that we don't build up a false idea about something.

In groups of 4-5, you are going to build a model of an atom. Before you start, think about these questions:

- What information do I know about the structure of the atom? (e.g. what parts make it up? how big is it?)
- What materials can I use to represent these parts of the atom as accurately as I can?
- How will I put all these different parts together in my model?

As a group, share your ideas and then plan how you will build your model. Once you have built your model, discuss the following questions:

- Does our model give a good idea of what the atom actually looks like?
- In what ways is our model *inaccurate*? For example, we know that electrons *move* around the atom's nucleus, but in your model, it might not have been possible for you to show this.
- Are there any ways in which our model could be improved?

Now look at what other groups have done. Discuss the same questions for each of the models you see and record your answers.

3.7 Ionisation Energy and the Periodic Table

3.7.1 Ions

In the previous section, we focused our attention on the electron configuration of *neutral* atoms. In a neutral atom, the number of protons is the same as the number of electrons. But what happens if an atom *gains* or *loses* electrons? Does it mean that the atom will still be part of the same element?

A change in the number of electrons of an atom does not change the type of atom that it is. However, the *charge* of the atom will change. If electrons are added, then the atom will become *more negative*. If electrons are taken away, then the atom will become *more positive*. The atom that is formed in either of these cases is called an **ion**. Put simply, an ion is a charged atom.

Definition: Ion
An ion is a charged atom. A positively charged ion is called a **cation** e.g. Na^+, and a negatively charged ion is called an **anion** e.g. F^-. The charge on an ion depends on the number of electrons that have been lost or gained.

Look at the following examples. Notice the number of valence electrons in the neutral atom, the number of electrons that are lost or gained, and the final charge of the ion that is formed.

Lithium

A lithium atoms loses one electrons to form a positive ion (figure 3.11).

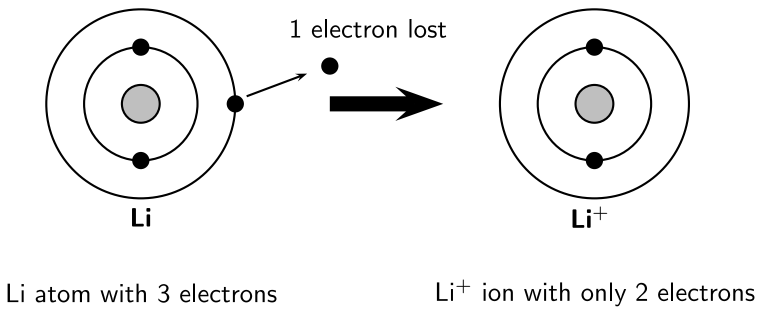

Figure 3.11: The arrangement of electrons in a lithium ion.

In this example, the lithium atom loses an electron to form the cation Li$^+$.

Fluorine

A fluorine atom gains one electron to form a negative ion (figure 3.12).

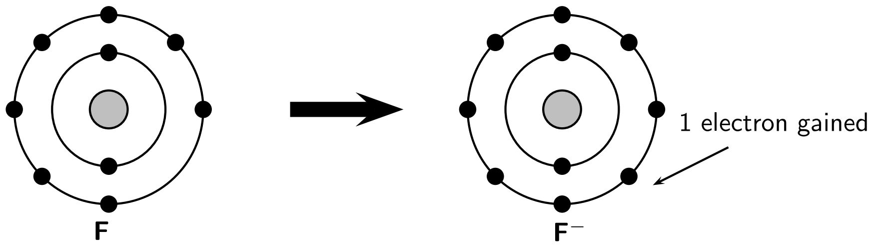

Figure 3.12: The arrangement of electrons in a fluorine ion.

Activity :: Investigation : The formation of ions

1. Use the diagram for lithium as a guide and draw similar diagrams to show how each of the following ions is formed:
   (a) Mg^{2+}
   (b) Na^{+}
   (c) Cl^{-}
   (d) O^{2-}
2. Do you notice anything interesting about the charge on each of these ions? Hint: Look at the number of valence electrons in the neutral atom and the charge on the final ion.

Observations:

Once you have completed the activity, you should notice that:

- In each case the number of electrons that is either gained or lost, is the same as the number of electrons that are needed for the atoms to achieve a full or an empty valence energy level.
- If you look at an energy level diagram for sodium (Na), you will see that in a neutral atom, there is only one valence electron. In order to achieve an empty valence level, and therefore a more stable state for the atom, this electron will be *lost*.
- In the case of oxygen (O), there are six valence electrons. To fill the valence energy level, it makes more sense for this atom to *gain* two electrons. A negative ion is formed.

3.7.2 Ionisation Energy

Ionisation energy is the energy that is needed to remove one electron from an atom. The ionisation energy will be different for different atoms.

The second ionisation energy is the energy that is needed to remove a second electron from an atom, and so on. As an energy level becomes more full, it becomes more and more difficult to remove an electron and the ionisation energy *increases*. On the Periodic Table of the Elements, a *group* is a vertical column of the elements, and a *period* is a horizontal row. In the periodic table, ionisation energy *increases* across a period, but *decreases* as you move down a group. The lower the ionisation energy, the more reactive the element will be because there is a greater chance of electrons being involved in chemical reactions. We will look at this in more detail in the next section.

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Exercise: The formation of ions

Match the information in column A with the information in column B by writing only the letter (A to I) next to the question number (1 to 7)

1. A positive ion that has 3 less electrons than its neutral atom	A. Mg^{2+}
2. An ion that has 1 more electron than its neutral atom	B. Cl^{-}
3. The anion that is formed when bromine gains an electron	C. CO_3^{2-}
4. The cation that is formed from a magnesium atom	D. Al^{3+}
5. An example of a compound ion	E. Br^{2-}
6. A positive ion with the electron configuration of argon	F. K^{+}
7. A negative ion with the electron configuration of neon	G. Mg^{+}
	H. O^{2-}
	I. Br^{-}

3.8 The Arrangement of Atoms in the Periodic Table

The **periodic table of the elements** is a tabular method of showing the chemical elements. Most of the work that was done to arrive at the periodic table that we know, can be attributed to a man called **Dmitri Mendeleev** in 1869. Mendeleev was a Russian chemist who designed the table in such a way that recurring ("periodic") trends in the properties of the elements could be shown. Using the trends he observed, he even left gaps for those elements that he thought were 'missing'. He even predicted the properties that he thought the missing elements would have when they were discovered. Many of these elements were indeed discovered and Mendeleev's predictions were proved to be correct.

To show the recurring properties that he had observed, Mendeleev began new rows in his table so that elements with similar properties were in the same vertical columns, called **groups**. Each row was referred to as a **period**. One important feature to note in the periodic table is that all the non-metals are to the right of the zig-zag line drawn under the element boron. The rest of the elements are metals, with the exception of hydrogen which occurs in the first block of the table despite being a non-metal.

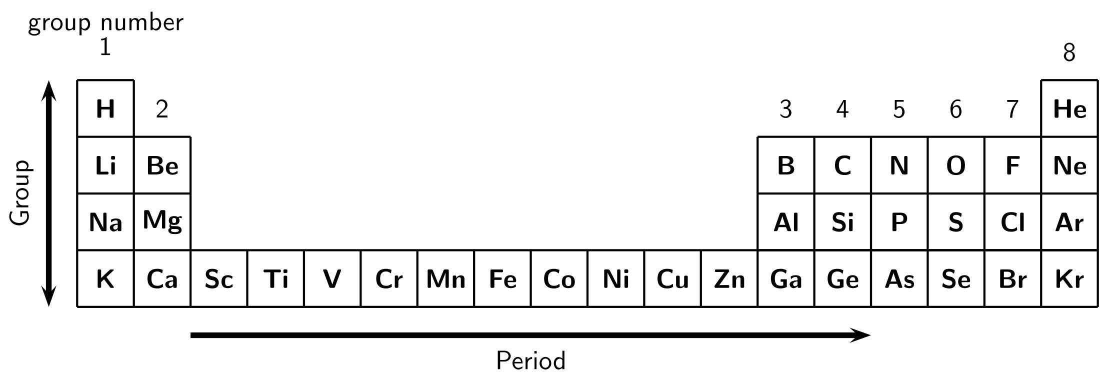

Figure 3.13: A simplified diagram showing part of the Periodic Table

3.8.1 Groups in the periodic table

A *group* is a vertical column in the periodic table, and is considered to be the most important way of classifying the elements. If you look at a periodic table, you will see the groups numbered at the top of each column. The groups are numbered from left to right as follows: 1, 2, then an open space which contains the **transition elements**, followed by groups 3 to 8. These numbers are normally represented using roman numerals. In some periodic tables, all the groups are numbered from left to right from number 1 to number 18. In some groups, the elements display very similar chemical properties, and the groups are even given separate names to identify them.

The characteristics of each group are mostly determined by the electron configuration of the atoms of the element.

- *Group 1:* These elements are known as the **alkali metals** and they are very reactive.

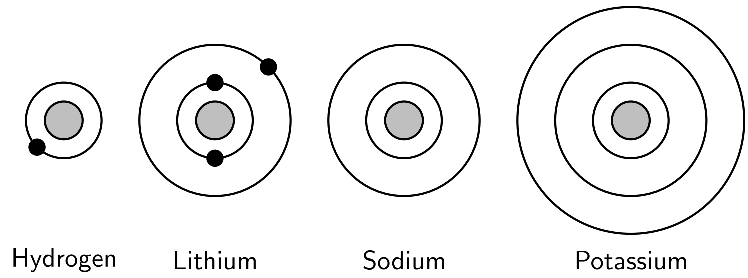

Figure 3.14: Electron diagrams for some of the Group 1 elements

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Activity :: Investigation : The properties of elements
Refer to figure 3.14.

1. Use a Periodic Table to help you to complete the last two diagrams for sodium (Na) and potassium (K).
2. What do you notice about the number of electrons in the valence energy level in each case?
3. Explain why elements from group 1 are more reactive than elements from group 2 on the periodic table (Hint: Think back to 'ionisation energy').

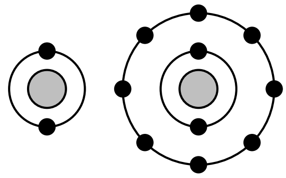
- *Group 2:* These elements are known as the **alkali earth metals**. Each element only has two valence electrons and so in chemical reactions, the group 2 elements tend to *lose* these electrons so that the energy shells are complete. These elements are less reactive than those in group 1 because it is more difficult to lose two electrons than it is to lose one. *Group 3* elements have three valence electrons.

!

Important: The number of valence electrons of an element corresponds to its group number on the periodic table.

- *Group 7:* These elements are known as the **halogens**. Each element is missing just one electron from its outer energy shell. These elements tend to *gain* electrons to fill this shell, rather than losing them.
- *Group 8:* These elements are the **noble gases**. All of the energy shells of the halogens are full, and so these elements are very unreactive.

- *Transition metals:* The differences between groups in the transition metals are not usually dramatic.



Helium Lithium

Figure 3.15: Electron diagrams for two of the noble gases, helium (He) and neon (Ne).

It is worth noting that in each of the groups described above, the **atomic diameter** of the elements increases as you move down the group. This is because, while the number of valence electrons is the same in each element, the number of core electrons increases as one moves down the group.

3.8.2 Periods in the periodic table

A **period** is a horizontal row in the periodic table of the elements. Some of the trends that can be observed within a period are highlighted below:

- As you move from one group to the next within a period, the number of valence electrons increases by one each time.
- Within a single period, all the valence electrons occur in the same energy shell. If the period increases, so does the energy shell in which the valence electrons occur.
- In general, the diameter of atoms decreases as one moves from left to right across a period. Consider the attractive force between the positively charged nucleus and the negatively charged electrons in an atom. As you move across a period, the number of protons in each atom increases. The number of electrons also increases, but these electrons will still be in the same energy shell. As the number of protons increases, the force of attraction between the nucleus and the electrons will increase and the atomic diameter will decrease.
- Ionisation energy increases as one moves from left to right across a period. As the valence electron shell moves closer to being full, it becomes more difficult to remove electrons. The opposite is true when you move down a *group* in the table because more energy shells are being added. The electrons that are closer to the nucleus 'shield' the outer electrons from the attractive force of the positive nucleus. Because these electrons are not being held to the nucleus as strongly, it is easier for them to be removed and the ionisation energy decreases.
- In general, the reactivity of the elements decreases from left to right across a period.

Exercise: Trends in ionisation energy

Refer to the data table below which gives the ionisation energy (in kJ/mol) and atomic number (Z) for a number of elements in the periodic table:

Z	Ionisation energy	Z	Ionisation energy
1	1310	10	2072
2	2360	11	494
3	517	12	734
4	895	13	575
5	797	14	783
6	1087	15	1051
7	1397	16	994
8	1307	17	1250
9	1673	18	1540

1. Draw a line graph to show the relationship between atomic number (on the x-axis) and ionisation energy (y-axis).
2. Describe any trends that you observe.
3. Explain why...
   (a) the ionisation energy for Z=2 is higher than for Z=1
   (b) the ionisation energy for Z=3 is lower than for Z=2
   (c) the ionisation energy increases between Z=5 and Z=7

?

Exercise: Elements in the Periodic Table
Refer to the elements listed below:

Lithium (Li); Chlorine (Cl); Magnesium (Mg); Neon (Ne); Oxygen (O); Calcium (Ca); Carbon (C)

Which of the elements listed above:

1. belongs to Group 1
2. is a halogen
3. is a noble gas
4. is an alkali metal
5. has an atomic number of 12
6. has 4 neutrons in the nucleus of its atoms
7. contains electrons in the 4th energy level
8. has only one valence electron
9. has all its energy orbitals full
10. will have chemical properties that are most similar
11. will form positive ions

3.9 Summary

- Much of what we know today about the atom, has been the result of the work of a number of scientists who have added to each other's work to give us a good understanding of atomic structure.

- Some of the important scientific contributors include **J.J.Thomson** (discovery of the electron, which led to the Plum Pudding Model of the atom), **Ernest Rutherford** (discovery that positive charge is concentrated in the centre of the atom) and **Niels Bohr** (the arrangement of electrons around the nucleus in energy levels).
- Because of the very small mass of atoms, their mass in measured in **atomic mass units** (u). 1 u $= 1.67 \times 10^{-24}$ g.
- An atom is made up of a central **nucleus** (containing **protons** and **neutrons**), surrounded by **electrons**.
- The **atomic number** (Z) is the number of protons in an atom.
- The **atomic mass number** (A) is the number of protons and neutrons in the nucleus of an atom.
- The **standard notation** that is used to write an element, is $^{A}_{Z}\text{X}$, where X is the element symbol, A is the atomic mass number and Z is the atomic number.
- The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons. This means that not all atoms of an element will have the same atomic mass.
- The **relative atomic mass** of an element is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units. The relative atomic mass is written under the elements' symbol on the Periodic Table.
- The energy of electrons in an atom is **quantised**. Electrons occur in specific energy levels around an atom's nucleus.
- Within each energy level, an electron may move within a particular shape of **orbital**. An orbital defines the space in which an electron is most likely to be found. There are different orbital shapes, including s, p, d and f orbitals.
- Energy diagrams such as **Aufbau diagrams** are used to show the electron configuration of atoms.
- The electrons in the outermost energy level are called **valence electrons**.
- The electrons that are not valence electrons are called **core electrons**.
- Atoms whose outermost energy level is full, are less chemically reactive and therefore more stable, than those atoms whose outer energy level is not full.
- An **ion** is a charged atom. A **cation** is a positively charged ion and an **anion** is a negatively charged ion.
- When forming an ion, an atom will lose or gain the number of electrons that will make its valence energy level full.
- An element's **ionisation energy** is the energy that is needed to remove one electron from an atom.
- Ionisation energy increases across a **period** in the Periodic Table.
- Ionisation energy decreases down a **group** in the Periodic Table.

?

Exercise: Summary

1. Write down only the word/term for each of the following descriptions.
   (a) The sum of the number of protons and neutrons in an atom

(b) The defined space around an atom's nucleus, where an electron is most likely to be found

2. For each of the following, say whether the statement is True or False. If it is False, re-write the statement correctly.
   (a) $^{20}_{10}$Ne and $^{22}_{10}$Ne each have 10 protons, 12 electrons and 12 neutrons.
   (b) The atomic mass of any atom of a particular element is always the same.
   (c) It is safer to use helium gas rather than hydrogen gas in balloons.
   (d) Group 1 elements readily form negative ions.

3. Multiple choice questions: In each of the following, choose the **one** correct answer.
   (a) The three basic components of an atom are:
      i. protons, neutrons, and ions
      ii. protons, neutrons, and electrons
      iii. protons, neutrinos, and ions
      iv. protium, deuterium, and tritium
   (b) The charge of an atom is...
      i. positive
      ii. neutral
      iii. negative
   (c) If Rutherford had used neutrons instead of alpha particles in his scattering experiment, the neutrons would...
      i. not deflect because they have no charge
      ii. have deflected more often
      iii. have been attracted to the nucleus easily
      iv. have given the same results
   (d) Consider the isotope $^{234}_{92}$U. Which of the following statements is *true*?
      i. The element is an isotope of $^{234}_{94}$Pu
      ii. The element contains 234 neutrons
      iii. The element has the same electron configuration as $^{238}_{92}$U
      iv. The element has an atomic mass number of 92
   (e) The electron configuration of an atom of chlorine can be represented using the following notation:
      i. $1s^2$ $2s^8$ $3s^7$
      ii. $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$
      iii. $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$
      iv. $1s^2$ $2s^2$ $2p^5$

4. The following table shows the first ionisation energies for the elements of period 1 and 2.

Period	Element	First ionisation energy ($kJ.mol^{-1}$)
1	H	1312
	He	2372
2	Li	520
	Be	899
	B	801
	C	1086
	N	1402
	O	1314
	F	1681
	Ne	2081

   (a) What is the meaning of the term *first ionisation energy*?
   (b) Identify the pattern of first ionisation energies in a period.
   (c) Which TWO elements exert the strongest attractive forces on their electrons? Use the data in the table to give a reason for your answer.

(d) Draw Aufbau diagrams for the TWO elements you listed in the previous question and explain why these elements are so stable.

(e) It is safer to use helium gas than hydrogen gas in balloons. Which property of helium makes it a safer option?

(f) 'Group 1 elements readily form positive ions'.
Is this statement correct? Explain your answer by referring to the table.

Chapter 4

Atomic Combinations - Grade 11

When you look at the matter around you, you will realise that atoms seldom exist on their own. More often, the things around us are made up of different atoms that have been joined together. This is called **chemical bonding**. Chemical bonding is one of the most important processes in chemistry because it allows all sorts of different molecules and combinations of atoms to form, which then make up the objects in the complex world around us. There are, however, some atoms that *do* exist on their own, and which do not bond with others. The **noble gases** in Group 8 of the Periodic Table, behave in this way. They include elements like neon (Ne), helium (He) and argon (Ar). The important question then is, why do some atoms bond but others do not?

4.1 Why do atoms bond?

As we begin this section, it's important to remember that what we will go on to discuss is a *model* of bonding, that is based on a particular *model* of the atom. You will remember from section 3.1 that a model is a *representation* of what is happening in reality. In the model of the atom that has been used so far, the atom is made up of a central nucleus, surrounded by electrons that are arranged in fixed energy levels (also sometimes called *shells*). Within each energy level, electrons move in *orbitals* of different shapes. The electrons in the outermost energy level of an atom are called the **valence electrons**. This model of the atom is useful in trying to understand how different types of bonding take place between atoms.

You will remember from these earlier discussions of electrons and energy levels in the atom, that electrons always try to occupy the *lowest* possible energy level. In the same way, an atom also prefers to exist at the lowest possible energy state so that it is most *stable*. An atom is most stable when all its valence electron orbitals are *full*. In other words, the outer energy level of the atom contains the maximum number of electrons that it can. A stable atom is also an *unreactive* one, and is unlikely to bond with other atoms. This explains why the noble gases are unreactive and why they exist as atoms, rather than as molecules. Look for example at the electron configuration of neon ($1s^2$ $2s^2$ $3p^6$). Neon has eight valence electrons in its valence energy shell. This is the maximum that it can hold and so neon is very stable and unreactive, and will not form new bonds. Other atoms, whose valence energy levels are not full, are more likely to bond in order to become more stable. We are going to look a bit more closely at some of the energy changes that take place when atoms bond.

4.2 Energy and bonding

Let's start by imagining that there are two hydrogen atoms approaching one another. As they move closer together, there are three forces that act on the atoms at the same time. These forces are shown in figure 4.1 and are described below:

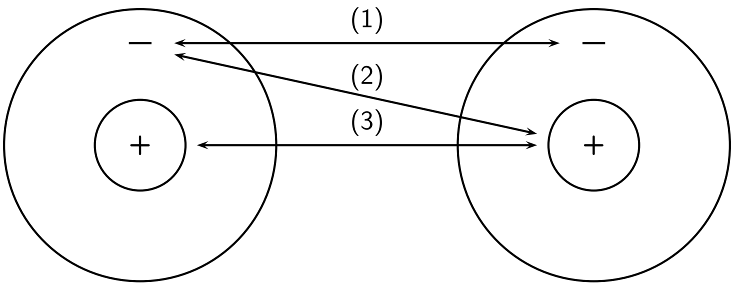


Figure 4.1: Forces acting on two approaching atoms: (1) repulsion between electrons, (2) attraction between protons and electrons and (3) repulsion between protons.

1. **repulsive force** between the electrons of the atoms, since like charges repel
2. **attractive force** between the nucleus of one atom and the electrons of another
3. **repulsive force** between the two positively-charged nuclei

Now look at figure 4.2 to understand the energy changes that take place when the two atoms move towards each other.

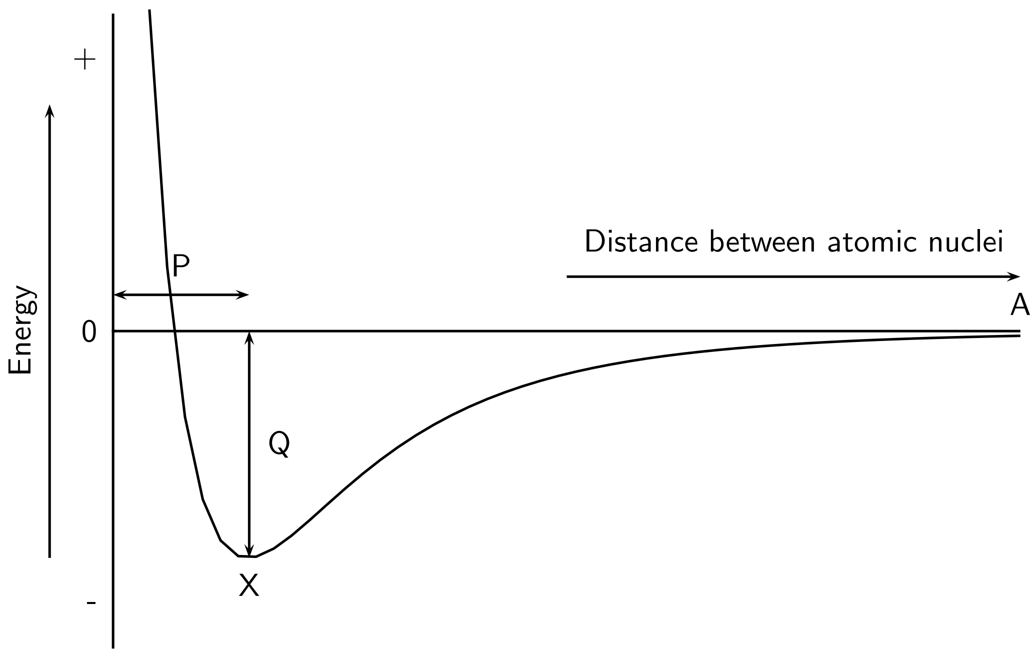


Figure 4.2: Graph showing the change in energy that takes place as atoms move closer together

In the example of the two hydrogen atoms, where the resultant force between them is attraction, the energy of the system is zero when the atoms are far apart (point A), because there is no interaction between the atoms. When the atoms are closer together, attractive forces dominate and the atoms are pulled towards each other. As this happens, the *potential energy* of the system decreases because energy would now need to be *supplied* to the system in order to move the atoms apart. However, as the atoms move closer together (i.e. *left* along the horizontal axis of the graph), repulsive forces start to dominate and this causes the potential energy of the system to rise again. At some point, the attractive and repulsive effects are balanced, and the energy of the system is at its minimum (point X). It is at this point, when the energy is at a minimum, that bonding takes place.

The distance marked 'P' is the **bond length**, i.e. the distance between the nuclei of the atoms when they bond. 'Q' represents the **bond energy** i.e. the amount of energy that must be added to the system to break the bonds that have formed. **Bond strength** means how strongly one atom attracts and is held to another. The strength of a bond is related to the bond length, the size of the bonded atoms and the number of bonds between the atoms. In general, the shorter the bond length, the stronger the bond between the atoms, and the smaller the atoms involved, the stronger the bond. The greater the number of bonds between atoms, the greater will be the bond strength.

4.3 What happens when atoms bond?

A **chemical bond** is formed when atoms are held together by attractive forces. This attraction occurs when electrons are *shared* between atoms, or when electrons are *exchanged* between the atoms that are involved in the bond. The sharing or exchange of electrons takes place so that the outer energy levels of the atoms involved are filled and the atoms are more stable. If an electron is **shared**, it means that it will spend its time moving in the electron orbitals around *both* atoms. If an electron is **exchanged** it means that it is transferred from one atom to another, in other words one atom *gains* an electron while the other *loses* an electron.

Definition: Chemical bond
A chemical bond is the physical process that causes atoms and molecules to be attracted to each other, and held together in more stable chemical compounds.

The type of bond that is formed depends on the elements that are involved. In this section, we will be looking at three types of chemical bonding: **covalent**, **ionic** and **metallic bonding**.

You need to remember that it is the *valence electrons* that are involved in bonding and that atoms will try to fill their outer energy levels so that they are more stable.

4.4 Covalent Bonding

4.4.1 The nature of the covalent bond

Covalent bonding occurs between the atoms of **non-metals**. The outermost orbitals of the atoms overlap so that unpaired electrons in each of the bonding atoms can be shared. By overlapping orbitals, the outer energy shells of all the bonding atoms are filled. The shared electrons move in the orbitals around *both* atoms. As they move, there is an attraction between these negatively charged electrons and the positively charged nuclei, and this force holds the atoms together in a covalent bond.

Definition: Covalent bond
Covalent bonding is a form of chemical bonding where pairs of electrons are shared between atoms.

Below are a few examples. Remember that it is only the *valence electrons* that are involved in bonding, and so when diagrams are drawn to show what is happening during bonding, it is only these electrons that are shown. Circles and crosses represent electrons in different atoms.

Worked Example 5: Covalent bonding

Question: How do hydrogen and chlorine atoms bond covalently in a molecule of hydrogen chloride?

Answer

Step 1 : Determine the electron configuration of each of the bonding atoms.

A chlorine atom has 17 electrons, and an electron configuration of $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$. A hydrogen atom has only 1 electron, and an electron configuration of $1s^1$.

Step 2 : Determine the number of valence electrons for each atom, and how many of the electrons are paired or unpaired.

Chlorine has 7 valence electrons. One of these electrons is unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3 : Look to see how the electrons can be shared between the atoms so that the outermost energy levels of both atoms are full.

The hydrogen atom needs one more electron to complete its valence shell. The chlorine atom also needs one more electron to complete its shell. Therefore *one pair of electrons* must be shared between the two atoms. In other words, one electron from the chlorine atom will spend some of its time orbiting the hydrogen atom so that hydrogen's valence shell is full. The hydrogen electron will spend some of its time orbiting the chlorine atom so that chlorine's valence shell is also full. A molecule of hydrogen chloride is formed (figure 4.3). Notice the shared electron pair in the overlapping orbitals.

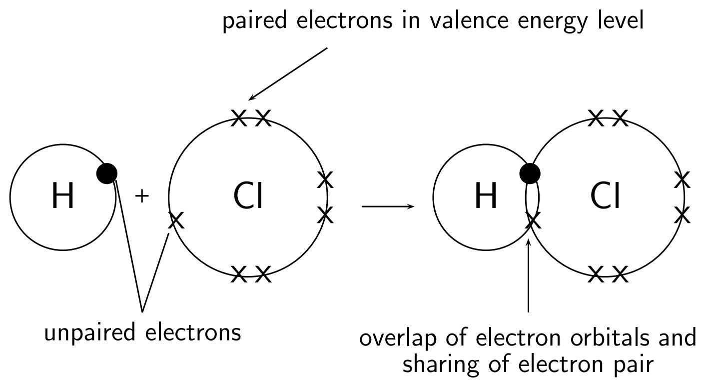


Figure 4.3: Covalent bonding in a molecule of hydrogen chloride

Worked Example 6: Covalent bonding involving multiple bonds

Question: How do nitrogen and hydrogen atoms bond to form a molecule of ammonia (NH_3)?

Answer

Step 1 : Determine the electron configuration of each of the bonding atoms.

A nitrogen atom has 7 electrons, and an electron configuration of $1s^2$ $2s^2$ $2p^3$. A hydrogen atom has only 1 electron, and an electron configuration of $1s^1$.

Step 2 : Determine the number of valence electrons for each atom, and how many of the electrons are paired or unpaired.

Nitrogen has 5 valence electrons meaning that 3 electrons are unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3 : Look to see how the electrons can be shared between the atoms so that the outer energy shells of all atoms are full.

Each hydrogen atom needs one more electron to complete its valence energy shell. The nitrogen atom needs three more electrons to complete its valence energy shell. Therefore *three pairs of electrons* must be shared between the four atoms involved. The nitrogen atom will share three of its electrons so that each of the hydrogen atoms now has a complete valence shell. Each of the hydrogen atoms will share its electron with the nitrogen atom to complete its valence shell (figure 4.4).

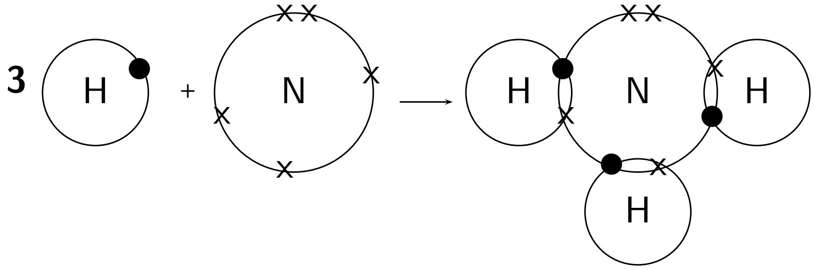


Figure 4.4: Covalent bonding in a molecule of ammonia

The above examples all show **single covalent bonds**, where only one pair of electrons is shared between *the same two atoms*. If two pairs of electrons are shared between the same two atoms, this is called a **double bond**. A **triple bond** is formed if three pairs of electrons are shared.

Worked Example 7: Covalent bonding involving a double bond

Question: How do oxygen atoms bond covalently to form an oxygen molecule?

Answer

Step 1 : Determine the electron configuration of the bonding atoms.

Each oxygen atom has 8 electrons, and their electron configuration is $1s^2$ $2s^2$ $2p^4$.

Step 2 : Determine the number of valence electrons for each atom and how many of these electrons are paired and unpaired.

Each oxygen atom has 6 valence electrons, meaning that each atom has 2 unpaired electrons.

Step 3 : Look to see how the electrons can be shared between atoms so that the outer energy shells of all the atoms are full.

Each oxygen atom needs two more electrons to complete its valence energy shell. Therefore *two pairs of electrons* must be shared between the two oxygen atoms so that both valence shells are full. Notice that the two electron pairs are being shared between *the same two* atoms, and so we call this a **double bond** (figure 4.5).

You will have noticed in the above examples that the number of electrons that are involved in bonding varies between atoms. We say that the **valency** of the atoms is different.

Definition: Valency
The number of electrons in an atom which are used to form a bond.

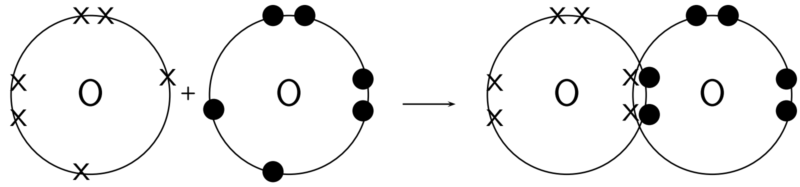


Figure 4.5: A double covalent bond in an oxygen molecule

In the first example, the valency of both hydrogen and chlorine is one, therefore there is a single covalent bond between these two atoms. In the second example, nitrogen has a valency of three and hydrogen has a valency of one. This means that three hydrogen atoms will need to bond with a single nitrogen atom. There are three *single* covalent bonds in a molecule of ammonia. In the third example, the valency of oxygen is two. This means that each oxygen atom will form two bonds with another atom. Since there is only one other atom in a molecule of O_2, a *double covalent* bond is formed between these two atoms.

Important: There is a relationship between the valency of an element and its position on the Periodic Table. For the elements in groups 1 to 4, the valency is the same as the group number. For elements in groups 5 to 7, the valency is calculated by subtracting the group number from 8. For example, the valency of fluorine (group 7) is 8-7=1, while the valency of calcium (group 2) is 2. Some elements have more than one possible valency, so you always need to be careful when you are writing a chemical formula. Often, the valency will be written in a bracket after the element symbol e.g. carbon (iv) oxide, means that in this molecule carbon has a valency of 4.

Exercise: Covalent bonding and valency

1. Explain the difference between the *valence electrons* and the *valency* of an element.
2. Complete the table below by filling in the number of valence electrons and the valency for each of the elements shown:

Element	**No. of valence electrons**	**No. of electrons needed to fill outer shell**	**Valency**
Mg			
K			
F			
Ar			
C			
N			
O			

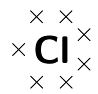
3. Draw simple diagrams to show how electrons are arranged in the following covalent molecules:
   (a) Calcium oxide (CaO)
   (b) Water (H_2O)
   (c) Chlorine (Cl_2)

4.5 Lewis notation and molecular structure

Although we have used diagrams to show the structure of molecules, there are other forms of notation that can be used, such as **Lewis notation** and **Couper notation**. **Lewis notation** uses dots and crosses to represent the **valence electrons** on different atoms. The chemical symbol of the element is used to represent the nucleus and the core electrons of the atom.

So, for example, a hydrogen atom would be represented like this:

A chlorine atom would look like this:



A molecule of hydrogen chloride would be shown like this:

The dot and cross inbetween the two atoms, represent the pair of electrons that are shared in the covalent bond.

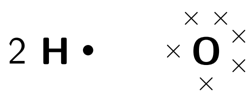
Worked Example 8: Lewis notation: Simple molecules

Question: Represent the molecule H_2O using Lewis notation

Answer

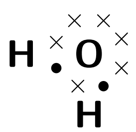
Step 1 : For each atom, determine the number of valence electrons in the atom, and represent these using dots and crosses.

The electron configuration of hydrogen is $1s^1$ and the electron configuration for oxygen is $1s^2$ $2s^2$ $2p^4$. Each hydrogen atom has one valence electron, which is unpaired, and the oxygen atom has six valence electrons with two unpaired.



Step 2 : Arrange the electrons so that the outermost energy level of each atom is full.

The water molecule is represented below.



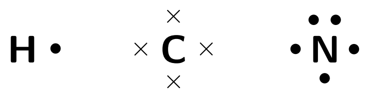
Worked Example 9: Lewis notation: Molecules with multiple bonds

Question: Represent the molecule HCN using Lewis notation

Answer

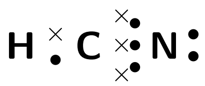
Step 1 : For each atom, determine the number of valence electrons that the atom has from its electron configuration.

The electron configuration of hydrogen is $1s^1$, the electron configuration of nitrogen is $1s^2\ 2s^2\ 2p^3$ and for carbon is $1s^2\ 2s^2\ 2p^2$. This means that hydrogen has one valence electron which is unpaired, carbon has four valence electrons, all of which are unpaired, and nitrogen has five valence electrons, three of which are unpaired.



Step 2 : Arrange the electrons in the HCN molecule so that the outermost energy level in each atom is full.

The HCN molecule is represented below. Notice the three electron pairs between the nitrogen and carbon atom. Because these three covalent bonds are between the same two atoms, this is a *triple* bond.



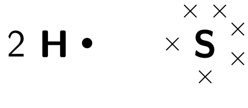
Worked Example 10: Lewis notation: Atoms with variable valencies

Question: Represent the molecule H_2S using Lewis notation

Answer

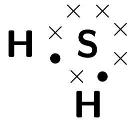
Step 1 : Determine the number of valence electrons for each atom.

Hydrogen has an electron configuration of $1s^1$ and sulfur has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. Each hydrogen atom has one valence electron which is unpaired, and sulfur has six valence electrons. Although sulfur has a variable valency, we know that the sulfur will be able to form 2 bonds with the hydrogen atoms. In this case, the valency of sulfur must be two.



Step 2 : Arrange the atoms in the molecule so that the outermost energy level in each atom is full.

The H_2S molecule is represented below.



Another way of representing molecules is using **Couper notation**. In this case, only the electrons that are involved in the bond between the atoms are shown. A line is used for each covalent bond. Using Couper notation, a molecule of water and a molecule of HCN would be represented as shown in figures 4.6 and 4.7 below.

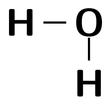


Figure 4.6: A water molecule represented using Couper notation

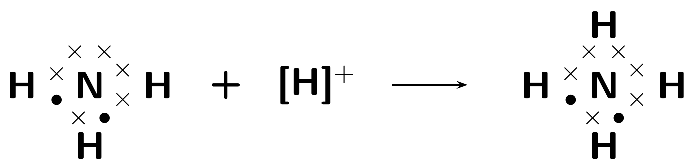
$$H-C\equiv N$$

Figure 4.7: A molecule of HCN represented using Couper notation

Extension: Dative covalent bonds

A **dative covalent bond** (also known as a coordinate covalent bond) is a description of covalent bonding between two atoms in which both electrons shared in the bond come from the same atom. This happens when a Lewis base (an electron donor) donates a pair of electrons to a Lewis acid (an electron acceptor). Lewis acids and bases will be discussed in section 15.1 in chapter 15.

One example of a molecule that contains a dative covalent bond is the ammonium ion (NH_4^+) shown in the figure below. The hydrogen ion H^+ does not contain any electrons, and therefore the electrons that are in the bond that forms between this ion and the nitrogen atom, come only from the nitrogen.



Exercise: Atomic bonding and Lewis notation

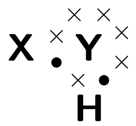
1. Represent each of the following *atoms* using Lewis notation:
   (a) beryllium
   (b) calcium
   (c) lithium
2. Represent each of the following *molecules* using Lewis notation:
   (a) bromine gas (Br_2)

(b) calcium chloride ($CaCl_2$)
(c) carbon dioxide (CO_2)

3. Which of the three molecules listed above contains a double bond?
4. Two chemical reactions are described below.
   - nitrogen and hydrogen react to form ammonia
   - carbon and hydrogen bond to form a molecule of methane (CH_4)

   For each reaction, give:

   (a) the valency of each of the atoms involved in the reaction
   (b) the Lewis structure of the product that is formed
   (c) the chemical formula of the product
   (d) the name of the product
5. A chemical compound has the following Lewis notation:



   (a) How many valence electrons does element Y have?
   (b) What is the valency of element Y?
   (c) What is the valency of element X?
   (d) How many covalent bonds are in the molecule?
   (e) Suggest a name for the elements X and Y.

4.6 Electronegativity

Electronegativity is a measure of how strongly an atom pulls a shared electron pair towards it. The table below shows the electronegativities of a number of elements:

Table 4.1: Table of electronegativities for selected elements

Element	**Electronegativity**
Hydrogen (H)	2.1
Sodium (Na)	0.9
Magnesium (Mg)	1.2
Calcium (Ca)	1.0
Chlorine (Cl)	3.0
Bromine (Br)	2.8

Definition: Electronegativity
Electronegativity is a chemical property which describes the power of an atom to attract electrons towards itself.

The greater the electronegativity of an element, the stronger its attractive pull on electrons. For example, in a molecule of hydrogen bromide (HBr), the electronegativity of bromine (2.8) is higher than that of hydrogen (2.1), and so the shared electrons will spend more of their time closer to the bromine atom. Bromine will have a slightly negative charge, and hydrogen will have a slightly positive charge. In a molecule like hydrogen (H_2) where the electronegativities of the atoms in the molecule are the same, both atoms have a neutral charge.

The concept of electronegativity was introduced by *Linus Pauling* in 1932, and this became very useful in predicting the nature of bonds between atoms in molecules. In 1939, he published a book called 'The Nature of the Chemical Bond', which became one of the most influential chemistry books ever published. For this work, Pauling was awarded the Nobel Prize in Chemistry in 1954. He also received the Nobel Peace Prize in 1962 for his campaign against above-ground nuclear testing.

4.6.1 Non-polar and polar covalent bonds

Electronegativity can be used to explain the difference between two types of covalent bonds. **Non-polar covalent bonds** occur between two identical non-metal atoms, e.g. H_2, Cl_2 and O_2. Because the two atoms have the same electronegativity, the electron pair in the covalent bond is shared equally between them. However, if two different non-metal atoms bond then the shared electron pair will be pulled more strongly by the atom with the highest electronegativity. As a result, a **polar covalent bond** is formed where one atom will have a slightly negative charge and the other a slightly positive charge. This is represented using the symbols δ^+ (slightly positive) and δ^- (slightly negative). So, in a molecule such as hydrogen chloride (HCl), hydrogen is H^{δ^+} and chlorine is Cl^{δ^-}.

4.6.2 Polar molecules

Some molecules with polar covalent bonds are **polar molecules**, e.g. H_2O. But not *all* molecules with polar covalent bonds are polar. An example is CO_2. Although CO_2 has two polar covalent bonds (between C^+ atom and the two O^- atoms), the molecule itself is not polar. The reason is that CO_2 is a linear molecule and is therefore symmetrical. So there is no difference in charge between the two ends of the molecule. The *polarity* of molecules affects properties such as *solubility*, *melting points* and *boiling points*.

Definition: Polar and non-polar molecules

A **polar molecule** is one that has one end with a slightly positive charge, and one end with a slightly negative charge. A **non-polar molecule** is one where the charge is equally spread across the molecule.

Exercise: Electronegativity

1. In a molecule of hydrogen chloride (HCl),
   (a) What is the electronegativity of hydrogen
   (b) What is the electronegativity of chlorine?
   (c) Which atom will have a slightly positive charge and which will have a slightly negative charge in the molecule?
   (d) Is the bond a non-polar or polar covalent bond?
   (e) Is the molecule polar or non-polar?

2. Complete the table below:

Molecule	Difference in electronegativity between atoms	Non-polar/polar covalent bond	Polar/non-polar molecule
H_2O			
HBr			
NO_2			
F_2			
CH_4			

4.7 Ionic Bonding

4.7.1 The nature of the ionic bond

You will remember that when atoms bond, electrons are either *shared* or they are *transferred* between the atoms that are bonding. In covalent bonding, electrons are shared between the atoms. There is another type of bonding, where electrons are *transferred* from one atom to another. This is called **ionic bonding**.

Ionic bonding takes place when the difference in electronegativity between the two atoms is more than 1.7. This usually happens when a metal atom bonds with a non-metal atom. When the difference in electronegativity is large, one atom will attract the shared electron pair much more strongly than the other, causing electrons to be transferred from one atom to the other.

Definition: Ionic bond
An ionic bond is a type of chemical bond based on the electrostatic forces between two oppositely-charged ions. When ionic bonds form, a metal donates an electron, due to a low electronegativity, to form a positive ion or cation. The non-metal atom has a high electronegativity, and therefore readily gains electrons to form negative ions or anions. The two or more ions are then attracted to each other by electrostatic forces.

Example 1:

In the case of NaCl, the difference in electronegativity is 2.1. Sodium has only one valence electron, while chlorine has seven. Because the electronegativity of chlorine is higher than the electronegativity of sodium, chlorine will attract the valence electron in the sodium atom very strongly. This electron from sodium is transferred to chlorine. Sodium has lost an electron and forms a Na^+ ion. Chlorine gains an electron and forms a Cl^- ion. The attractive force between the positive and negative ion is what holds the molecule together.

The balanced equation for the reaction is:

$$Na + Cl \rightarrow NaCl$$

This can be represented using Lewis notation:

Example 2:

Another example of ionic bonding takes place between magnesium (Mg) and oxygen (O) to form magnesium oxide (MgO). Magnesium has two valence electrons and an electronegativity of 1.2, while oxygen has six valence electrons and an electronegativity of 3.5. Since oxygen has

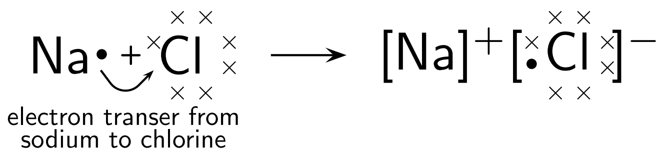


Figure 4.8: Ionic bonding in sodium chloride

a higher electronegativity, it attracts the two valence electrons from the magnesium atom and these electrons are transferred from the magnesium atom to the oxygen atom. Magnesium loses two electrons to form Mg^{2+}, and oxygen gains two electrons to form O^{2-}. The attractive force between the oppositely charged ions is what holds the molecule together.

The balanced equation for the reaction is:

$$2Mg + O_2 \rightarrow 2MgO$$

Because oxygen is a diatomic molecule, two magnesium atoms will be needed to combine with two oxygen atoms to produce two molecules of magnesium oxide (MgO).

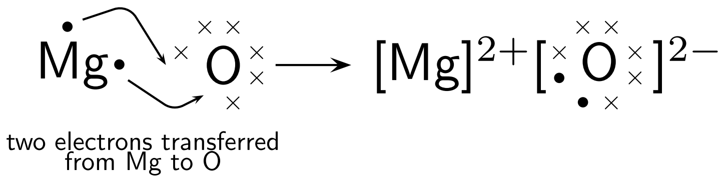


Figure 4.9: Ionic bonding in magnesium oxide

Important: Notice that the number of electrons that is either lost or gained by an atom during ionic bonding, is the same as the **valency** of that element

Exercise: Ionic compounds

1. Explain the difference between a *covalent* and an *ionic* bond.
2. Magnesium and chlorine react to form magnesium chloride.
   (a) What is the difference in electronegativity between these two elements?
   (b) Give the chemical formula for:
      - a magnesium ion
      - a choride ion
      - the ionic compound that is produced during this reaction
   (c) Write a balanced chemical equation for the reaction that takes place.
3. Draw Lewis diagrams to represent the following ionic compounds:
   (a) sodium iodide (NaI)
   (b) calcium bromide ($CaBr_2$)
   (c) potassium chloride (KCl)

4.7.2 The crystal lattice structure of ionic compounds

Ionic substances are actually a combination of lots of ions bonded together into a giant molecule. The arrangement of ions in a regular, geometric structure is called a **crystal lattice**. So in fact NaCl does not contain one Na and one Cl ion, but rather a lot of these two ions arranged in a crystal lattice where the ratio of Na to Cl ions is 1:1. The structure of a crystal lattice is shown in figure 4.10.

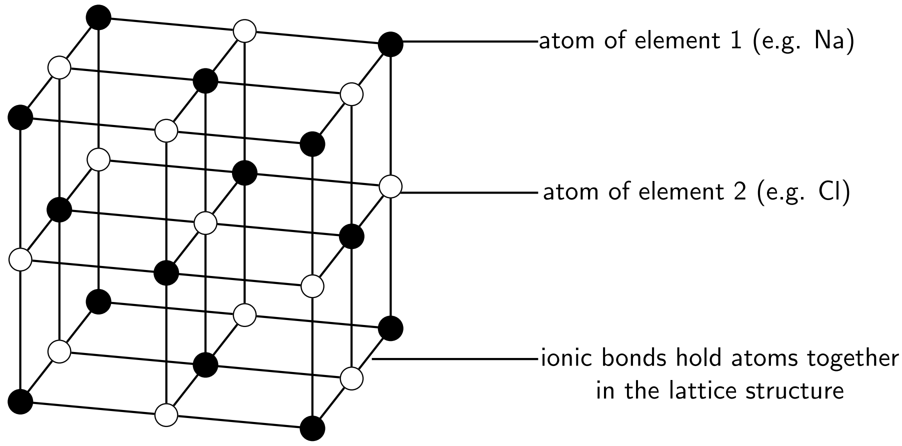


Figure 4.10: The crystal lattice arrangement in an ionic compound (e.g. NaCl)

4.7.3 Properties of Ionic Compounds

Ionic compounds have a number of properties:

- Ions are arranged in a lattice structure
- Ionic solids are crystalline at room temperature
- The ionic bond is a strong electrical attraction. This means that ionic compounds are often hard and have high melting and boiling points
- Ionic compounds are brittle, and bonds are broken along planes when the compound is stressed
- Solid crystals don't conduct electricity, but ionic solutions do

4.8 Metallic bonds

4.8.1 The nature of the metallic bond

The structure of a metallic bond is quite different from covalent and ionic bonds. In a metal bond, the valence electrons are *delocalised*, meaning that an atom's electrons do not stay around that one nucleus. In a metallic bond, the positive atomic nuclei (sometimes called the 'atomic kernels') are surrounded by a sea of delocalised electrons which are attracted to the nuclei (figure 4.11).

Definition: Metallic bond
Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal.

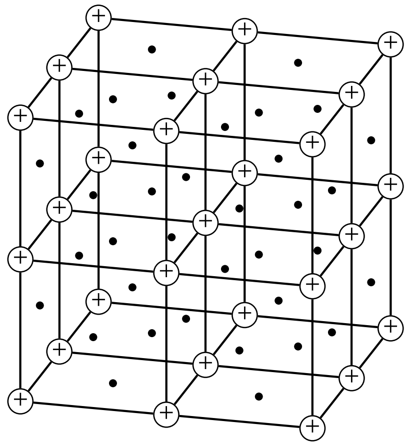


Figure 4.11: Positive atomic nuclei (+) surrounded by delocalised electrons (•)

4.8.2 The properties of metals

Metals have several unique properties as a result of this arrangement:

- *Thermal conductors*

  Metals are good conductors of heat and are therefore used in cooking utensils such as pots and pans. Because the electrons are loosely bound and are able to move, they can transport heat energy from one part of the material to another.

- *Electrical conductors*

  Metals are good conductors of electricity, and are therefore used in electrical conducting wires. The loosely bound electrons are able to move easily and to transfer charge from one part of the material to another.

- *Shiny metallic lustre*

  Metals have a characteristic shiny appearance and are often used to make jewellery. The loosely bound electrons are able to absorb and reflect light at all frequencies, making metals look polished and shiny.

- *Malleable and ductile*

  This means that they can be bent into shape without breaking (malleable) and can be stretched into thin wires (ductile) such as copper, which can then be used to conduct electricity. Because the bonds are not fixed in a particular direction, atoms can slide easily over one another, making metals easy to shape, mould or draw into threads.

- *Melting point*

  Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged. The high melting point is due to the high strength of metallic bonds.

- *Density*

  Metals have a high density because their atoms are packed closely together.

?

Exercise: Chemical bonding

1. Give two examples of everyday objects that contain..
   (a) covalent bonds
   (b) ionic bonds

(c) metallic bonds

2. Complete the table which compares the different types of bonding:

	Covalent	**Ionic**	**Metallic**
Types of atoms involved			
Nature of bond between atoms			
Melting Point (high/low)			
Conducts electricity? (yes/no)			
Other properties			

3. Complete the table below by identifying the type of bond (covalent, ionic or metallic) in each of the compounds:

Molecular formula	**Type of bond**
H_2SO_4	
FeS	
NaI	
$MgCl_2$	
Zn	

4. Which of these substances will conduct electricity most effectively? Give a reason for your answer.
5. Use your knowledge of the different types of bonding to explain the following statements:
   (a) Swimming during an electric thunderstorm (i.e. where there is lightning) can be very dangerous.
   (b) Most jewellery items are made from metals.
   (c) Plastics are good insulators.

4.9 Writing chemical formulae

4.9.1 The formulae of covalent compounds

To work out the formulae of covalent compounds, we need to use the valency of the atoms in the compound. This is because the valency tells us how many bonds each atom can form. This in turn can help to work out how many atoms of each element are in the compound, and therefore what its formula is. The following are some examples where this information is used to write the chemical formula of a compound.

Worked Example 11: Formulae of covalent compounds

Question: Write the chemical formula for water

Answer

Step 1 : Write down the elements that make up the compound.

A molecule of water contains the elements *hydrogen* and *oxygen*.

Step 2 : Determine the valency of each element

The valency of hydrogen is 1 and the valency of oxygen is 2. This means that oxygen can form two bonds with other elements and each of the hydrogen atoms can form one.

Step 3 : Write the chemical formula
Using the valencies of hydrogen and oxygen, we know that in a single water molecule, two hydrogen atoms will combine with one oxygen atom. The chemical formula for water is therefore:

$$\mathbf{H_2O}.$$

Worked Example 12: Formulae of covalent compounds

Question: Write the chemical formula for magnesium oxide

Answer
Step 1 : Write down the elements that make up the compound.
A molecule of magnesium oxide contains the elements *magnesium* and *oxygen*.

Step 2 : Determine the valency of each element
The valency of magnesium is 2, while the valency of oxygen is also 2. In a molecule of magnesium oxide, one atom of magnesium will combine with one atom of oxygen.

Step 3 : Write the chemical formula
The chemical formula for magnesium oxide is therefore:

$$\mathbf{MgO}$$

Worked Example 13: Formulae of covalent compounds

Question: Write the chemical formula for copper (II) chloride.

Answer
Step 1 : Write down the elements that make up the compound.
A molecule of copper (II) chloride contains the elements *copper* and *chlorine*.

Step 2 : Determine the valency of each element
The valency of copper is 2, while the valency of chlorine is 1. In a molecule of copper (II) chloride, two atoms of chlorine will combine with one atom of copper.

Step 3 : Write the chemical formula
The chemical formula for copper (II) chloride is therefore:

$$\mathbf{CuCl_2}$$

4.9.2 The formulae of ionic compounds

The overall charge of an ionic compound will always be zero and so the negative and positive charge must be the same size. We can use this information to work out what the chemical formula of an ionic compound is if we know the charge on the individual ions. In the case of NaCl for example, the charge on the sodium is +1 and the charge on the chlorine is -1. The charges balance (+1-1=0) and therefore the ionic compound is neutral. In MgO, magnesium has a charge of +2 and oxygen has a charge of -2. Again, the charges balance and the compound is neutral. Positive ions are called **cations** and negative ions are called **anions**.

Some ions are made up of groups of atoms, and these are called **compound ions**. It is a good idea to learn the compound ions that are shown in table 4.2

Table 4.2: Table showing common compound ions and their formulae

Name of compound ion	formula
Carbonate	$CO_3{}^{2-}$
sulphate	$SO_4{}^{2-}$
Hydroxide	OH^-
Ammonium	$NH_4{}^+$
Nitrate	$NO_3{}^-$
Hydrogen carbonate	$HCO_3{}^-$
Phosphate	$PO_4{}^{3-}$
Chlorate	$ClO_3{}^-$
Cyanide	CN^-
Chromate	$CrO_4{}^{2-}$
Permanganate	$MnO_4{}^-$

In the case of ionic compounds, the valency of an ion is the same as its charge (Note: valency is always expressed as a *positive* number e.g. valency of the chloride ion is 1 and not -1). Since an ionic compound is always *neutral*, the positive charges in the compound must balance out the negative. The following are some examples:

Worked Example 14: Formulae of ionic compounds

Question: Write the chemical formula for potassium iodide.

Answer

Step 1 : Write down the ions that make up the compound.

Potassium iodide contains potassium and iodide ions.

Step 2 : Determine the valency and charge of each ion.

Potassium iodide contains the ions K^+ (valency = 1; charge = +1) and I^- (valency = 1; charge = -1). In order to balance the charge in a single molecule, one atom of potassium will be needed for every one atom of iodine.

Step 3 : Write the chemical formula

The chemical formula for potassium iodide is therefore:

KI

Worked Example 15: Formulae of ionic compounds

Question: Write the chemical formula for sodium sulphate.

Answer

Step 1 : Write down the ions that make up the compound.

Sodium sulphate contains sodium ions and sulphate ions.

Step 2 : Determine the valency and charge of each ion.

Na^+ (valency = 1; charge = +1) and $SO_4{}^{2-}$ (valency = 2; charge = -2).

Step 3 : Write the chemical formula.

Two sodium ions will be needed to balance the charge of the sulphate ion. The chemical formula for sodium sulphate is therefore:

$$\mathbf{Na_2SO_4{}^{2-}}$$

Worked Example 16: Formulae of ionic compounds

Question: Write the chemical formula for calcium hydroxide.

Answer

Step 1 : Write down the ions that make up the compound.

Calcium hydroxide contains calcium ions and hydroxide ions.

Step 2 : Determine the valency and charge of each ion.

Calcium hydroxide contains the ions Ca^{2+} (charge = +2) and OH^- (charge = -1). In order to balance the charge in a single molecule, two hydroxide ions will be needed for every ion of calcium.

Step 3 : Write the chemical formula.

The chemical formula for calcium hydroxide is therefore:

$$\mathbf{Ca(OH)_2}$$

Exercise: Chemical formulae

1. Copy and complete the table below:

Compound	Cation	Anion	Formula
	Na^+	Cl^-	
potassium bromide		Br^-	
	NH_4^+	Cl^-	
potassium chromate			
			PbI
potassium permanganate			
calcium phosphate			

2. Write the chemical formula for each of the following compounds:
   (a) hydrogen cyanide
   (b) carbon dioxide
   (c) sodium carbonate
   (d) ammonium hydroxide
   (e) barium sulphate
   (f) potassium permanganate

4.10 The Shape of Molecules

4.10.1 Valence Shell Electron Pair Repulsion (VSEPR) theory

The shape of a covalent molecule can be predicted using the Valence Shell Electron Pair Repulsion (VSEPR) theory. This is a model in chemistry that tries to predict the shapes of molecules. Very simply, VSEPR theory says that the electron pairs in a molecule will arrange themselves around the central atom of the molecule so that the repulsion between their negative charges is as small as possible. In other words, the electron pairs arrange themselves so that they are as far apart as they can be. Depending on the number of electron pairs in the molecule, it will have a different shape.

Definition: Valence Shell Electron Pair Repulsion Theory

Valence shell electron pair repulsion theory (VSEPR) is a model in chemistry, which is used to predict the shape of individual molecules, based upon the extent of their electron-pair repulsion.

VSEPR theory is based on the idea that the geometry of a molecule is mostly determined by repulsion among the pairs of electrons around a central atom. The pairs of electrons may be bonding or non-bonding (also called lone pairs). Only valence electrons of the central atom influence the molecular shape in a meaningful way.

4.10.2 Determining the shape of a molecule

To predict the shape of a covalent molecule, follow these steps:

Step 1:

Draw the molecule using Lewis notation. Make sure that you draw *all* the electrons around the molecule's central atom.

Step 2:

Count the number of electron pairs around the central atom.

Step 3:

Determine the basic geometry of the molecule using the table below. For example, a molecule with two electron pairs around the central atom has a *linear* shape, and one with four electron pairs around the central atom would have a *tetrahedral* shape. The situation is actually more

Table 4.3: The effect of electron pairs in determining the shape of molecules

Number of electron pairs	Geometry
2	linear
3	trigonal planar
4	tetrahedral
5	trigonal bipyramidal
6	octahedral

complicated than this, but this will be discussed later in this section.

Figure 4.12 shows each of these shapes. Remember that the shapes are 3-dimensional, and so you need to try to imagine them in this way. In the diagrams, the shaded part represents those parts of the molecule that are 'in front', while the dashed lines represent those parts that are 'at the back' of the molecule.

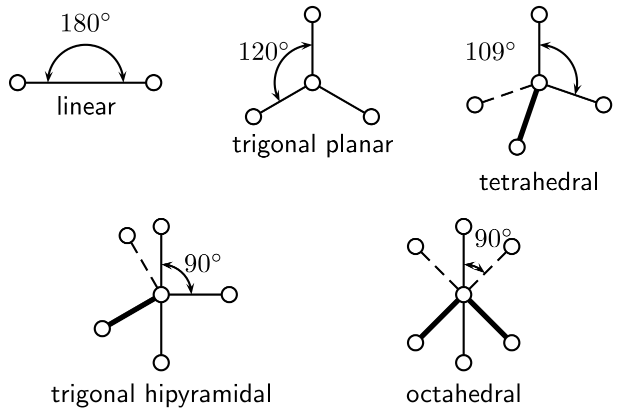


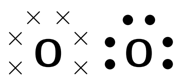
Figure 4.12: Some common molecular shapes

Worked Example 17: The shapes of molecules

Question: Determine the shape of a molecule of O_2

Answer

Step 1 : Draw the molecule using Lewis notation



Step 2 : Count the number of electron pairs around the central atom

There are two electron pairs.

Step 3 : Determine the basic geometry of the molecule

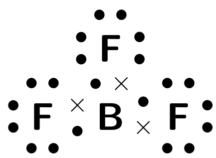
Since there are two electron pairs, the molecule must be linear.

Worked Example 18: The shapes of molecules

Question: Determine the shape of a molecule of BF_3

Answer

Step 1 : Draw the molecule using Lewis notation



Step 2 : Count the number of electron pairs around the central atom

There are three electron pairs.

Step 3 : Determine the basic geometry of the molecule

Since there are three electron pairs, the molecule must be trigonal planar.

Extension: More about molecular shapes

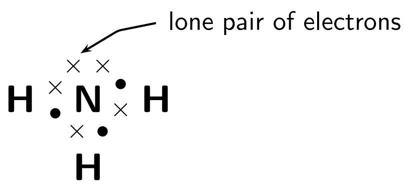
Determining the shape of a molecule can be a bit more complicated. In the examples we have used above, we looked only at the number of **bonding electron pairs** when we were trying to decide on the molecules' shape. But there are also other electron pairs in the molecules. These electrons, which are not involved in bonding but which are also around the central atom, are called **lone pairs**. The worked example below will give you an indea of how these lone pairs can affect the shape of the molecule.

Worked Example 19: Advanced

Question: Determine the shape of a molecule of NH_3

Answer

Step 1 : Draw the molecule using Lewis notation



Step 2 : Count the number of electron pairs around the central atom

There are four electron pairs.

Step 3 : Determine the basic geometry of the molecule
Since there are four electron pairs, the molecule must be tetrahedral.

Step 4 : Determine how many lone pairs are around the central atom
There is one lone pair of electrons and this will affect the shape of the molecule.

Step 5 : Determine the final shape of the molecule
The lone pair needs more space than the bonding pairs, and therefore pushes the three hydrogen atoms together a little more. The bond angles between the hydrogen and nitrogen atoms in the molecule become 106 degrees, rather than the usual 109 degrees of a tetrahedral molecule. The shape of the molecule is *trigonal pyramidal*.

Activity :: Group work : Building molecular models
In groups, you are going to build a number of molecules using marshmallows to represent the atoms in the molecule, and toothpicks to represent the bonds between the atoms. In other words, the toothpicks will hold the atoms (marshmallows) in the molecule together. Try to use different coloured marshmallows to represent different elements.
You will build models of the following molecules:
HCl, CH_4, H_2O, HBr and NH_3
For each molecule, you need to:

- Determine the basic geometry of the molecule
- Build your model so that the atoms are as far apart from each other as possible (remember that the electrons around the central atom will try to avoid the repulsions between them).
- Decide whether this shape is accurate for that molecule or whether there are any lone pairs that may influence it.
- Adjust the position of the atoms so that the bonding pairs are further away from the lone pairs.
- How has the shape of the molecule changed?
- Draw a simple diagram to show the shape of the molecule. It doesn't matter if it is not 100% accurate. This exercise is only to help you to visualise the 3-dimensional shapes of molecules.

Do the models help you to have a clearer picture of what the molecules look like? Try to build some more models for other molecules you can think of.

4.11 Oxidation numbers

When reactions occur, an exchange of electrons takes place. **Oxidation** is the *loss* of electrons from an atom, while **reduction** is the *gain* of electrons by an atom. By giving elements an oxidation number, it is possible to keep track of whether that element is losing or gaining electrons during a chemical reaction. The loss of electrons in one part of the reaction, must be balanced by a gain of electrons in another part of the reaction.

Definition: Oxidation number
A simplified way of understanding an oxidation number is to say that it is the charge an atom would have if it was in a compound composed of ions.

There are a number of rules that you need to know about oxidation numbers, and these are listed below. These will probably not make much sense at first, but once you have worked through some examples, you will soon start to understand!

1. **Rule 1:** An element always has an oxidation number of zero, since it is neutral.

   In the reaction $H_2 + Br_2 \rightarrow 2HBr$, the oxidation numbers of hydrogen and bromine on the left hand side of the equation are both zero.
2. **Rule 2:** In most cases, an atom that is part of a molecule will have an oxidation number that has the same numerical value as its valency.
3. **Rule 3:** Monatomic ions have an oxidation number that is equal to the charge on the ion.

   The chloride ion Cl^- has an oxidation number of -1, and the magnesium ion Mg^{2+} has an oxidation number of $+2$.
4. **Rule 4:** In a molecule, the oxidation number for the whole molecule will be zero, unless the molecule has a charge, in which case the oxidation number is equal to the charge.
5. **Rule 5:** Use a table of electronegativities to determine whether an atom has a positive or a negative oxidation number. For example, in a molecule of water, oxygen has a higher electronegativity so it must be negative because it attracts electrons more strongly. It will have a negative oxidation number (-2). Hydrogen will have a positive oxidation number ($+1$).
6. **Rule 6:** An oxygen atom usually has an oxidation number of -2, although there are some cases where its oxidation number is -1.
7. **Rule 7:** The oxidation number of hydrogen is usually $+1$. There are some exceptions where its oxidation number is -1.
8. **Rule 8:** In most compounds, the oxidation number of the halogens is -1.

Important: You will notice that the oxidation number of an atom is the same as its valency. Whether an oxidation number os positive or negative, is determined by the electronegativities of the atoms involved.

Worked Example 20: Oxidation numbers

Question: Give the oxidation numbers for all the atoms in the reaction between sodium and chlorine to form sodium chloride.

$$Na + Cl \rightarrow NaCl$$

Answer

Step 1 : Determine which atom will have a positive or negative oxidation number

Sodium will have a positive oxidation number and chlorine will have a negative oxidation number.

Step 2 : Determine the oxidation number for each atom

Sodium (group 1) will have an oxidation number of $+1$. Chlorine (group 7) will have an oxidation number of -1.

Step 3 : Check whether the oxidation numbers add up to the charge on the molecule

In the equation $Na + Cl \rightarrow NaCl$, the overall charge on the NaCl molecule is +1-1=0. This is correct since NaCl is neutral. This means that, in a molecule of NaCl, sodium has an oxidation number of +1 and chlorine has an oxidation number of -1.

Worked Example 21: Oxidation numbers

Question: Give the oxidation numbers for all the atoms in the reaction between hydrogen and oxygen to produce water. The unbalanced equation is shown below:

$$H_2 + O_2 \rightarrow H_2O$$

Answer

Step 1 : Determine which atom will have a positive or negative oxidation number

Hydrogen will have a positive oxidation number and oxygen will have a negative oxidation number.

Step 2 : Determine the oxidation number for each atom

Hydrogen (group 1) will have an oxidation number of $+1$. Oxygen (group 6) will have an oxidation number of -2.

Step 3 : Check whether the oxidation numbers add up to the charge on the molecule

In the reaction $H_2 + O_2 \rightarrow H_2O$, the oxidation numbers for hydrogen and oxygen (on the left hand side of the equation) are zero since these are elements. In the water molecule, the sum of the oxidation numbers is 2(+1)-2=0. This is correct since the oxidation number of water is zero. Therefore, in water, hydrogen has an oxidation number of +1 and oxygen has an oxidation number of -2.

Worked Example 22: Oxidation numbers

Question: Give the oxidation number of sulfur in a sulphate (SO_4^{2-}) ion

Answer

Step 1 : Determine which atom will have a positive or negative oxidation number

Sulfur has a positive oxidation number and oxygen will have a negative oxidation number.

Step 2 : Determine the oxidation number for each atom

Oxygen (group 6) will have an oxidation number of -2. The oxidation number of sulfur at this stage is uncertain.

Step 3 : Determine the oxidation number of sulfur by using the fact that the oxidation numbers of the atoms must add up to the charge on the molecule

In the polyatomic SO_4^{2-} ion, the sum of the oxidation numbers must be -2. Since there are four oxygen atoms in the ion, the total charge of the oxygen is -8. If the overall charge of the ion is -2, then the oxidation number of sulfur must be $+6$.

?

Exercise: Oxidation numbers

1. Give the oxidation numbers for each element in the following chemical compounds:
   (a) NO_2
   (b) $BaCl_2$
   (c) H_2SO_4
2. Give the oxidation numbers for the reactants and products in each of the following reactions:
   (a) $C + O_2 \rightarrow CO_2$
   (b) $N_2 + 3H_2 \rightarrow 2NH_3$
   (c) Magnesium metal burns in oxygen

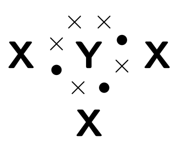
4.12 Summary

- A **chemical bond** is the physical process that causes atoms and molecules to be attracted together and to be bound in new compounds.
- Atoms are more **reactive**, and therefore more likely to bond, when their outer electron orbitals are not full. Atoms are less reactive when these outer orbitals contain the maximum number of electrons. This explains why the noble gases do not combine to form molecules.
- There are a number of **forces** that act between atoms: attractive forces between the positive nucleus of one atom and the negative electrons of another; repulsive forces between like-charged electrons, and repulsion between like-charged nuclei.
- Chemical bonding occurs when the **energy** of the system is at its lowest.
- **Bond length** is the distance between the nuclei of the atoms when they bond.
- **Bond energy** is the energy that must be added to the system for the bonds to break.
- When atoms bond, electrons are either shared or exchanged.
- **Covalent bonding** occurs between the atoms of non-metals and involves a sharing of electrons so that the orbitals of the outermost energy levels in the atoms are filled.
- The **valency** of an atom is the number of bonds that it can form with other atoms.
- A **double** or **triple bond** occurs if there are two or three electron pairs that are shared between the same two atoms.
- A **dative covalent bond** is a bond between two atoms in which both the electrons that are shared in the bond come from the same atom.
- **Lewis** and **Couper** notation are two ways of representing molecular structure. In Lewis notation, dots and crosses are used to represent the valence electrons around the central atom. In Couper notation, lines are used to represent the bonds between atoms.
- **Electronegativity** measures how strongly an atom draws electrons to it.
- Electronegativity can be used to explain the difference between two types of covalent bonds: **polar covalent bonds** (between non-identical atoms) and **non-polar covalent bonds** (between identical atoms).
- An **ionic bond** occurs between atoms where the difference in electronegativity is greater than 2.1. An exchange of electrons takes place and the atoms are held together by the electrostatic force of attraction between oppositely-charged ions.
- Ionic solids are arranged in a **crystal lattice** structure.

- Ionic compounds have a number of specific **properties**, including their high melting and boiling points, brittle nature, the arrangement of solids in a lattice structure and the ability of ionic solutions to conduct electricity.
- A **metallic bond** is the electrostatic attraction between the positively charged nuclei of metal atoms and the delocalised electrons in the metal.
- Metals also have a number of properties, including their ability to conduct heat and electricity, their metallic lustre, the fact that they are both malleable and ductile, and their high melting point and density.
- The valency of atoms, and the way they bond, can be used to determine the **chemical formulae** of compounds.
- The **shape of molecules** can be predicted using the VSEPR theory, which uses the arrangement of electrons around the central atom to determine the most likely shape of the molecule.
- **Oxidation numbers** are used to determine whether an atom has gained or lost electrons during a chemical reaction.

Exercise: Summary exercise

1. Give **one word/term** for each of the following descriptions.
   (a) The distance between two atoms in a molecule
   (b) A type of chemical bond that involves the transfer of electrons from one atom to another.
   (c) A measure of an atom's ability to attract electrons to it.
2. Which ONE of the following best describes the bond formed between an H^+ ion and the NH_3 molecule?
   (a) Covalent bond
   (b) Dative covalent (coordinate covalent) bond
   (c) Ionic Bond
   (d) Hydrogen Bond
3. Explain the meaning of each of the following terms:
   (a) valency
   (b) bond energy
   (c) covalent bond
4. Which of the following reactions will *not* take place? Explain your answer.
   (a) $\mathrm{H} + \mathrm{H} \rightarrow \mathrm{H_2}$
   (b) $\mathrm{Ne} + \mathrm{Ne} \rightarrow \mathrm{Ne_2}$
   (c) $\mathrm{I} + \mathrm{I} \rightarrow \mathrm{I_2}$
5. Draw the Lewis structure for each of the following:
   (a) calcium
   (b) iodine
   (c) hydrogen bromide (HBr)
   (d) nitrogen dioxide (NO_2)
6. Given the following Lewis structure, where X and Y each represent a different element...



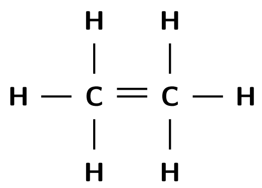
   (a) What is the valency of X?

(b) What is the valency of Y?

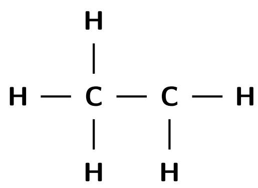
(c) Which elements could X and Y represent?

7. A molecule of ethane has the formula C_2H_6. Which of the following diagrams (Couper notation) accurately represents this molecule?

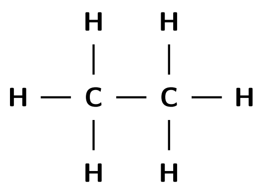
(a)



(b)



(c)



8. Potassium dichromate is dissolved in water.

   (a) Give the name and chemical formula for each of the ions in solution.

   (b) What is the chemical formula for potassium dichromate?

   (c) Give the oxidation number for each element in potassium dichromate.

Chapter 5

Intermolecular Forces - Grade 11

In the previous chapter, we discussed the different forces that exist *between atoms* (intramolecular forces). When atoms are joined to one another they form molecules, and these molecules in turn have forces that bind them together. These forces are known as **intermolecular forces**, and we are going to look at them in more detail in this next section.

> **Definition: Intermolecular forces**
> Intermolecular forces are forces that act between stable molecules.

You will also remember from the previous chapter, that we can describe molecules as being either **polar** or **non-polar**. A polar molecule is one in which there is a difference in electronegativity between the atoms in the molecule, such that the shared electron pair spends more time close to the atom that attracts it more strongly. The result is that one end of the molecule will have a slightly positive charge (δ^+), and the other end will have a slightly negative charge (δ^+). The molecule is said to be a **dipole**. However, it is important to remember that just because the bonds within a molecule are polar, the molecule itself may not necessarily be polar. The shape of the molecule may also affect its polarity. A few examples are shown in table 5.1 to refresh your memory!

5.1 Types of Intermolecular Forces

It is important to be able to recognise whether the molecules in a substance are polar or non-polar because this will determine what type of inermolecular forces there are. This is important in explaining the properties of the substance.

1. **Van der Waals forces**

   These intermolecular forces are named after a Dutch physicist called Johannes van der Waals (1837 -1923), who recognised that there were weak attractive and repulsive forces between the molecules of a gas, and that these forces caused gases to deviate from 'ideal gas' behaviour. Van der Waals forces are *weak* intermolecular forces, and can be divided into three types:

   (a) *Dipole-dipole forces*

   Figure 5.1 shows a simplified dipole molecule, with one end slightly positive and the other slightly negative.

   When one dipole molecule comes into contact with another dipole molecule, the positive pole of the one molecule will be attracted to the negative pole of the other, and the molecules will be held together in this way (figure 5.2). Examples of molecules that are held together by dipole-dipole forces are HCl, FeS, KBr, SO_2 and NO_2.

   (b) *Ion-dipole forces*

   As the name suggests, this type of intermolecular force exists between an ion and a dipole molecule. You will remember that an *ion* is a charged atom, and this will

Table 5.1: Polarity in molecules with different atomic bonds and molecular shapes

Molecule	**Chemical formula**	**Bond between atoms**	**Shape of molecule**	**Polarity of molecule**
Hydrogen	H_2	Covalent	Linear molecule **H** —— **H**	Non-polar
Hydrogen chloride	HCl	Polar covalent	Linear molecule $\mathbf{H}^{\delta+}$ —— $\mathbf{Cl}^{\delta-}$	Polar
Carbon tetrafluoromethane	CF_4	Polar covalent	Tetrahedral molecule $F^{\delta-}$ | $F^{\delta-}$ —— $C^{\delta+}$ —— $F^{\delta-}$ | $F^{\delta-}$	Non-polar

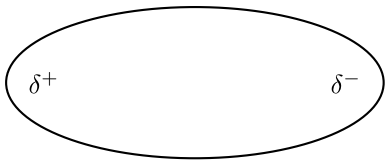


Figure 5.1: A simplified diagram of a dipole molecule

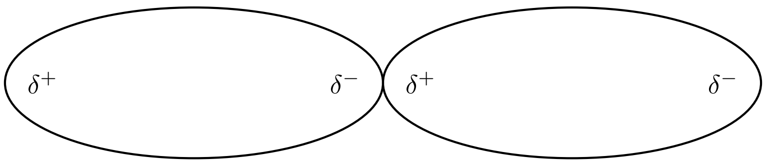


Figure 5.2: Two dipole molecules are held together by the attractive force between their oppositely charged poles

be attracted to one of the charged ends of the polar molecule. A positive ion will be attracted to the negative pole of the polar molecule, while a negative ion will be attracted to the positive pole of the polar molecule. This can be seen when sodium chloride (NaCl) dissolves in water. The positive sodium ion (Na^+) will be attracted to the slightly negative oxygen atoms in the water molecule, while the negative chloride ion (Cl^-) is attracted to the slightly positive hydrogen atom. These intermolecular forces weaken the ionic bonds between the sodium and chloride ions so that the sodium chloride dissolves in the water (figure 5.3).

(c) *London forces*

These intermolecular forces are also sometimes called 'dipole- induced dipole' or 'momentary dipole' forces. Not all molecules are polar, and yet we know that there are also intermolecular forces between non-polar molecules such as carbon dioxide. In

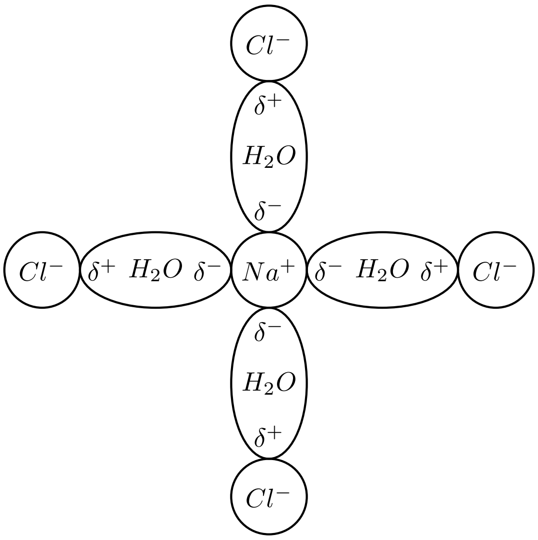


Figure 5.3: Ion-dipole forces in a sodium chloride solution

non-polar molecules the electronic charge is evenly distributed but it is possible that at a particular moment in time, the electrons might not be evenly distributed. The molecule will have a *temporary dipole*. In other words, each end of the molecules has a slight charge, either positive or negative. When this happens, molecules that are next to each other attract each other very weakly. These London forces are found in the halogens (e.g. F_2 and I_2), the noble gases (e.g. Ne and Ar) and in other non-polar molecules such as carbon dioxide and methane.

2. **Hydrogen bonds**

   As the name implies, this type of intermolecular bond involves a hydrogen atom. The hydrogen must be attached to another atom that is strongly electronegative, such as oxygen, nitrogen or fluorine. Water molecules for example, are held together by hydrogen bonds between the hydrogen atom of one molecule and the oxygen atom of another (figure 5.4). Hydrogen bonds are stronger than van der Waals forces. It is important to note however, that both van der Waals forces and hydrogen bonds are weaker than the covalent and ionic bonds that exist between *atoms*.

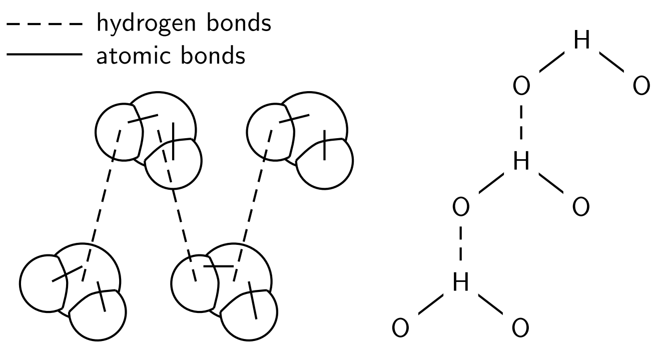


Figure 5.4: Two representations showing the hydrogen bonds between water molecules: space-filling model and structural formula.

Exercise: Types of intermolecular forces

1. Complete the following table by placing a tick to show which type of intermolecular force occurs in each substance:

Formula	**Dipole-dipole**	**Momentary dipole**	**Ion-dipole**	**hydrogen bond**
HCl				
CO_2				
I_2				
H_2O				
KI(aq)				
NH_3				

2. In which of the substances above are the intermolecular forces...
   (a) strongest
   (b) weakest

5.2 Understanding intermolecular forces

The types of intermolecular forces that occur in a substance will affect its properties, such as its **phase**, **melting point** and **boiling point**. You should remember, if you think back to the kinetic theory of matter, that the *phase* of a substance is determined by how strong the forces are between its particles. The weaker the forces, the more likely the substance is to exist as a gas because the particles are far apart. If the forces are very strong, the particles are held closely together in a solid structure. Remember also that the *temperature* of a material affects the energy of its particles. The more energy the particles have, the more likely they are to be able to overcome the forces that are holding them together. This can cause a change in phase.

Definition: Boiling point
The temperature at which a material will change from being a liquid to being a gas.

Definition: Melting point
The temperature at which a material will change from being a solid to being a liquid.

Now look at the data in table 5.2.

Formula	**Formula mass**	**Melting point** (^{0}C)	**Boiling point** (^{0}C) at 1 atm
He	4	-270	-269
Ne	20	-249	-246
Ar	40	-189	-186
F_2	38	-220	-188
Cl_2	71	-101	-35
Br_2	160	-7	58
NH_3	17	-78	-33
H_2O	18	0	100
HF	20	-83	20

Table 5.2: Melting point and boiling point of a number of chemical substances

The melting point and boiling point of a substance, give us information about the *phase* of the substance at room temperature, and the *strength of the intermolecular forces*. The examples below will help to explain this.

Example 1: Fluorine (F_2)

Phase at room temperature

Fluorine (F_2) has a melting point of -220^{0}C and a boiling point of -188^{0}C. This means that for any temperature that is greater than -188^{0}C, fluorine will be a gas. At temperatures below -220^{0}C, fluorine would be a solid, and at any temperature inbetween these two, fluorine will be a liquid. So, at room temperature, fluorine exists as a gas.

Strength of intermolecular forces

What does this information tell us about the intermolecular forces in fluorine? In fluorine, these forces must be very weak for it to exist as a gas at room temperature. Only at temperatures below -188^{0}C will the molecules have a low enough energy that they will come close enough to each other for forces of attraction to act between the molecules. The intermolecular forces in fluorine are very weak **van der Waals** forces because the molecules are *non-polar*.

Example 2: Hydrogen fluoride (HF)

Phase at room temperature

For temperatures below -83^{0}C, hydrogen fluoride is a solid. Between -83^{0}C and 20^{0}C, it exists as a liquid, and if the temperature is increased above 20^{0}C, it will become a gas.

Strength of intermolecular forces

What does this tell us about the intermolecular forces in hydrogen fluoride? The forces are stronger than those in fluorine, because more energy is needed for fluorine to change into the gaseous phase. In other words, more energy is needed for the intermolecular forces to be overcome so that the molecules can move further apart. Intermolecular forces will exist between the hydrogen atom of one molecule and the fluorine atom of another. These are **hydrogen bonds**, which are stronger than van der Waals forces.

What do you notice about water? Luckily for us, water behaves quite differently from the rest of the halides. Imagine if water were like ammonia (NH_3), which is a gas above a temperature of -33^{0}C! There would be no liquid water on the planet, and that would mean that no life would be able to survive here. The hydrogen bonds in water are particularly strong and this gives water unique qualities when compared to other molecules with hydrogen bonds. This will be discussed more in chapter **??** deals with this in more detail. You should also note that the strength of the intermolecular forces increases with an increase in formula mass. This can be seen by the increasing melting and boiling points of substances as formula mass increases.

?

Exercise: Applying your knowledge of intermolecular forces

Refer to the data in table 5.2 and then use your knowledge of different types of intermolecular forces to explain the following statements:

- The boiling point of F_2 is much lower than the boiling point of NH_3
- At room temperature, many elements exist naturally as gases
- The boiling point of HF is higher than the boiling point of Cl_2
- The boiling point of water is much higher than HF, even though they both contain hydrogen bonds

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5.3 Intermolecular forces in liquids

Intermolecular forces affect a number of properties in liquids:

- **Surface tension**

  You may have noticed how some insects are able to walk across a water surface, and how leaves float in water. This is because of surface tension. In water, each molecule is held to the surrounding molecules by strong hydrogen bonds. Molecules in the centre of the liquid are completely surrounded by other molecules, so these forces are exerted in all directions. However, molecules at the surface do not have any water molecules above them to pull them upwards. Because they are only pulled sideways and downwards, the water molecules at the surface are held more closely together. This is called **surface tension**.

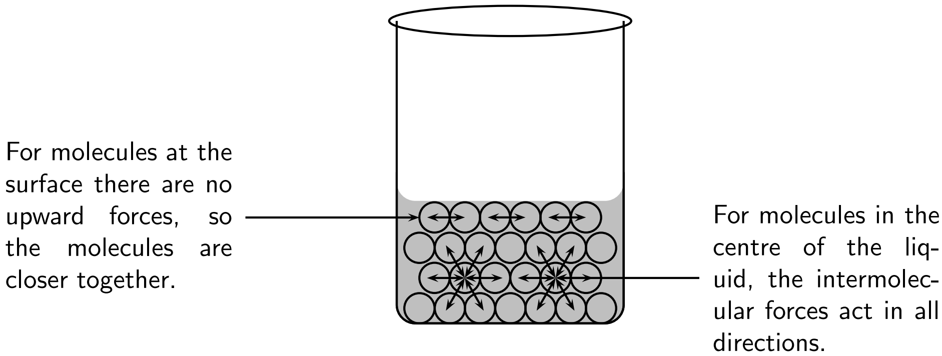
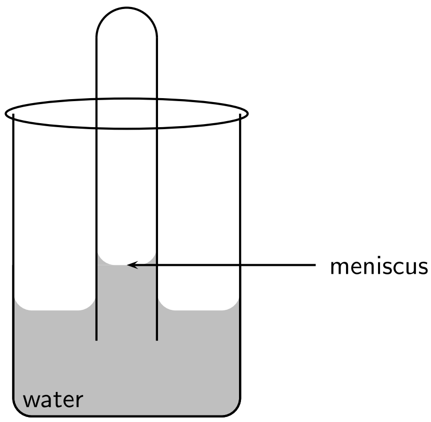


Figure 5.5: Surface tension in a liquid

- **Capillarity**

Activity :: Investigation : Capillarity
Half fill a beaker with water and hold a hollow glass tube in the centre as shown below. Mark the level of the water in the glass tube, and look carefully at the shape of the air-water interface in the tube. What do you notice?



At the air-water interface, you will notice a **meniscus**, where the water appears to dip in the centre. In the glass tube, the attractive forces between the glass and the water are stronger than the intermolecular forces between the water molecules. This causes the water to be held more closely to the glass, and a meniscus forms. The forces between the glass and the water also mean that the water can be 'pulled up' higher when it is in the tube than when it is in teh beaker. Capillarity is the surface tension that occurs in liquids that are inside tubes.

- **Evaporation**

  In a liquid, each particle has kinetic energy, but some particles will have more energy than others. We therefore refer to the *average* kinetic energy of the molecules when we describe the liquid. When the liquid is heated, those particles which have the highest energy will be able to overcome the intermolecular forces holding them in the liquid phase, and will become a gas. This is called **evaporation**. Evaporation occurs when a liquid changes to a gas. The stronger the intermolecular forces in a liquid, the higher the temperature of the molecules will have to be for it to become a gas. You should note that a liquid doesn't necessarily have to reach boiling point before evaporation can occur. Evaporation takes place all the time. You will see this if you leave a glass of water outside in the sun. Slowly the water level will drop over a period of time.

  What happens then to the molecules of water that remain in the liquid? Remember that it was the molecules with the highest energy that left the liquid. This means that the average kinetic energy of the remaining molecules will decrease, and so will the *temperature* of the liquid.

  A similar process takes place when a person sweats during exercise. When you exercise, your body temperature increases and you begin to release moisture (sweat) through the pores in your skin. The sweat quickly evaporates and causes the temperature of your skin to drop. This helps to keep your body temperature at a level that is suitable forit to function properly.

Interesting Fact

Transpiration in plants - Did you know that plants also 'sweat'? In plants, this is called *transpiration*, and a plant will lose water through spaces in the leaf surface called *stomata*. Although this water loss is important in the survival of a plant, if a plant loses too much water, it will die. Plants that live in very hot, dry places such as deserts, must be specially adapted to reduce the amount of water that transpires (evaporates) from their leaf surface. Desert plants have some amazing adaptations to deal with this problem! Some have hairs on their leaves, which reflect sunlight so that the temperature is not as high as it would be, while others have a thin waxy layer covering their leaves, which reduces water loss. Some plants are even able to close their stomata during the day when temperatures (and therefore transpiration) are highest.

Important: In the same way that intermolecular forces affect the properties of liquids, they also affect the properties of solids. For example, the stronger the intermolecular forces between the particles that make up the solid, the *harder* the solid is likely to be, and the higher its *melting point* is likely to be.

5.4 Summary

- **Intermolecular forces** are the forces that act between stable molecules.
- The **type** of intermolecular force in a substance, will depend on the **nature of the molecules**.
- **Polar molecules** have an unequal distribution of charge, meaning that one part of the molecule is slightly positive and the other part is slightly negative. **Non-polar molecules** have an equal distribution of charge.

- There are three types of **Van der Waal's forces**. These are dipole-dipole, ion-dipole and London forces (momentary dipole).
- **Dipole-dipole** forces exist between two **polar molecules**, for example between two molecules of hydrogen chloride.
- **Ion-dipole** forces exist between **ions and dipole molecules**. The ion is attracted to the part of the molecule that has an opposite charge to its own. An example of this is when an ionic solid such as sodium chloride dissolves in water.
- **Momentary dipole** forces occur between two **non-polar molecules**, where at some point there is an uequal distribution of charge in the molecule. For example, there are London forces between two molecules of carbon dioxide.
- **Hydrogen bonds** occur between **hydrogen atoms** and other **atoms that have a high electronegativity** such as oxygen, nitrogen and fluorine. The hydrogen atom in one molecule will be attracted to the nitrogen atom in another molecule, for example. There are hydrogen bonds between water molecules and between ammonia molecules.
- Intermolecular forces affect the **properties** of substances. For example, the stronger the intermolecular forces, the higher the **melting point** of that substance, and the more likely that substance is to exist as a solid or liquid. Its **boiling point** will also be higher.
- In **liquids**, properties such as **surface tension**, **capillarity** and **evaporation** are the result of intermolecular forces.

?

Exercise: Summary Exercise

1. Give one word or term for each of the following descriptions:
   (a) The tendency of an atom in a molecule to attract bonding electrons.
   (b) A molecule that has an unequal distribution of charge.
   (c) A charged atom.
2. For each of the following questions, choose the one correct answer from the list provided.
   (a) The following table gives the melting points of various hydrides:

Hydride	**Melting point** (^{0}C)
HI	-50
NH_3	-78
H_2S	-83
CH_4	-184

   In which of these hydrides does hydrogen bonding occur?
   i. HI only
   ii. NH_3 only
   iii. HI and NH_3 only
   iv. HI, NH_3 and H_2S
   (IEB Paper 2, 2003)
   (b) Refer to the list of substances below:
   HCl, Cl_2, H_2O, NH_3, N_2, HF
   Select the true statement from the list below:
   i. NH_3 is a non-polar molecule
   ii. The melting point of NH_3 will be higher than for Cl_2
   iii. Ion-dipole forces exist between molecules of HF
   iv. At room temperature N_2 is usually a liquid

3. The respective boiling points for four chemical substances are given below:
   Hydrogen sulphide -60^{0}C
   Ammonia -33^{0}C
   Hydrogen fluoride 20^{0}C
   Water 100^{0}C
   (a) Which one of the substances exhibits the strongest forces of attraction between its molecules in the liquid state?
   (b) Give the name of the force responsible for the relatively high boiling points of ammonia and water and explain how this force originates.
   (c) The shapes of the molecules of hydrogen sulphide and water are similar, yet their boiling points differ. Explain.

   (IEB Paper 2, 2002)

Chapter 6

Solutions and solubility - Grade 11

We are surrounded by different types of solutions in our daily lives. Any solution is made up of a **solute** and a **solvent**. A **solute** is a substance that dissolves in a solvent. In the case of a salt (NaCl) solution, the salt crystals are the solute. A **solvent** is the substance in which the solute dissolves. In the case of the NaCl solution, the solvent would be the water. In most cases, there is always more of the solvent than there is of the solute in a solution.

Definition: Solutes and solvents
A **solute** is a substance that is dissolved in another substance. A solute can be a solid, liquid or gas. A **solvent** is the liquid that dissolves a solid, liquid, or gaseous solute.

6.1 Types of solutions

When a solute is mixed with a solvent, a *mixture* is formed, and this may be either *heterogeneous* or *homogeneous*. If you mix sand and water for example, the sand does not dissolve in the water. This is a **heterogeneous** mixture. When you mix salt and water, the resulting mixture is **homogeneous** because the solute has dissolved in the solvent.

Definition: Solution
In chemistry, a solution is a homogeneous mixture that consists of a solute that has been dissolved in a solvent.

A solution then is a homogeneous mixture of a solute and a solvent. Examples of solutions are:

- A *solid solute* dissolved in a *liquid solvent* e.g. sodium chloride dissolved in water.
- A *gas solute* dissolved in a *liquid solvent* e.g. carbon dioxide dissolved in water (fizzy drinks) or oxygen dissolved in water (aquatic ecosystems).
- A *liquid solute* dissolved in a *liquid solvent* e.g. ethanol in water.
- A *solid solute* in a *solid solvent* e.g. metal alloys.
- A *gas solute* in a *gas solvent* e.g. the homogeneous mixture of gases in the air that we breathe.

While there are many different types of solutions, most of those we will be discussing are *liquids*.

6.2 Forces and solutions

An important question to ask is why some solutes dissolve in certain solvents and not in others. The answer lies in understanding the interaction between the intramolecular and intermolecular forces between the solute and solvent particles.

Activity :: Experiment : Solubility

Aim:

To investigate the solubility of solutes in different solvents.

Apparatus:

Salt, vinegar, iodine, ethanol

Method:

1. Mix half a teaspoon of salt in 100cm^3 of water
2. Mix half a teaspoon of vinegar (acetic acid) in 100cm^3 of water
3. Mix a few grains of iodine in ethanol
4. Mix a few grains of iodine in 100cm^3 of water

Results:

Record your observations in the table below:

Solute	Polar, non-polar or ionic solute	Solvent	Polar, non-polar or ionic solvent	Does solute dissolve?
Iodine		Ethanol		
Iodine		Water		
Vinegar		Water		
Salt		Water		

You should have noticed that in some cases, the solute dissolves in the solvent, while in other cases it does not.

Conclusions:

In general, polar and ionic solutes dissolve well in polar solvents, while non-polar solutes dissolve well in non-polar solvents. An easy way to remember this is that 'like dissolves like', in other words, if the solute and the solvent have similar intermolecular forces, there is a high possibility that dissolution will occur. This will be explained in more detail below.

- **Non-polar solutes and non-polar solvents** (e.g. iodine and ether)

  Iodine molecules are non-polar, and the forces between the molecules are weak van der Waals forces. There are also weak van der Waals forces between ether molecules. Because the intermolecular forces in both the solute and the solvent are similar, it is easy for these to be broken in the solute, allowing the solute to move into the spaces between the molecules of the solvent. The solute dissolves in the solvent.

- **Polar solutes and polar solvents** (e.g. salt and water)

  There are strong electrostatic forces between the ions of a salt such as sodium chloride. There are also strong hydrogen bonds between water molecules. Because the strength of the intermolecular forces in the solute and solvent are similar, the solute will dissolve in the solvent.

6.3 Solubility

You may have noticed sometimes that, if you try to dissolve salt (or some other solute) in a small amount of water, it will initially dissolve, but then appears not to be able to dissolve any further when you keep adding more solute to the solvent. This is called the **solubility** of the solution. Solubility refers to the maximum amount of solute that will dissolve in a solvent under certain conditions.

Definition: Solubility
Solubility is the ability of a given substance, the solute, to dissolve in a solvent. If a substance has a high solubility, it means that lots of the solute is able to dissolve in the solvent.

So what factors affect solubility? Below are some of the factors that affect solubility:

- the quantity of solute and solvent in the solution
- the temperature of the solution
- other compounds in the solvent affect solubility because they take up some of the spaces between molecules of the solvent, that could otherwise be taken by the solute itself
- the strength of the forces between particles of the solute, and the strength of forces between particles of the solvent

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Activity :: Experiment : Factors affecting solubility
Aim:
To determine the effect of temperature on solubility
Method:

1. Measure 100cm^3 of water into a beaker
2. Measure 100 g of salt and place into another beaker
3. Slowly pour the salt into the beaker with the water, stirring it as you add. Keep adding salt until you notice that the salt is not dissolving anymore.
4. Record the amount of salt that has been added to the water and the temperature of the solution.
5. Now increase the temperature of the water by heating it over a bunsen burner.
6. Repeat the steps above so that you obtain the solubility limit of salt at this higher temperature. You will need to start again with new salt and water!
7. Continue to increase the temperature as many times as possible and record your results.

Results:

Record your results in the table below:

Temp (^{0}C)	**Amount of solute that dissolves in 100 cm^3 of water** (g)

As you increase the temperature of the water, are you able to dissolve *more* or *less* salt?

Conclusions:

As the temperature of the solution increases, so does the amount of salt that will dissolve. The solubility of sodium chloride increases as the temperature increases.

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Exercise: Investigating the solubility of salts

The data table below gives the solubility (measured in grams of salt per 100 g water) of a number of different salts at various temperatures. Look at the data and then answer the questions that follow.

	Solubility (g salt per 100 g H_2O)		
Temp (^{0}C)	**KNO_3**	**K_2SO_4**	**NaCl**
0	13.9	7.4	35.7
10	21.2	9.3	35.8
20	31.6	11.1	36.0
30	45.3	13.0	36.2
40	61.4	14.8	36.5
50	83.5	16.5	36.8
60	106.0	18.2	37.3

1. On the same set of axes, draw line graphs to show how the solubility of the three salts changes with an increase in temperature.
2. Describe what happens to salt solubility as temperature increases. Suggest a reason why this happens.
3. Write an equation to show how each of the following salts ionises in water:
   (a) KNO_3
   (b) K_2SO_4
4. You are given three beakers, each containing the same amount of water. 5 g KNO_3 is added to beaker 1, 5 g K_2SO_4 is added to beaker 2 and 5 g NaCl is added to beaker 3. The beakers are heated over a bunsen burner until the temperature of their solutions is 60^{0}C.
   (a) Which salt solution will have the highest conductivity under these conditions?
   (b) Explain your answer.

?

Exercise: Experiments and solubility

Two grade 10 learners, Siphiwe and Ann, wish to separately investigate the solubility of potassium chloride at room temperature. They follow the list of instructions shown below, using the apparatus that has been given to them:

Method:

1. Determine the mass of an empty, dry evaporating basin using an electronic balance and record the mass.
2. Pour 50 ml water into a 250 ml beaker.

3. Add potassium chloride crystals to the water in the beaker in small portions.
4. Stir the solution until the salt dissolves.
5. Repeat the addition of potassium chloride (steps a and b) until no more salt dissolves and some salt remains undissolved.
6. Record the temperature of the potassium chloride solution.
7. Filter the solution into the evaporating basin.
8. Determine the mass of the evaporating basin containing the solution that has passed through the filter (the filtrate) on the electronic balance and record the mass.
9. Ignite the Bunsen burner.
10. Carefully heat the filtrate in the evaporating basin until the salt is dry.
11. Place the evaporating basin in the desiccator (a large glass container in which there is a dehydrating agent like calcium sulphate that absorbs water) until it reaches room temperature.
12. Determine the mass of the evaporating basin containing the dry cool salt on the electronic balance and record the mass.

On completion of the experiment, their results were as follows:

	Siphiwe's results	Ann's results
Temperature (^{0}C)	15	26
Mass of evaporating basin (g)	65.32	67.55
Mass of evaporating basin + salt solution (g)	125.32	137.55
Mass of evaporating basin + salt (g)	81.32	85.75

1. Calculate the solubility of potassium chloride, using the data recorded by
   (a) Siphiwe
   (b) Ann
   A reference book lists the solubility of potassium chloride as 35.0 g per 100 ml of water at 25^{0}C.
   (c) Give a reason why you think each obtained results different from each other and the value in the reference book.
2. Siphiwe and Ann now expand their investigation and work together. They investigate the solubility of potassium chloride at different temperatures but also the solubility of copper (II) sulphate at these same temperatures. They collect and write up their results as follows:

   In each experiment we used 50 ml of water in the beaker. We found the following masses of substance dissolved in the 50 ml of water. At 0^{0}C, mass of potassium chloride is 14.0 g and copper sulphate is 14.3 g. At 10^{0}C, 15.6 g and 17.4 g respectively. At 20^{0}C, 17.3 g and 20.7 g respectively. At 40^{0}C, potassium chloride mass is 20.2 g and copper sulphate is 28.5 g, at 60^{0}C, 23.1 g and 40.0 g and lastly at 80^{0}C, the masses were 26.4 g and 55.0 g respectively.

   (a) From the record of data provided above, draw up a neat table to record Siphiwe and Ann's results.
   (b) Identify the dependent and independent variables in their investigation.
   (c) Choose an appropriate scale and plot a graph of these results.
   (d) From the graph, determine:
      i. the temperature at which the solubility of copper sulphate is 50 g per 50 ml of water.
      ii. the maximum number of grams of potassium chloride which will dissolve in 100 ml of water at 70^{0}C.

   (IEB Exemplar Paper 2, 2006)

6.4 Summary

- In chemistry, a **solution** is a homogenous mixture of a solute in a solvent.
- A **solute** is a substance that dissolves in a solute. A solute can be a solid, liquid or gas.
- A **solvent** is a substance in which a solute dissolves. A solvent can also be a solid, liquid or gas.
- Examples of solutions include salt solutions, metal alloys, the air we breathe and gases such as oxygen and carbon dioxide dissolved in water.
- Not all solutes will dissolve in all solvents. A general rule is the **like dissolves like**. Solutes and solvents that have similar intermolecular forces are more likely to dissolve.
- Polar and ionic solutes will be more likely to dissolve in polar solvents, while non-polar solutes will be more likely to dissolve in polar solvents.
- **Solubility** is the extent to which a solute is able to dissolve in a solvent under certain conditions.
- Factors that affect solubility are the **quantity of solute and solvent**, **temperature**, the **intermolecular forces** in the solute and solvent and **other substances** that may be in the solvent.

?

Exercise: Summary Exercise

1. Give one word or term for each of the following descriptions:
   (a) A type of mixture where the solute has completely dissolved in the solvent.
   (b) A measure of how much solute is dissolved in a solution.
   (c) Forces between the molecules in a substance.
2. For each of the following questions, choose the one correct answer from the list provided.
   A Which one of the following will readily dissolve in water?
      i. I_2(s)
      ii. NaI(s)
      iii. CCl_4(l)
      iv. $BaSO_4$(s)
      (IEB Paper 2, 2005)
   b In which of the following pairs of substances will the dissolving process happen most readily?

	Solute	Solvent
A	S_8	H_2O
B	KCl	CCl_4
C	KNO_3	H_2O
D	NH_4Cl	CCl_4

   (IEB Paper 2, 2004)
3. Which one of the following three substances is the most soluble in pure water at room temperature?

   Hydrogen sulphide, ammonia and hydrogen fluoride
4. Briefly explain in terms of intermolecular forces why solid iodine does not dissolve in pure water, yet it dissolves in xylene, an organic liquid at room temperature.
   (IEB Paper 2, 2002)

Chapter 7

Atomic Nuclei - Grade 11

Nuclear physics is the branch of physics which deals with the **nucleus** of the atom. Within this field, some scientists focus their attention on looking at the *particles* inside the nucleus and understanding how they interact, while others classify and interpret the *properties* of nuclei. This detailed knowledge of the nucleus makes it possible for *technological advances* to be made. In this next chapter, we are going to touch on each of these different areas within the field of nuclear physics.

7.1 Nuclear structure and stability

You will remember from an earlier chapter that an atom is made up of different types of particles: protons (positive charge) neutrons (neutral) and electrons (negative charge). The nucleus is the part of the atom that contains the protons and the neutrons, while the electrons are found in energy orbitals around the nucleus. The protons and neutrons together are called **nucleons**. It is the nucleus that makes up most of an atom's *atomic mass*, because an electron has a very small mass when compared with a proton or a neutron.

Within the nucleus, there are different forces which act between the particles. The **strong nuclear force** is the force between two or more nucleons, and this force binds protons and neutrons together inside the nucleus. This force is most powerful when the nucleus is small, and the nucleons are close together. The **electromagnetic force** causes the repulsion between like-charged (positive) protons. In a way then, these forces are trying to produce opposite effects in the nucleus. The strong nuclear force acts to hold all the protons and neutrons close together, while the electromagnetic force acts to push protons further apart. In atoms where the nuclei are small, the strong nuclear force overpowers the electromagnetic force. However, as the nucleus gets bigger (in elements with a higher number of nucleons), the electromagnetic force becomes greater than the strong nuclear force. In these nuclei, it becomes possible for particles and energy to be ejected from the nucleus. These nuclei are called **unstable**. The particles and energy that a nucleus releases are referred to as **radiation**, and the atom is said to be **radioactive**. We are going to look at these concepts in more detail in the next few sections.

7.2 The Discovery of Radiation

Radioactivity was first discovered in 1896 by a French scientist called Henri Becquerel while he was working on phosphorescent materials. He happened to place some uranium crystals on black paper that he had used to cover a piece of film. When he looked more carefully, he noticed that the film had lots of patches on it, and that this did not happen when other elements were placed on the paper. He eventually concluded that some rays must be coming out of the uranium crystals to produce this effect.

His observations were taken further by the Polish scientist Marie Curie and her husband Pierre, who increased our knowledge of radioactive elements. In 1903, Henri, Marie and Pierre were awarded the Nobel Prize in Physics for their work on radioactive elements. This award made Marie the first woman ever to receive a Nobel Prize. Marie Curie and her husband went on to discover two new elements, which they named **polonium** (Po) after Marie's home country, and **radium** (Ra) after its highly radioactive characteristics. For these dicoveries, Marie was awarded a Nobel Prize in Chemistry in 1911, making her one of very few people to receive two Nobel Prizes.

Marie Curie died in 1934 from aplastic anemia, which was almost certainly partly due to her massive exposure to radiation during her lifetime. Most of her work was carried out in a shed without safety measures, and she was known to carry test tubes full of radioactive isotopes in her pockets and to store them in her desk drawers. By the end of her life, not only was she very ill, but her hands had become badly deformed due to their constant exposure to radiation. Unfortunately it was only later in her life that the full dangers of radiation were realised.

7.3 Radioactivity and Types of Radiation

In section 7.1, we discussed that when a nucleus is unstable it can emit particles and energy. This is called **radioactive decay**.

Definition: Radioactive decay
Radioactive decay is the process in which an unstable atomic nucleus loses energy by emitting particles or electromagnetic waves. **Radiation** is the name for the emitted particles or electromagnetic waves.

When a nucleus undergoes radioactive decay, it emits radiation and the nucleus is called radioactive. We are exposed to small amounts of radiation all the time. Even the rocks around us emit radiation! However some elements are far more radioactive than others. *Isotopes* tend to be less stable because they contain a larger number of nucleons than 'non-isotopes' of the same element. These radioactive isotopes are called **radioisotopes**.

Radiation can be emitted in different forms. There are three main types of radiation: alpha, beta and gamma radiation. These are shown in figure 7.1, and are described below.

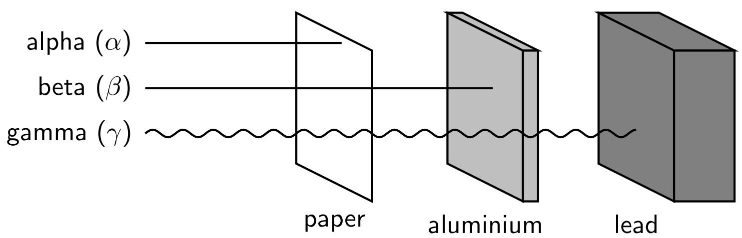


Figure 7.1: Types of radiation

7.3.1 Alpha (α) particles and alpha decay

An alpha particle is made up of two protons and two neutrons bound together. This type of radiation has a *positive charge*. An alpha particle is sometimes represented using the chemical symbol He^{2+}, because it has the same structure as a Helium atom (two neutrons and two protons) which is missing its two electrons, hence the overall charge of $+2$. Alpha particles have very low penetration power. Penetration power describes how easily the particles can pass through another material. Because alpha particles have a *low* penetration power, it means that even something as thin as a piece of paper or the outside layer of the human skin, will absorb these particles so that they can't go any further.

Alpha decay occurs because the nucleus has too many protons, and this causes a lot of repulsion between these like charges. To try to reduce this repulsion, the nucleus emits an α particle. This can be seen in the decay of Americium (Am) to Neptunium (Np).

Example:

$$^{241}_{95}\mathrm{Am} \rightarrow ^{237}_{93}\mathrm{Np} + \alpha\mathrm{particle}$$

Let's take a closer look at what has happened during this reaction. Americium (Z = 95; A = 241) undergoes α decay and releases one alpha particle (i.e. 2 protons and 2 neutrons). The atom now has only 93 protons (Z = 93). On the periodic table, the element which has 93 protons (Z = 93) is called Neptunium. Therefore, the Americium atom has become a Neptunium atom. The atomic mass of the neptunium atom is 237 (A = 237) because 4 nucleons (2 protons and 2 neutrons) were emitted from the atom of Americium.

7.3.2 Beta (β) particles and beta decay

In certain types of radioactive nuclei that have too many neutrons, a neutron may be converted into a proton, an electron and another particle (called a *neutrino*). The high energy electrons that are released in this way are called **beta particles**. Beta particles have a higher penetration power than alpha particles and are able to pass through thicker materials such as paper.

The diagram below shows what happens during β decay:

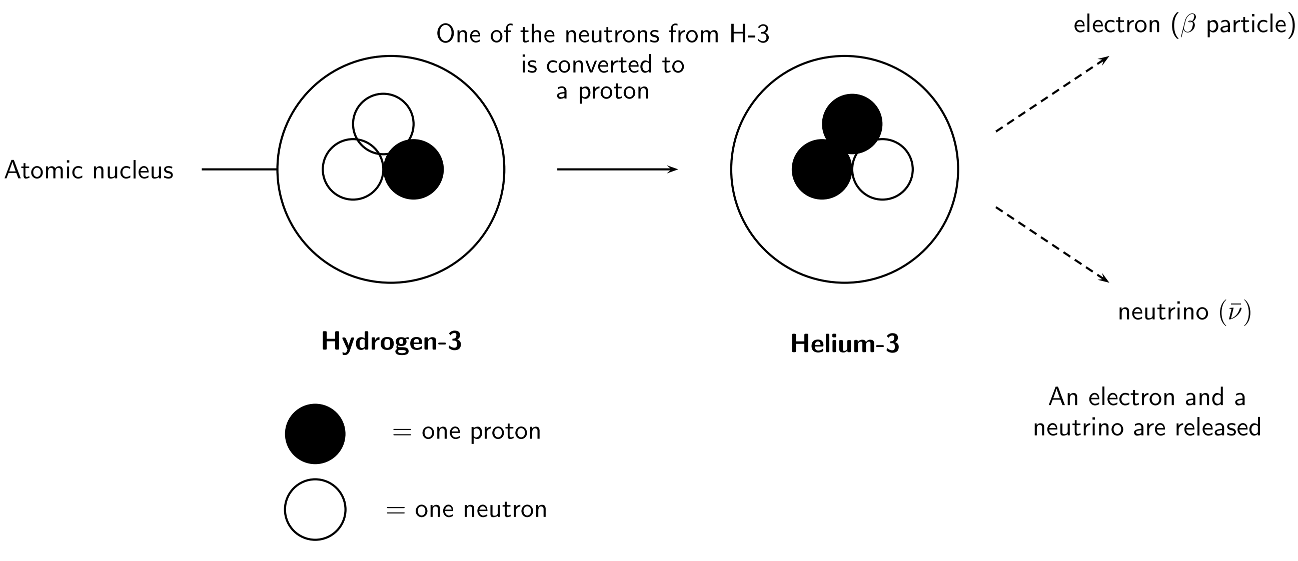


Figure 7.2: β decay in a hydrogen atom

During beta decay, the number of neutrons in the atom decreases by one, and the number of protons increases by one. Since the number of protons before and after the decay is different, the atom has changed into a different element. In figure 7.2, Hydrogen has become Helium. The beta decay of the Hydrogen-3 atom can be represented as follows:

$$^{3}_{1}\mathrm{H} \rightarrow ^{3}_{2}\mathrm{He} + \beta\mathrm{particle} + \bar{\nu}$$

Interesting Fact

When scientists added up all the energy from the neutrons, protons and electrons involved in β-decays, they noticed that there was always some energy missing. We know that energy is always conserved, which led Wolfgang Pauli in 1930 to come up with the idea that another particle, which was not detected yet, also had to be involved in the decay. He called this particle the neutrino (Italian for "little neutral one"), because he knew it had to be neutral, have little or no mass, and interact only very weakly, making it very hard to find experimentally! The neutrino was finally identified experimentally about 25 years after Pauli first thought of it.

Due to the radioactive processes inside the sun, each 1 cm^2 patch of the earth receives 70 billion (70×10^9) neutrinos each second! Luckily neutrinos only interact very weakly so they do not harm our bodies when billions of them pass through us every second.

7.3.3 Gamma (γ) rays and gamma decay

When particles inside the nucleus collide during radioactive decay, energy is released. This energy can leave the nucleus in the form of waves of electromagnetic energy called gamma rays. Gamma radiation is part of the electromagnetic spectrum, just like visible light. However, unlike visible light, humans cannot see gamma rays because they are at a higher frequency and a higher energy. Gamma radiation has no mass or charge. This type of radiation is able to penetrate most common substances, including metals. The only substances that can absorb this radiation are thick lead and concrete.

Gamma decay occurs if the nucleus is at too high an energy level. Since gamma rays are part of the electromagnetic spectrum, they can be thought of as waves *or* particles. Therefore in gamma decay, we can think of a ray or a particle (called a photon) being released. The atomic number and atomic mass remain unchanged.

Table 7.1 summarises and compares the three types of radioactive decay that have been discussed.

Table 7.1: A comparison of alpha, beta and gamma decay

Type of decay	**Particle/ray released**	**Change in element**	**Penetration power**
Alpha (α)	α particle (2 protons and 2 neutrons)	Yes	Low
Beta (β)	β particle (electron)	Yes	Medium
Gamma (γ)	γ ray (electromagnetic energy)	No	High

Worked Example 23: Radioactive decay

Question: The isotope $^{241}_{95}$Pb undergoes radioactive decay and loses two alpha particles.

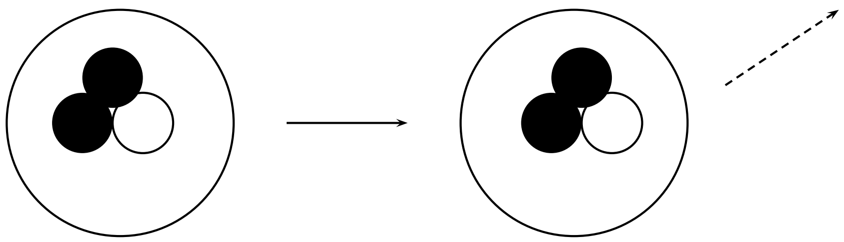


Figure 7.3: γ decay in a helium atom

1. Write the chemical formula of the element that is produced as a result of the decay.
2. Write an equation for this decay process.

Answer

Step 1 : Work out the number of protons and/or neutrons that the radioisotope loses during radioactive decay

One α particle consists of two protons and two neutrons. Since two α particles are released, the total number of protons lost is four and the total number of neutrons lost is also four.

Step 2 : Calculate the atomic number (Z) and atomic mass number (A) of the element that is formed.

$$Z = 95 - 4 = 91$$

$$A = 241 - 4 = 237$$

Step 3 : Refer to the periodic table to see which element has the atomic number that you have calculated.

The element that has Z = 91 is Protactinium (Pa).

Step 4 : Write the symbol for the element that has formed as a result of radioactive decay.

$$^{237}_{91}\text{Pa}$$

Step 5 : Write an equation for the decay process.

$$^{241}_{95}Pb \rightarrow ^{237}_{91} Pa + 2 \text{ protons} + 2 \text{ neutrons}$$

Activity :: Discussion : Radiation

In groups of 3-4, discuss the following questions:

- Which of the three types of radiation is most dangerous to living creatures (including humans!)

- What can happen to people if they are exposed to high levels of radiation?
- What can be done to protect yourself from radiation (Hint: Think of what the radiologist does when you go for an X-ray)?

Exercise: Radiation and radioactive elements

1. There are two main forces inside an atomic nucleus:
   (a) Name these two forces.
   (b) Explain why atoms that contain a greater number of nucleons are more likely to be radioactive.
2. The isotope $^{241}_{95}\text{Pb}$ undergoes radioactive decay and loses three alpha particles.
   (a) Write the chemical formula of the element that is produced as a result of the decay.
   (b) How many nucleons does this element contain?
3. Complete the following equation:
$$^{210}_{82}Am \rightarrow \text{(alpha decay)}$$
4. Radium-228 decays by emitting a beta particle. Write an equation for this decay process.
5. Describe how gamma decay differs from alpha and beta decay.

7.4 Sources of radiation

The sources of radiation can be either **natural** or **man-made**.

7.4.1 Natural background radiation

- *Cosmic radiation*

  The Earth, and all living things on it, are constantly bombarded by radiation from space. Charged particles from the sun and stars interact with the Earth's atmosphere and magnetic field to produce a shower of radiation, mostly beta and gamma radiation. The amount of cosmic radiation varies in different parts of the world because of differences in elevation and also the effects of the Earth's magnetic field.

- *Terrestrial Radiation*

  Radioactive material is found throughout nature. It occurs naturally in the soil, water, and vegetation. The major isotopes that are of concern are uranium and the decay products of uranium, such as thorium, radium, and radon. Low levels of uranium, thorium, and their decay products are found everywhere. Some of these materials are ingested (taken in) with food and water, while others are breathed in. The dose of radiation from terrestrial sources varies in different parts of the world.

Interesting Fact

Cosmic and terrestrial radiation are not the only natural sources. All people have radioactive potassium-40, carbon-14, lead-210 and other isotopes inside their bodies from birth.

7.4.2 Man-made sources of radiation

Although all living things are exposed to natural background radiation, there are other sources of radiation. Some of these will affect most members of the public, while others will only affect those people who are exposed to radiation through their work.

- *Members of the Public*

  Man-made radiation sources that affect members of the public include televisions, tobacco (polonium-210), combustible fuels, smoke detectors (americium), luminous watches (tritium) and building materials. By far, the most significant source of man-made radiation exposure to the public is from medical procedures, such as diagnostic x-rays, nuclear medicine, and radiation therapy. Some of the major isotopes involved are I-131, Tc-99m, Co-60, Ir-192, and Cs-137. The production of nuclear fuel using uranium is also a source of radiation for the public, as is fallout from nuclear weapons testing or use.

- *Individuals who are exposed through their work*

  Any people who work in the following environments are exposed to radiation at some time: radiology (X-ray) departments, nuclear power plants, nuclear medicine departments and radiation oncology (the study of cancer) departments. Some of the isotopes that are of concern are cobalt-60, cesium-137, americium-241, and others.

Radiation therapy (or radiotherapy) uses ionising radiation as part of cancer treatment to control malignant cells. In cancer, a malignant cell is one that divides very rapidly to produce many more cells. These groups of dividing cells can form a growth or **tumour**. The malignant cells in the tumour can take nutrition away from other healthy body cells, causing them to die, or can increase the pressure in parts of the body because of the space that they take up. Radiation therapy uses radiation to try to target these malignant cells and kill them. However, the radiation can also damage other, healthy cells in the body. To stop this from happening, shaped radiation beams are aimed from several angles to intersect at the tumour, so that the radiation dose here is much higher than in the surrounding, healthy tissue. But even doing this doesn't protect all the healthy cells, and that is why people have side-effects to this treatment.

Note that radiation therapy is different from chemotherapy, which uses *chemicals*, rather than radiation, to destroy malignant cells. Generally, the side effects of chemotherapy are greater because the treatment is not as localised as it is with radiation therapy. The chemicals travel throughout the body, affecting many healthy cells.

7.5 The 'half-life' of an element

Definition: Half-life
The half-life of an element is the time it takes for half the atoms of a radioisotope to decay into other atoms.

Table 7.2 gives some examples of the half-lives of different elements.

Table 7.2: Table showing the half-life of a number of elements

Radioisotope	Chemical symbol	Half-life
Polonium-212	Po-212	0.16 seconds
Sodium-24	Na-24	15 hours
Strontium-90	Sr-90	28 days
Cobalt-60	Co-60	5.3 years
Caesium-137	Cs-137	30 years
Carbon-14	C-14	5 760 years
Calcium	Ca	100 000 years
Beryllium	Be	2 700 000 years
Uranium-235	U-235	7.1 billion years

So, in the case of Sr-90, it will take 28 days for half of the atoms to decay into other atoms. It will take another 28 days for half of the remaining atoms to decay. Let's assume that we have a sample of strontium that weighs 8g. After the first 28 days there will be:

$$1/2 \times 8 = 4 \text{ g Sr-90 left}$$

After 56 days, there will be:

$$1/2 \times 4 \text{ g} = 2 \text{ g Sr-90 left}$$

After 84 days, there will be:

$$1/2 \times 2 \text{ g} = 1 \text{ g Sr-90 left}$$

If we convert these amounts to a *fraction* of the original sample, then after 28 days 1/2 of the sample remains undecayed. After 56 days 1/4 is undecayed and after 84 days, 1/8 and so on.

Activity :: Group work : Understanding half-life

Work in groups of 4-5

You will need:

16 sheets of A4 paper per group, scissors, 2 boxes per group, a marking pen and timer/stopwatch.

What to do:

- Your group should have two boxes. Label one 'decayed' and the other 'radioactive'.
- Take the A4 pages and cut each into 4 pieces of the same size. You should now have 64 pieces of paper. Stack these neatly and place them in the 'radioactive' box. The paper is going to represent some radioactive material.
- Set the timer for one minute. After one minute, remove half the sheets of paper from the radioactive box and put them in the 'decayed' box.
- Set the timer for another minute and repeat the previous step, again removing half the pieces of paper that are left in the radioactive box and putting them in the decayed box.
- Repeat this process until 8 minutes have passed. You may need to start cutting your pieces of paper into even smaller pieces as you progress.

Questions:

1. How many pages were left in the radioactive box after...
   (a) 1 minute
   (b) 3 minutes

   (c) 5 minutes
2. What percentage (%) of the pages were left in the radioactive box after...
   (a) 2 minutes
   (b) 4 minutes
3. After how many minutes is there 1/128 of radioactive material remaining?
4. What is the half-life of the 'radioactive' material in this exercise?

Worked Example 24: Half-life 1

Question: A 100 g sample of Cs-137 is allowed to decay. Calculate the mass of Cs-137 that will be left after 90 years

Answer
Step 1 : You need to know the half-life of Cs-137
The half-life of Cs-137 is 30 years.

Step 2 : Determine how many times the quantity of sample will be halved in 90 years.
If the half-life of Cs-137 is 30 years, and the sample is left to decay for 90 years, then the number of times the quantity of sample will be halved is 90/30 = 3.

Step 3 : Calculate the quantity that will be left by halving the mass of Cs-137 three times
1. After 30 years, the mass left is 100 g $\times$ 1/2 = 50 g
2. After 60 years, the mass left is 50 g $\times$ 1/2 = 25 g
3. After 90 years, the mass left is 25 g $\times$ 1/2 = 12.5 g

Note that a quicker way to do this calculation is as follows:
Mass left after 90 years = $(1/2)^3 \times 100$ g = 12.5 g (The exponent is the number of times the quantity is halved)

Worked Example 25: Half-life 2

Question: An 80 g sample of Po-212 decays until only 10 g is left. How long did it take for this decay to take place?

Answer
Step 1 : Calculate the fraction of the original sample that is left after decay
Fraction remaining = 10 g/80 g = 1/8

Step 2 : Calculate how many half-life periods of decay (x) must have taken place for 1/8 of the original sample to be left

$$(\frac{1}{2})^x = \frac{1}{8}$$

Therefore, x = 3

Step 3 : Use the half-life of Po-212 to calculate how long the sample was left to decay

The half-life of Po-212 is 0.16 seconds. Therefore if there were three periods of decay, then the total time is 0.16×3. The time that the sample was left to decay is 0.48 seconds.

?

Exercise: Looking at half life

1. Imagine that you have 100 g of Na-24.
   (a) What is the half life of Na-24?
   (b) How much of this isotope will be left after 45 hours?
   (c) What percentage of the original sample will be left after 60 hours?
2. A sample of Sr-90 is allowed to decay. After 84 days, 10 g of the sample remains.
   (a) What is the half life of Sr-90?
   (b) How much Sr-90 was in the original sample?
   (c) How much Sr-90 will be left after 112 days?

7.6 The Dangers of Radiation

Natural radiation comes from a variety of sources such as the rocks, sun and from space. However, when we are exposed to large amounts of radiation, this can cause damage to cells. γ radiation is particularly dangerous because it is able to penetrate the body, unlike α and β particles whose penetration power is less. Some of the dangers of radiation are listed below:

- **Damage to cells**

  Radiation is able to penetrate the body, and also to penetrate the membranes of the cells within our bodies, causing massive damage. *Radiation poisoning* occurs when a person is exposed to large amounts of this type of radiation. Radiation poisoning damages tissues within the body, causing symptoms such as diarrhoea, vomiting, loss of hair and convulsions.

- **Genetic abnormalities**

  When radiation penetrates cell membranes, it can damage chromosomes within the nucleus of the cell. The chromosomes contain all the genetic information for that person. If the chromosomes are changed, this may lead to genetic abnormalities in any children that are born to the person who has been exposed to radiation. Long after the nuclear disaster of Chernobyl in Russia in 1986, babies were born with defects such as missing limbs and abnormal growths.

- **Cancer**

  Small amounts of radiation can cause cancers such as leukemia (cancer of the blood)

7.7 The Uses of Radiation

However, despite the many dangers of radiation, it does have many powerful uses, some of which are listed below:

- **Medical Field**

  Radioactive *chemical tracers* emitting γ rays can give information about a person's internal anatomy and the functioning of specific organs. The radioactive material may be injected into the patient, from where it will target specific areas such as bones or tumours. As the material decays and releases radiation, this can be seen using a special type of camera or other instrument. The radioactive material that is used for this purpose must have a short half-life so that the radiation can be detected quickly and also so that the material is quickly removed from the patient's body. Using radioactive materials for this purpose can mean that a tumour or cancer may be diagnosed long before these would have been detected using other methods such as X-rays.

  Radiation may also be used to sterilise medical equipment.

Activity :: Research Project : The medical uses of radioisotopes

Carry out your own research to find out more about the radioisotopes that are used to diagnose diseases in the following parts of the body:

- thyroid gland
- kidneys
- brain

In each case, try to find out...

1. which radioisotope is used
2. what the sources of this radioisotope are
3. how the radioisotope enters the patient's body and how it is monitored

---

- **Biochemistry and Genetics**

  Radioisotopes may be used as tracers to label molecules so that chemical processes such as DNA replication or amino acid transport can be traced.

- **Food preservation**

  Irradiation of food can stop vegetables or plants from sprouting after they have been harvested. It also kills bacteria and parasites, and controls the ripening of fruits.

- **Environment**

  Radioisotopes can be used to trace and analyse pollutants.

- **Archaeology and Carbon dating**

  Natural radioisotopes such as C-14 can be used to determine the age of organic remains. All living organisms (e.g. trees, humans) contain carbon. Carbon is taken in by plants and trees through the process of photosynthesis in the form of carbon dioxide and is then converted into organic molecules. When animals feed on plants, they also obtain carbon through these organic compounds. Some of the carbon in carbon dioxide is the radioactive C-14, while the rest is a non-radioactive form of carbon. When an organism dies, no more carbon is taken in and there is a limited amount of C-14 in the body. From this point onwards, C-14 begins its radioactive decay. When scientists uncover remains, they are able to estimate the age of the remains by seeing how much C-14 is left in the body relative to the amount of non-radioactive carbon. The less C-14 there is, the older the remains because radioactive decay must have been taking place for a long time. Because scientists know the exact rate of decay of C-14, they can calculate a very accurate estimate of the age of the remains. Carbon dating has been a very important tool in building up accurate historical records.

Activity :: Case Study : Using radiocarbon dating

Radiocarbon dating has played an important role in uncovering many aspects of South Africa's history. Read the following extract from an article that appeared in Afrol news on 10th February 2007 and then answer the questions that follow.

> The world famous rock art in South Africa's uKhahlamba-Drakensberg, a World Heritage Site, is three times older than previously thought, archaeologists conclude in a new study. The more than 40,000 paintings were made by the San people some 3000 years ago, a new analysis had shown.
> Previous work on the age of the rock art in uKhahlamba-Drakensberg concluded it is less than 1,000 years old. But the new study - headed by a South African archaeologist leading a team from the University of Newcastle upon Tyne (UK) and Australian National University in Canberra - estimates the panels were created up to 3,000 years ago. They used the latest radio-carbon dating technology.
> The findings, published in the current edition of the academic journal 'South African Humanities', have "major implications for our understanding of how the rock artists lived and the social changes that were taking place over the last three millennia," according to a press release from the British university.

Questions:

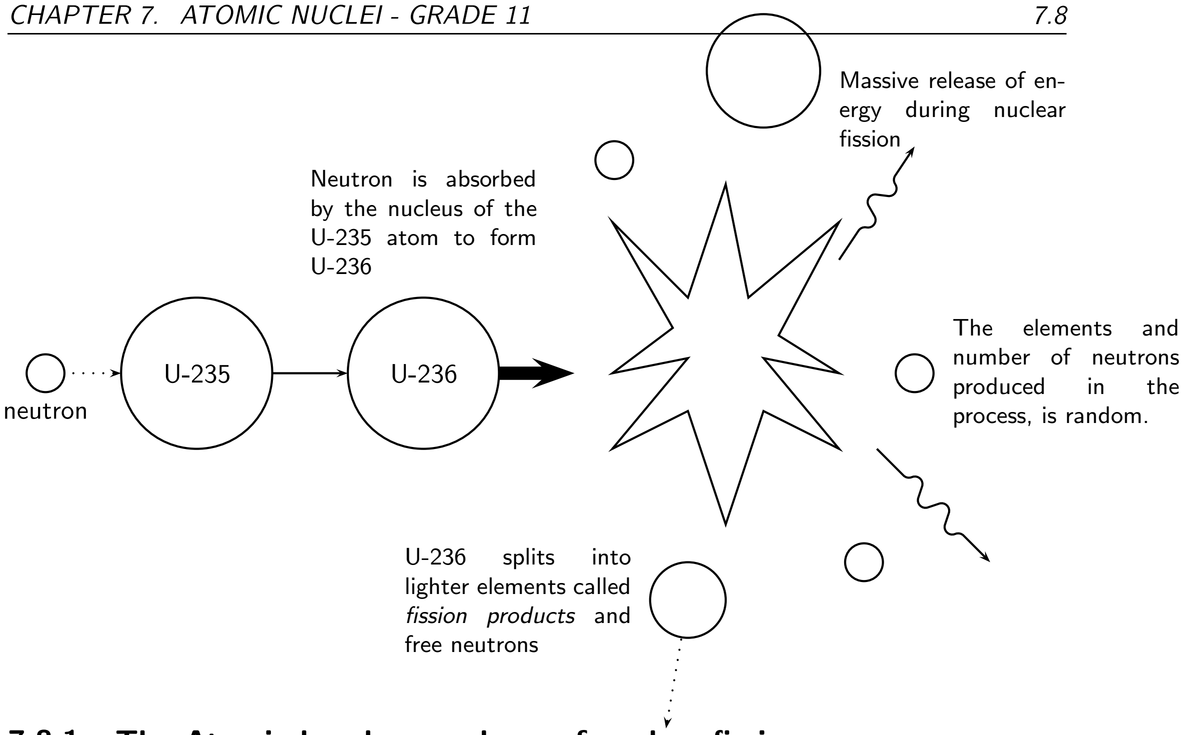
1. What is the half-life of carbon-14?
2. In the news article, what role did radiocarbon dating play in increasing our knowledge of South Africa's history?
3. Radiocarbon dating can also be used to analyse the remains of once-living organisms. Imagine that a set of bones are found between layers of sediment and rock in a remote area. A group of archaeologists carries out a series of tests to try to estimate the age of the bones. They calculate that the bones are approximately 23 040 years old.
   What percentage of the original carbon-14 must have been left in the bones for them to arrive at this estimate?

7.8 Nuclear Fission

Nuclear fission is a process where the nucleus of an atom is split into two or more smaller nuclei, known as *fission products*. The fission of heavy elements is an **exothermic reaction** and huge amounts of energy are released in the process. This energy can be used to produce *nuclear power* or to make *nuclear weapons*, both of which we will discuss a little later.

Definition: Nuclear fission
The splitting of an atomic nucleus

Below is a diagram showing the nuclear fission of Uranium-235. An atom of Uranium-235 is bombarded with a neutron to initiate the fission process. This neutron is absorbed by Uranium-235, to become Uranium-236. Uranium-236 is highly unstable and breaks down into a number of lighter elements, releasing energy in the process. Free neutrons are also produced during this process, and these are then available to bombard other fissionable elements. This process is known as a **fission chain reaction**, and occurs when one nuclear reaction starts off another, which then also starts off another one so that there is a rapid increase in the number of nuclear reactions that are taking place.



7.8.1 The Atomic bomb - an abuse of nuclear fission

A nuclear chain reaction can happen very quickly, releasing vast amounts of energy in the process. In 1939, it was discovered that Uranium could undergo nuclear fission. In fact, it was uranium that was used in the first atomic bomb. The bomb contained huge amounts of Uranium-235, enough to start a runaway nuclear fission chain reaction. Because the process was uncontrolled, the energy from the fission reactions was released in a matter of *seconds*, resulting in the massive explosion of that first bomb. Since then, more atomic bombs have been dropped, causing massive destruction and loss of life.

Activity :: Discussion : Nuclear weapons testing - an ongoing issue

Read the article below which has been adapted from one that appeared in 'The Globe' in Washington on 10th October 2006, and then answer the questions that follow.

> US officials and arms control specialists warned yesterday that North Korea's test of a small nuclear device could start an arms race in the region and threaten the landmark global treaty designed nearly four decades ago to halt the spread of nuclear weapons. US officials expressed concern that North Korea's neighbors, including Japan, Taiwan, and South Korea, could eventually decide to develop weapons of their own. They also fear that North Korea's moves could embolden Iran, and that this in turn could encourage Saudi Arabia or other neighbours in the volatile Middle East to one day seek nuclear deterrents, analysts say.
>
> North Korea is the first country to conduct a nuclear test after pulling out of the Nuclear Nonproliferation Treaty. The treaty, which was created in 1968, now includes 185 nations (nearly every country in the world). Under the treaty, the five declared nuclear powers at the time (United States, the Soviet Union, France, China, and Great Britain) agreed to reduce their supplies of nuclear weapons. The treaty has also helped to limit the number of new nuclear weapons nations.
>
> But there have also been serious setbacks. India and Pakistan, which never signed the treaty, became new nuclear powers, shocking the world with test explosions in 1998. The current issue of nuclear weapons testing in North Korea, is another such setback and a blow to the treaty.

Group discussion questions:

1. Discuss what is meant by an 'arms race' and a 'treaty'.

2. Do you think it is important to have such treaties in place to control the testing and use of nuclear weapons? Explain your answer.
3. Discuss some of the reasons why countries might not agree to be part of a nuclear weapons treaty.
4. How would you feel if South Africa decided to develop its own nuclear weapons?

7.8.2 Nuclear power - harnessing energy

However, nuclear fission can also be carried out in a controlled way in a *nuclear reactor*. A nuclear reactor is a piece of equipment where nuclear chain reactions can be started in a controlled and sustained way. This is different from a nuclear *explosion* where the chain reaction occurs in seconds. The most important use of nuclear reactors at the moment is to produce **electrical power**, and most of these nuclear reactors use nuclear fission. A **nuclear fuel** is a chemical isotope that can keep a fission chain reaction going. The most common isotopes that are used are Uranium-235 and Plutonium-239. The amount of free energy that is in nuclear fuels is far greater than the energy in a similar amount of other fuels such as gasoline. In many countries, nuclear power is seen as a relatively environmentally friendly alternative to fossil fuels, which release large amounts of greenhouse gases, and are also non-renewable resources. However, one of the concerns around the use of nuclear power, is the production of *nuclear waste* which contains radioactive chemical elements.

Activity :: Debate : Nuclear Power

The use of nuclear power as a source of energy has been a subject of much debate. There are many advantages of nuclear power over other energy sources. These include the large amount of energy that can be produced at a small plant, little atmospheric pollution and the small quantity of waste. However there are also disadvantages. These include the expense of maintaining nuclear power stations, the huge impact that an accident could have as well as the disposal of dangerous nuclear waste.

Use these ideas as a starting point for a class debate.

Nuclear power - An energy alternative or environmental hazard?

Your teacher will divide the class into teams. Some of the teams will be 'pro' nuclear power while the others will be 'anti' nuclear power.

7.9 Nuclear Fusion

Nuclear fusion is the joining together of the nuclei of two atoms to form a heavier nucleus. If the atoms involved are small, this process is accompanied by the release of energy. It is the nuclear fusion of elements that causes stars to shine and hydrogen bombs to explode. As with nuclear *fission* then, there are both positive and negative uses of nuclear fusion.

Definition: Nuclear fusion
The joining together of the nuclei of two atoms.

You will remember that nuclei naturally repel one another because of the electrostatic force between their positively charged protons. So, in order to bring two nuclei together, a lot of energy must be supplied if fusion is to take place. If two nuclei can be brought close enough together however, the electrostatic force is overwhelmed by the more powerful strong nuclear force which only operates over short distances. If this happens, nuclear fusion can take place. Inside the cores of stars, the temperature is high enough for hydrogen fusion to take place but scientists have so far been unsuccessful in making it work in the laboratory. One of the huge advantages of nuclear fusion, if it could be made to happen, is that it is a relatively environmentally friendly source of energy. The helium that is produced is not radioactive or poisonous and does not carry the dangers of nuclear fission.

7.10 Nucleosynthesis

An astronomer named Edwin Hubble discovered in the 1920's that the universe is expanding. He measured that far-away galaxies are moving away from the earth at great speed, and the further away they are, the faster they are moving.

Extension: What are galaxies?

Galaxies are huge clusters of stars and matter in the universe. The earth is part of the Milky Way galaxy which is shaped like a very large spiral. Astronomers can measure the light coming from distant galaxies using telescopes. Edwin Hubble was also able to measure the velocities of galaxies.

These observations led people to see that the universe is expanding. It also led to the *Big Bang* hypothesis. The 'Big Bang' hypothesis is an idea about how the universe may have started. According to this theory, the universe started off at the beginning of time as a point which then exploded and expanded into the universe we live in today. This happened between 10 and 14 billion years ago.

Just after the Big Bang, when the universe was only 10^{-43}s old, it was very hot and was made up of quarks and leptons (an example of a lepton is the electron). As the universe expanded, ($\sim 10^{-2}$s) and cooled, the quarks started binding together to form protons and neutrons (together called *nucleons*).

7.10.1 Age of Nucleosynthesis (225 s - 10^3 s)

About 225 s after the Big Bang, the protons and neutrons started binding together to form simple *nuclei*. The process of forming nuclei is called *nucleosynthesis*. When a proton and a neutron bind together, they form the *deuteron*. The deuteron is like a hydrogen nucleus (which is just a proton) with a neutron added to it so it can be written as $^2\mathrm{H}$. Using protons and neutrons as building blocks, more nuclei can be formed as shown below. For example, the Helium-4 nucleus (also called an *alpha particle*) can be formed in the following ways:

$$\begin{aligned} {}^2\mathrm{H} + \mathrm{n} &\rightarrow {}^3\mathrm{H} \\ \mathrm{deuteron} + \mathrm{neutron} &\rightarrow \mathrm{triton} \end{aligned}$$

then:

$$\begin{aligned} {}^3\mathrm{H} + \mathrm{p} &\rightarrow {}^4\mathrm{He} \\ \mathrm{triton} + \mathrm{proton} &\rightarrow \mathrm{Helium4\ (alpha\ particle)} \end{aligned}$$

or

$$\begin{aligned} {}^{2}\mathrm{H} + \mathrm{p} &\rightarrow {}^{3}\mathrm{He} \\ \mathrm{deuteron} + \mathrm{proton} &\rightarrow \mathrm{Helium3} \end{aligned}$$

then:

$$\begin{aligned} {}^{3}\mathrm{He} + \mathrm{n} &\rightarrow {}^{4}\mathrm{He} \\ \mathrm{Helium3} + \mathrm{neutron} &\rightarrow \mathrm{Helium4\ (alpha\ particle)} \end{aligned}$$

Some $^{7}\mathrm{Li}$ nuclei could also have been formed by the fusion of $^{4}\mathrm{He}$ and $^{3}\mathrm{H}$.

7.10.2 Age of Ions (10^{3} s - 10^{13} s)

However, at this time the universe was still very hot and the electrons still had too much energy to become bound to the alpha particles to form helium *atoms*. Also, the nuclei with mass numbers greater than 4 (i.e. greater than $^{4}\mathrm{He}$) are very short-lived and would have decayed almost immediately after being formed. Therefore, the universe moved through a stage called the Age of Ions when it consisted of free positively charged H^{+} ions and $^{4}\mathrm{He}$ ions, and negatively charged electrons not yet bound into atoms.

7.10.3 Age of Atoms (10^{13} s - 10^{15} s)

As the universe expanded further, it cooled down until the electrons were able to bind to the hydrogen and helium nuclei to form hydrogen and helium atoms. Earlier, during the Age of Ions, both the hydrogen and helium ions were positively charged which meant that they repelled each other (electrostatically). During the Age of Atoms, the hydrogen and helium along with the electrons, were in the form of atoms which are electrically neutral and so they no longer repelled each other and instead pulled together under gravity to form clouds of gas, which evetually formed stars.

7.10.4 Age of Stars and Galaxies (the universe today)

Inside the core of stars, the densities and temperatures are high enough for fusion reactions to occur. Most of the heavier nuclei that exist today were formed inside stars from thermonuclear reactions! (It's interesting to think that the atoms that we are made of were actually manufactured inside stars!). Since stars are mostly composed of hydrogen, the first stage of thermonuclear reactions inside stars involves hydrogen and is called **hydrogen burning**. The process has three steps and results in four hydrogen atoms being formed into a helium atom with (among other things) two photons (light!) being released.

The next stage is **helium burning** which results in the formation of carbon. All these reactions release a large amount of energy and heat the star which causes heavier and heavier nuclei to fuse into nuclei with higher and higher atomic numbers. The process stops with the formation of $^{56}\mathrm{Fe}$, which is the most strongly bound nucleus. To make heavier nuclei, even higher energies are needed than is possible inside normal stars. These nuclei are most likely formed when huge amounts of energy are released, for example when stars explode (an exploding star is called a **supernova**). This is also how all the nuclei formed inside stars get "recycled" in the universe to become part of new stars and planets.

7.11 Summary

- **Nuclear physics** is the branch of physics that deals with the nucleus of an atom.

- There are two forces between the particles of the nucleus. The **strong nuclear force** is an attractive force between the neutrons and the **electromagnetic force** is the repulsive force between like-charged protons.
- In atoms with large nuclei, the electromagnetic force becomes greater than the strong nuclear force and particles or energy may be released from the nucleus.
- **Radioactive decay** occurs when an unstable atomic nucleus loses energy by emitting particles or electromagnetic waves.
- The particles and energy released are called **radiation** and the atom is said to be **radioactive**.
- Radioactive isotopes are called **radioisotopes**.
- Radioactivity was first discovered by Marie Curie and her husband Pierre.
- There are three types of radiation from radioactive decay: **alpha (α)**, **beta (β)** and **gamma (γ)** radiation.
- During **alpha decay**, an alpha particle is released. An alpha particle consists of two protons and two neutrons bound together. Alpha radiation has low penetration power.
- During **beta decay**, a beta particle is released. During beta decay, a neutron is converted to a proton, an electron and a neutrino. A beta particle is the electron that is released. Beta radiation has greater penetration power than alpha radiation.
- During **gamma decay**, electromagnetic energy is released as gamma rays. Gamma radiation has the highest penetration power of the three radiation types.
- There are many **sources of radiation**. Some of natural and others are man-made.
- Natural sources of radiation include cosmic and terrestrial radiation.
- Man-made sources of radiation include televisions, smoke detectors, X-rays and radiation therapy.
- The **half-life** of an element is the time it takes for half the atoms of a radioisotope to decay into other atoms.
- Radiation can be very damaging. Some of the **negative impacts of radiation** exposure include damage to cells, genetic abnormalities and cancer.
- However, radiation can also have many **positive uses**. These include use in the medical field (e.g. chemical tracers), biochemistry and genetics, use in food preservation, the environment and in archaeology.
- **Nuclear fission** is the splitting of an atomic nucleus into smaller fission products. Nuclear fission produces large amounts of energy, which can be used to produce nuclear power, and to make nuclear weapons.
- **Nuclear fusion** is the joining together of the nuclei of two atoms to form a heavier nucleus. In stars, fusion reactions involve the joining of hydrogen atoms to form helium atoms.
- **Nucleosynthesis** is the process of forming nuclei. This was very important in helping to form the universe as we know it.

?

Exercise: Summary exercise

1. Explain each of the following terms:
   (a) electromagnetic force
   (b) radioactive decay
   (c) radiocarbon dating

2. For each of the following questions, choose the **one correct answer**:
   (a) The part of the atom that undergoes radioactive decay is the...
      i. neutrons
      ii. nucleus
      iii. electrons
      iv. entire atom
   (b) The radio-isotope Po-212 undergoes alpha decay. Which of the following statements is **true**?
      i. The number of protons in the element remains unchanged.
      ii. The number of nucleons after decay is 212.
      iii. The number of protons in the element after decay is 82.
      iv. The end product after decay is Po-208.
3. 20 g of sodium-24 undergoes radoactive decay. Calculate the percentage of the original sample that remains after 60 hours.
4. Nuclear physics can be controversial. Many people argue that studying the nucleus has led to devastation and huge loss of life. Others would argue that the benefits of nuclear physics far outweigh the negative things that have come from it.
   (a) Outline some of the ways in which nuclear physics has been used in negative ways.
   (b) Outline some of the benefits that have come from nuclear physics.

Chapter 8

Thermal Properties and Ideal Gases - Grade 11

We are surrounded by gases in our atmosphere which support and protect life on this planet. In this chapter, we are going to try to understand more about gases, and learn how to predict how they will behave under different conditions. The kinetic theory of matter was discussed in chapter 2. This theory is very important in understanding how gases behave.

8.1 A review of the kinetic theory of matter

The main assumptions of the kinetic theory of matter are as follows:

- Matter is made up of **particles** (e.g. atoms or molecules)
- These particles are constantly moving because they have kinetic energy. The space in which the particles move is the **volume** of the gas.
- There are **spaces** between the particles
- There are **attractive forces** between particles and these become stronger as the particles move closer together.
- All particles have **energy**. The **temperature** of a substance is a measure of the **average kinetic energy** of the particles.
- A change in **phase** may occur when the energy of the particles is changed.

The kinetic theory applies to all matter, including gases. In a gas, the particles are far apart and have a high kinetic energy. They move around freely, colliding with each other or with the sides of the container if the gas is enclosed. The **pressure** of a gas is a measure of the frequency of collisions of the gas particles with each other and with the sides of the container that they are in. If the gas is heated, the average kinetic energy of the gas particles will increase and if the temperature is decreased, so does their energy. If the energy of the particles decreases significantly, the gas liquifies. An **ideal gas** is one that obeys all the assumptions of the kinetic theory of matter. A **real gas** behaves like an ideal gas, except at high pressures and low temperatures. This will be discussed in more detail later in this chapter.

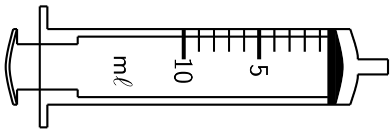
Definition: Ideal gas
An ideal gas or perfect gas is a hypothetical gas that obeys all the assumptions of the kinetic theory of matter. In other words, an ideal gas would have identical particles of zero volume, with no intermolecular forces between them. The atoms or molecules in an ideal gas would also undergo elastic collisions with the walls of their container.

Definition: Real gas
Real gases behave more or less like ideal gases except under certain conditions e.g. high pressures and low temperatures.

There are a number of laws that describe how gases behave. It will be easy to make sense of these laws if you understand the kinetic theory of gases that was discussed above.

8.2 Boyle's Law: Pressure and volume of an enclosed gas

Activity :: Demonstration : Boyle's Law
If you have ever tried to force in the plunger of a syringe or a bicycle pump while sealing the opening with a finger, you will have seen Boyle's Law in action! This will now be demonstrated using a 10 ml syringe.
Aim:
To demonstrate Boyle's law.
Apparatus:
You will only need a syringe for this demonstration.



Method:

1. Hold the syringe in one hand, and with the other pull the plunger out towards you so that the syringe is now full of air.
2. Seal the opening of the syringe with your finger so that no air can escape the syringe.
3. Slowly push the plunger in, and notice whether it becomes *more* or *less* difficult to push the plunger in.

Results:
What did you notice when you pushed the plunger in? What happens to the **volume** of air inside the syringe? Did it become *more* or *less* difficult to push the plunger in as the volume of the air in the syringe decreased? In other words, did you have to apply more or less **pressure** to the plunger as the volume of air in the syringe decreased?
As the volume of air in the syringe decreases, you have to apply more pressure to the plunger to keep forcing it down. The pressure of the gas inside the syringe pushing back on the plunger is greater. Another way of saying this is that as the volume of the gas in the syringe *decreases*, the pressure of that gas *increases*.
Conclusion:
If the volume of the gas decreases, the pressure of the gas increases. If the volume of the gas increases, the pressure decreases. These results support Boyle's law.

In the previous demonstration, the volume of the gas decreased when the pressure increased, and the volume increased when the pressure decreased. This is called an **inverse relationship**. The inverse relationship between pressure and volume is shown in figure 8.1.

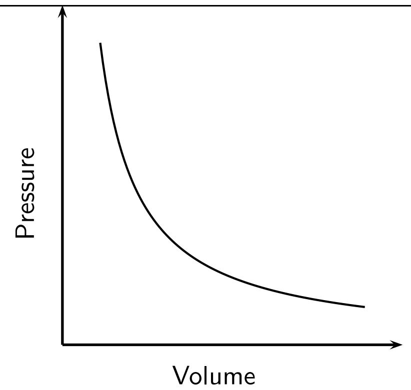


Figure 8.1: Graph showing the inverse relationship between pressure and volume

Can you use the kinetic theory of gases to explain this inverse relationship between the pressure and volume of a gas? Let's think about it. If you decrease the volume of a gas, this means that the same number of gas particles are now going to come into contact with each other and with the sides of the container much more often. You may remember from earlier that we said that *pressure* is a measure of the *frequency of collisions* of gas particles with each other and with the sides of the container they are in. So, if the volume decreases, the pressure will naturally increase. The opposite is true if the volume of the gas is increased. Now, the gas particles collide less frequently and the pressure will decrease.

It was an Englishman named Robert Boyle who was able to take very accurate measurements of gas pressures and volumes using excellent vacuum pumps. He discovered the startlingly simple fact that the pressure and volume of a gas are not just vaguely inversely related, but are *exactly* **inversely proportional**. This can be seen when a graph of pressure against the inverse of volume is plotted. When the values are plotted, the graph is a straight line. This relationship is shown in figure 8.2.

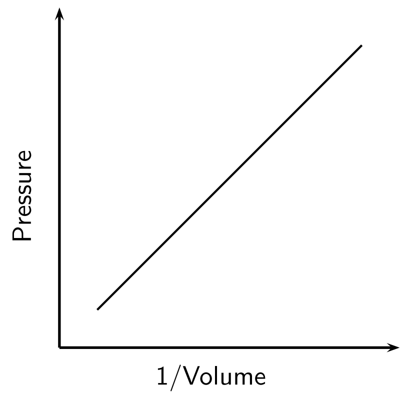


Figure 8.2: The graph of pressure plotted against the inverse of volume, produces a straight line. This shows that pressure and volume are exactly inversely proportional.

Definition: Boyle's Law
The pressure of a fixed quantity of gas is inversely proportional to the volume it occupies so long as the temperature remains constant.

Important: Proportionality

During this chapter, the terms **directly proportional** and **inversely proportional** will be used a lot, and it is important that you understand their meaning. Two quantities are said to be **proportional** if they vary in such a way that one of the quantities is a constant multiple of the other, or if they have a constant ratio. We will look at two examples to show the difference between *directly proportional* and *inversely proportional*.

1. *Directly proportional*
   A car travels at a constant speed of 120 km/h. The time and the distance covered are shown in the table below.

Time (mins)	**Distance** (km)
10	20
20	40
30	60
40	80

   What you will notice is that the two quantities shown are constant multiples of each other. If you divide each distance value by the time the car has been driving, you will always get 2. This shows that the values are proportional to each other. They are **directly proportional** because both values are increasing. In other words, as the driving time increases, so does the distance covered. The same is true if the values decrease. The shorter the driving time, the smaller the distance covered. This relationship can be described mathematically as:

   $$y = kx$$

   where y is distance, x is time and k is the *proportionality constant*, which in this case is 2. Note that this is the equation for a straight line graph! The symbol $\propto$ is also used to show a directly proportional relationship.

2. *Inversely proportional*
   Two variables are inversely proportional if one of the variables is directly proportional to the multiplicative inverse of the other. In other words,

   $$y \propto \frac{1}{x}$$

   or

   $$y = \frac{k}{x}$$

   This means that as one value gets bigger, the other value will get smaller. For example, the time taken for a journey is inversely proportional to the speed of travel. Look at the table below to check this for yourself. For this example, assume that the distance of the journey is 100 km.

Speed (km/h)	**Time** (mins)
100	60
80	75
60	100
40	150

   According to our definition, the two variables are inversely proportional is one variable is *directly* proportional to the *inverse* of the other. In other words, if we divide one of the variables by the inverse of the other, we should always get the same number. For example,

   $$\frac{100}{1/60} = 6000$$

   If you repeat this using the other values, you will find that the answer is always 6000. The variables are inversely proportional to each other.

We know now that the pressure of a gas is *inversely proportional* to the volume of the gas, provided the temperature stays the same. We can write this relationship symbolically as

$$p \propto \frac{1}{V}$$

This equation can also be written as follows:

$$p = \frac{k}{V}$$

where k is a proportionality constant. If we rearrange this equation, we can say that:

$$pV = k$$

This equation means that, assuming the temperature is constant, multiplying any pressure and volume values for a fixed amount of gas will always give the same value. So, for example, $p_1V_1 = k$ and $p_2V_2 = k$, where the subscripts 1 and 2 refer to two pairs of pressure and volume readings for the same mass of gas at the same temperature.

From this, we can then say that:

$$p_1V_1 = p_2V_2$$

In the gas equations, k is a "variable constant". This means that k is constant in a particular set of situations, but in two different sets of situations it has different constant values.

Important: Remember that Boyle's Law requires two conditions. First, the amount of gas must stay constant. Clearly, if you let a little of the air escape from the container in which it is enclosed, the pressure of the gas will decrease along with the volume, and the inverse proportion relationship is broken. Second, the temperature must stay constant. Cooling or heating matter generally causes it to contract or expand. In our original syringe demonstration, if you were to heat up the gas in the syringe, it would expand and force you to apply a greater force to keep the plunger at a given position. Again, the proportionality would be broken.

Activity :: Investigation : Boyle's Law

Here are some of Boyle's original data. Note that pressure would originally have been measured using a *mercury manometer* and the units for pressure would have been *millimetres mercury* or mm Hg. However, to make things a bit easier for you, the pressure data have been converted to a unit that is more familiar. Note that the volume is given in terms of arbitrary marks (evenly made).

Volume (graduation mark)	**Pressure** (kPa)	**Volume** (graduation mark)	**Pressure** (kPa)
12	398	28	170
14	340	30	159
16	298	32	150
18	264	34	141
20	239	36	133
22	217	38	125
24	199	40	120
26	184		

1. Plot a graph of pressure (p) against volume (V). Volume will be on the x-axis and pressure on the y-axis. Describe the relationship that you see.

2. Plot a graph of p against $1/V$. Describe the relationship that you see.
3. Do your results support Boyle's Law? Explain your answer.

Did you know that the mechanisms involved in *breathing* also relate to Boyle's Law? Just below the lungs is a muscle called the **diaphragm**. When a person breathes in, the diaphragm moves down and becomes more 'flattened' so that the volume of the lungs can increase. When the lung volume *increases*, the pressure in the lungs *decreases* (Boyle's law). Since air always moves from areas of high pressure to areas of lower pressure, air will now be drawn into the lungs because the air pressure *outside* the body is higher than the pressure *in* the lungs. The opposite process happens when a person breathes out. Now, the diaphragm moves upwards and causes the volume of the lungs to *decrease*. The pressure in the lungs will *increase*, and the air that was in the lungs will be forced out towards the lower air pressure outside the body.

Worked Example 26: Boyle's Law 1

Question: A sample of helium occupies a volume of 160 cm^3 at 100 kPa and 25 °C. What volume will it occupy if the pressure is adjusted to 80 kPa and if the temperature remains unchanged?

Answer

Step 4 : Write down all the information that you know about the gas.

V_1 = 160 cm^3 and V_2 = ?
p_1 = 100 kPa and p_2 = 80 kPa

Step 1 : Use an appropriate gas law equation to calculate the unknown variable.

Because the temperature of the gas stays the same, the following equation can be used:

$$p_1V_1 = p_2V_2$$

If the equation is rearranged, then

$$V_2 = \frac{p_1V_1}{p_2}$$

Step 2 : Substitute the known values into the equation, making sure that the units for each variable are the same. Calculate the unknown variable.

$$V_2 = \frac{100 \times 160}{80} = 200cm^3$$

The volume occupied by the gas at a pressure of 80kPa, is 200 cm^3

Worked Example 27: Boyle's Law 2

Question: The pressure on a 2.5 l volume of gas is increased from 695 Pa to 755 Pa while a constant temperature is maintained. What is the volume of the gas under these pressure conditions?

Answer

Step 1 : Write down all the information that you know about the gas.

V_1 = 2.5 l and V_2 = ?

p_1 = 695 Pa and p_2 = 755 Pa

Step 2 : Choose a relevant gas law equation to calculate the unknown variable.

At constant temperature,

$$p_1V_1 = p_2V_2$$

Therefore,

$$V_2 = \frac{p_1V_1}{p_2}$$

Step 3 : Substitute the known values into the equation, making sure that the units for each variable are the same. Calculate the unknown variable.

$$V_2 = \frac{695 \times 2.5}{755} = 2.3l$$

Important:

It is not necessary to convert to Standard International (SI) units in the examples we have used above. Changing pressure and volume into different units involves *multiplication*. If you were to change the units in the above equation, this would involve multiplication on both sides of the equation, and so the conversions cancel each other out. However, although *SI* units don't have to be used, you must make sure that for each variable you use the *same* units throughout the equation. This is not true for some of the calculations we will do at a later stage, where SI units *must* be used.

Exercise: Boyle's Law

1. An unknown gas has an initial pressure of 150 kPa and a volume of 1 L. If the volume is increased to 1.5 L, what will the pressure now be?
2. A bicycle pump contains 250 cm^3 of air at a pressure of 90 kPa. If the air is compressed, the volume is reduced to 200 cm^3. What is the pressure of the air inside the pump?
3. The air inside a syringe occupies a volume of 10 cm^3 and exerts a pressure of 100 kPa. If the end of the syringe is sealed and the plunger is pushed down, the pressure increases to 120 kPa. What is the volume of the air in the syringe?
4. During an investigation to find the relationship between the pressure and volume of an enclosed gas at constant temperature, the following results were obtained.

Volume (cm^3)	Pressure (kPa)
40	125.0
30	166.7
25	200.0

(a) For the results given in the above table, plot a graph of **pressure** (y-axis) against the **inverse of volume** (x-axis).

(b) From the graph, deduce the relationship between the pressure and volume of an enclosed gas at constant temperature.

(c) Use the graph to predict what the volume of the gas would be at a pressure of 40 kPa. Show on your graph how you arrived at your answer.

(*IEB 2004 Paper 2*)

8.3 Charles's Law: Volume and Temperature of an enclosed gas

Charles's law describes the relationship between the **volume** and **temperature** of a gas. The law was first published by Joseph Louis Gay-Lussac in 1802, but he referenced unpublished work by Jacques Charles from around 1787. This law states that at constant pressure, the volume of a given mass of an ideal gas increases or decreases by the same factor as its temperature (in kelvin) increases or decreases. Another way of saying this is that temperature and volume are **directly proportional** (figure **??**).

Definition: Charles's Law
The volume of an enclosed sample of gas is directly proportional to its absolute temperature provided the pressure is kept constant.

Interesting Fact

Charles's Law is also known as Gay-Lussac's Law. This is because Charles did not publish his discovery, and it was rediscovered independently by another French Chemist Joseph Louis Gay-Lussac some years later.

Activity :: Demonstration : Charles's Law
Aim:
To demonstrate Charles's Law using simple materials.
Apparatus:
glass bottle (e.g. empty glass coke bottle), balloon, bunsen burner, retort stand
Method:

1. Place the balloon over the opening of the empty bottle.
2. Place the bottle on the retort stand over the bunsen burner and allow it to heat up. Observe what happens to the balloon. WARNING: Be careful when handling the heated bottle. You may need to wear gloves for protection.

Results:

You should see that the balloon starts to expand. As the air inside the bottle is heated, the pressure also increases, causing the volume to increase. Since the volume of the glass bottle can't increase, the air moves into the balloon, causing it to expand.

Conclusion:

The temperature and volume of the gas are directly related to each other. As one increases, so does the other.

Mathematically, the relationship between temperature and pressure can be represented as follows:

$$V \propto T$$

or

$$V = kT$$

If the equation is rearranged, then...

$$\frac{V}{T} = k$$

and, following the same logic that was used for Boyle's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

The equation relating volume and temperature produces a straight line graph (refer back to the notes on proportionality if this is unclear). This relationship is shown in figure 8.3.

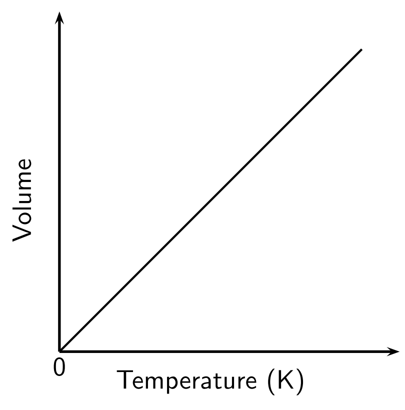


Figure 8.3: The volume of a gas is directly proportional to its temperature, provided the pressure of the gas is constant.

However, if this graph is plotted on a **celsius** temperature scale, the zero point of temperature doesn't correspond to the zero point of volume. When the volume is zero, the temperature is actually -273.15^{0}C (figure 8.4.

A new temperature scale, the **Kelvin scale** must be used instead. Since zero on the Celsius scale corresponds with a Kelvin temperature of -273.15^{0}C, it can be said that:

Kelvin temperature (T) = Celsius temperature (t) + 273.15

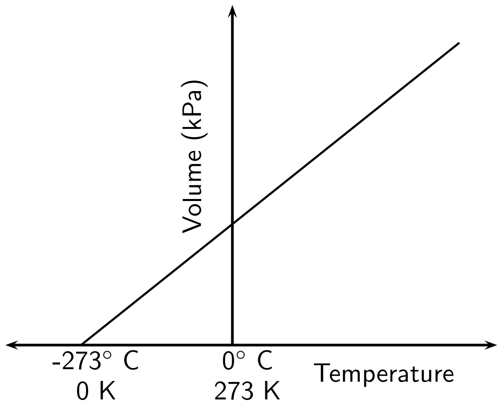


Figure 8.4: The relationship between volume and temperature, shown on a Celsius temperature scale.

At school level, you can simplify this slightly and convert between the two temperature scales as follows:

$$T = t + 273$$
or
$$t = T - 273$$

Can you explain Charles's law in terms of the kinetic theory of gases? When the temperature of a gas increases, so does the average speed of its molecules. The molecules collide with the walls of the container more often and with greater impact. These collisions will push back the walls, so that the gas occupies a greater volume than it did at the start. We saw this in the first demonstration. Because the glass bottle couldn't expand, the gas pushed out the balloon instead.

?

Exercise: Charles's law

The table below gives the temperature (in ^{0}C) of a number of gases under different volumes at a constant pressure.

Volume (l)	He	H_2	N_2O
0	-272.4	-271.8	-275.0
0.25	-245.5	-192.4	-123.5
0.5	-218.6	-113.1	28.1
0.75	-191.8	-33.7	179.6
1.0	-164.9	45.7	331.1
1.5	-111.1	204.4	634.1
2	-57.4	363.1	937.2
2.5	-3.6	521.8	1240.2
3.0	50.2	680.6	1543.2
3.5	103.9	839.3	1846.2

1. On the same set of axes, draw graphs to show the relationship between temperature and volume for each of the gases.
2. Describe the relationship you observe.
3. If you extrapolate the graphs (in other words, extend the graph line even though you may not have the exact data points), at what temperature do they intersect?
4. What is significant about this temperature?

Worked Example 28: Charles's Law 1

Question: Ammonium chloride and calcium hydroxide are allowed to react. The ammonia that is released in the reaction is collected in a gas syringe and sealed in. This gas is allowed to come to room temperature which is 32°C. The volume of the ammonia is found to be 122 ml. It is now placed in a water bath set at 7°C. What will be the volume reading after the syringe has been left in the bath for a good while (assume the plunger moves completely freely)?

Answer

Step 1 : Write down all the information that you know about the gas.

V_1 = 122 ml and V_2 = ?

T_1 = 32^0C and T_2 = 7^0C

Step 2 : Convert the known values to SI units if necessary.

Here, temperature must be converted into Kelvin, therefore:

T_1 = 32 + 273 = 305 K

T_2 = 7 + 273 = 280 K

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Therefore,

$$V_2 = \frac{V_1 \times T_2}{T_1}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$V_2 = \frac{122 \times 280}{305} = 112ml$$

Important:

Note that here the temperature must be converted to Kelvin (SI) since the change from degrees Celcius involves addition, not multiplication by a fixed conversion ratio (as is the case with pressure and volume.)

Worked Example 29: Charles's Law 2

Question: At a temperature of 298 K, a certain amount of CO_2 gas occupies a volume of 6 l. What volume will the gas occupy if its temperature is reduced to 273 K?

Answer

Step 1 : Write down all the information that you know about the gas.
$V_1 = 6$ l and $V_2 = ?$
$T_1 = 298$ K and $T_2 = 273$ K

Step 2 : Convert the known values to SI units if necessary.
Temperature data is already in Kelvin, and so no conversions are necessary.

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Therefore,

$$V_2 = \frac{V_1 \times T_2}{T_1}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$V_2 = \frac{6 \times 273}{298} = 5.5l$$

8.4 The relationship between temperature and pressure

The pressure of a gas is directly proportional to its temperature, if the volume is kept constant (figure 8.5). When the temperature of a gas increases, so does the energy of the particles. This causes them to move more rapidly and to collide with each other and with the side of the container more often. Since pressure is a measure of these collisions, the pressure of the gas increases with an increase in temperature. The pressure of the gas will decrease if its temperature decreases.

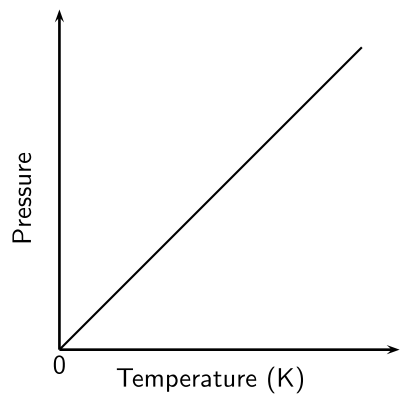


Figure 8.5: The relationship between the temperature and pressure of a gas

In the same way that we have done for the other gas laws, we can describe the relationship between temperature and pressure using symbols, as follows:

$$T \propto p, \text{ therefore } p = kT$$

We can also say that:

$$\frac{p}{T} = k$$

and that, provided the *amount* of gas stays the same...

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

?

Exercise: More gas laws

1. A gas of unknown volume has a temperature of 14°C. When the temperature of the gas is increased to 100°C, the volume is found to be 5.5 L. What was the initial volume of the gas?
2. A gas has an initial volume of 2600 mL and a temperature of 350 K.
   (a) If the volume is reduced to 1500 mL, what will the temperature of the gas be in Kelvin?
   (b) Has the temperature *increased* or *decreased*?
   (c) Explain this change, using the kinetic theory of matter.
3. A cylinder of propane gas at a temperature of 20°C exerts a pressure of 8 atm. When a cylinder has been placed in sunlight, its temperature increases to 25°C. What is the pressure of the gas inside the cylinder at this temperature?

8.5 The general gas equation

All the gas laws we have described so far rely on the fact that at least one variable (T, p or V) remains constant. Since this is unlikely to be the case most times, it is useful to combine the relationships into one equation. These relationships are as follows:

Boyle's law: p $\propto \frac{1}{V}$ (constant T)

Relationship between p and T: p $\propto$ T (constant V)

If we combine these relationships, we get p $\propto \frac{T}{V}$

If we introduce the proportionality constant k, we get p = k$\frac{T}{V}$

or, rearranging the equation...

$$pV = kT$$

We can also rewrite this relationship as follows:

$$\frac{pV}{T} = k$$

Provided the mass of the gas stays the same, we can also say that:

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

In the above equation, the subscripts 1 and 2 refer to two pressure and volume readings for the same mass of gas under different conditions. This is known as the **general gas equation**.

Temperature is always in kelvin and the units used for pressure and volume must be the same on both sides of the equation.

Important:
Remember that the general gas equation only applies if the mass of the gas is fixed.

Worked Example 30: General Gas Equation 1

Question: At the beginning of a journey, a truck tyre has a volume of 30 dm^3 and an internal pressure of 170 kPa. The temperature of the tyre is 16^{0}C. By the end of the trip, the volume of the tyre has increased to 32 dm^3 and the temperature of the air inside the tyre is 35^{0}C. What is the tyre pressure at the end of the journey?

Answer
Step 1 : Write down all the information that you know about the gas.
p_1 = 170 kPa and p_2 = ?
V_1 = 30 dm^3 and V_2 = 32 dm^3
T_1 = 16^{0}C and T_2 = 40^{0}C

Step 2 : Convert the known values to SI units if necessary.
Here, temperature must be converted into Kelvin, therefore:
T_1 = 16 + 273 = 289 K
T_2 = 40 + 273 = 313 K

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.
Use the general gas equation to solve this problem:

$$\frac{p_1 \times V_1}{T_1} = \frac{p_2 \times V_2}{T_2}$$

Therefore,

$$p_2 = \frac{p_1 \times V_1 \times T_2}{T_1 \times V_2}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$p_2 = \frac{170 \times 30 \times 313}{289 \times 32} = 173kPa$$

The pressure of the tyre at the end of the journey is 173 kPa.

Worked Example 31: General Gas Equation 2

Question: A cylinder that contains methane gas is kept at a temperature of 15^{0}C and exerts a pressure of 7 atm. If the temperature of the cylinder increases to 25^{0}C, what pressure does the gas now exert? (Refer to table 8.1 to see what an 'atm' is.

Answer
Step 1 : Write down all the information that you know about the gas.
p_1 = 7 atm and p_2 = ?

$T_1 = 15^0C$ and $T_2 = 25^0C$

Step 2 : Convert the known values to SI units if necessary.
Here, temperature must be converted into Kelvin, therefore:
$T_1 = 15 + 273 = 288$ K
$T_2 = 25 + 273 = 298$ K

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.
Since the volume of the cylinder is constant, we can write:

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

Therefore,

$$p_2 = \frac{p_1 \times T_2}{T_1}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$p_2 = \frac{7 \times 298}{288} = 7.24atm$$

The pressure of the gas is 7.24 atm.

Worked Example 32: General Gas Equation 3

Question: A gas container can withstand a pressure of 130 kPa before it will start to leak. Assuming that the volume of the gas in the container stays the same, at what temperature will the container start to leak if the gas exerts a pressure of 100 kPa at 15^0C?

Answer
Step 1 : Write down all the information that you know about the gas.
$p_1 = 100$ kPa and $p_2 = 130$ kPa
$T_1 = 15^0C$ and $T_2 = ?$

Step 2 : Convert the known values to SI units if necessary.
Here, temperature must be converted into Kelvin, therefore:
$T_1 = 15 + 273 = 288$ K

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.
Since the volume of the container is constant, we can write:

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

Therefore,

$$\frac{1}{T_2} = \frac{p_1}{T_1 \times p_2}$$

Therefore,

$$T_2 = \frac{T_1 \times p_2}{p_1}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$T_2 = \frac{288 \times 130}{100} = 374.4K = 101.4^0C$$

Exercise: The general gas equation

1. A closed gas system initially has a volume of 8 L and a temperature of 100°C. The pressure of the gas is unknown. If the temperature of the gas decreases to 50°C, the gas occupies a volume of 5 L. If the pressure of the gas under these conditions is 1.2 atm, what was the initial pressure of the gas?
2. A balloon is filled with helium gas at 27°C and a pressure of 1.0 atm. As the balloon rises, the volume of the balloon increases by a factor of 1.6 and the temperature decreases to 15°C. What is the final pressure of the gas (assuming none has escaped)?
3. 25 cm^3 of gas at 1 atm has a temperature of 20°C. When the gas is compressed to 20 cm^3, the temperature of the gas increases to 28°C. Calculate the final pressure of the gas.

8.6 The ideal gas equation

In the early 1800's, Amedeo Avogadro hypothesised that if you have samples of different gases, of the same volume, at a fixed temperature and pressure, then the samples must contain the same number of freely moving particles (i.e. atoms or molecules).

Definition: Avogadro's Law
Equal volumes of gases, at the same temperature and pressure, contain the same number of molecules.

You will remember from an earlier section, that we combined different gas law equations to get *one* that included temperature, volume and pressure. In this equation, pV = kT, the value of k is different for different masses of gas. If we were to measure the amount of gas in moles, then k = nR, where n is the number of moles of gas and R is the universal gas constant. The value of R is 8.3143 J.K^{-1}, or for most calculations, 8.3 J.K^{-1}. So, if we replace k in the general gas equation, we get the following **ideal gas equation**.

$$pV = nRT$$

Important:

1. The value of R is the same for all gases
2. All quantities in the equation pV = nRT must be in the same units as the value of R. In other words, SI units must be used throughout the equation.

The following table may help you when you convert to SI units.

Table 8.1: Conversion table showing different units of measurement for volume, pressure and temperature.

Variable	Pressure (**p**)	Volume (**V**)	moles (**n**)	universal gas constant (**R**)	temperature (**K**)
SI unit	Pascals (Pa)	m^3	mol	J.mol.K^{-1}	kelvin (K)
Other units and conversions	760 mm Hg = 1 atm = 101325 Pa = 101.325 kPa	1 m^3 = 1000000 cm^3 = 1000 dm^3 = 1000 litres			K = ^{0}C + 273

Worked Example 33: Ideal gas equation 1

Question: Two moles of oxygen (O_2) gas occupy a volume of 25 dm^3 at a temperature of 40^{0}C. Calculate the pressure of the gas under these conditions.

Answer

Step 1 : Write down all the information that you know about the gas.

p = ?
V = 25 dm^3
n = 2
T = 40^{0}C

Step 2 : Convert the known values to SI units if necessary.

$$V = \frac{25}{1000} = 0.025m^3$$

$$T = 40 + 273 = 313K$$

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$pV = nRT$$

Therefore,

$$p = \frac{nRT}{V}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$p = 2 \times 8.3 \times 3130.025 = 207832Pa = 207.8kPa$$

Worked Example 34: Ideal gas equation 2

Question: Carbon dioxide (CO_2) gas is produced as a result of the reaction between calcium carbonate and hydrochloric acid. The gas that is produced is collected in a 20 dm^3 container. The pressure of the gas is 105 kPa at a temperature of 20^{0}C.

What mass of carbon dioxide was produced?

Answer

Step 1 : Write down all the information that you know about the gas.

p = 105 kPa
V = 20 dm^3
T = 20^{0}C

Step 2 : Convert the known values to SI units if necessary.

$$p = 105 \times 1000 = 105000 Pa$$

$$T = 20 + 273 = 293K$$

$$V = \frac{20}{1000} = 0.02 m^3$$

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$pV = nRT$$

Therefore,

$$n = \frac{pV}{RT}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$n = \frac{105000 \times 0.02}{8.3 \times 293} = 0.86 moles$$

Step 5 : Calculate mass from moles

$$n = \frac{m}{M}$$

Therefore,

$$m = n \times M$$

The molar mass of CO_2 is calculated as follows:

$$M = 12 + (2 \times 16) = 44 g.mol^{-1}$$

Therefore,

$$m = 0.86 \times 44 = 37.84 g$$

Worked Example 35: Ideal gas equation 3

Question: 1 mole of nitrogen (N_2) reacts with hydrogen (H_2) according to the following equation:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

The ammonia (NH_3) gas is collected in a separate gas cylinder which has a volume of 25 dm^3. The temperature of the gas is 22^0C. Calculate the pressure of the gas inside the cylinder.

Answer

Step 1 : Write down all the information that you know about the gas.

V = 25 dm^3

n = 2 (Calculate this by looking at the mole ratio of nitrogen to ammonia, which is 1:2)

T = 22^0C

Step 2 : Convert the known values to SI units if necessary.

$$V = \frac{25}{1000} = 0.025m^3$$

$$T = 22 + 273 = 295K$$

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$pV = nRT$$

Therefore,

$$p = \frac{nRT}{V}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$p = \frac{2 \times 8.3 \times 295}{0.025} = 195880Pa = 195.89kPa$$

Worked Example 36: Ideal gas equation 4

Question: Calculate the number of air particles in a 10 m by 7 m by 2 m classroom on a day when the temperature is 23°C and the air pressure is 98 kPa.

Answer

Step 1 : Write down all the information that you know about the gas.

V = 10 m × 7 m × 2m = 140 m^3

p = 98 kPa

T = 23^0C

Step 2 : Convert the known values to SI units if necessary.

$$p = 98 \times 1000 = 98000Pa$$

$$T = 23 + 273 = 296K$$

Step 3 : Choose a relevant gas law equation that will allow you to calculate the unknown variable.

$$pV = nRT$$

Therefore,

$$n = \frac{pV}{RT}$$

Step 4 : Substitute the known values into the equation. Calculate the unknown variable.

$$n = \frac{98000 \times 140}{8.3 \times 296} = 5584.5mol$$

Worked Example 37: Applying the gas laws

Question: Most modern cars are equipped with airbags for both the driver and the passenger. An airbag will completely inflate in 0,05 s. This is important because a typical car collision lasts about 0,125 s. The following reaction of sodium azide (a compound found in airbags) is activated by an electrical signal:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

1. Calculate the mass of $N_2(g)$ needed to inflate a sample airbag to a volume of 65 dm^3 at 25 °C and $99{,}3$ kPa. Assume the gas temperature remains constant during the reaction.
2. In reality the above reaction is exothermic. Describe, in terms of the kinetic molecular theory, how the pressure in the sample airbag will change, if at all, as the gas temperature returns to 25 °C.

Answer

Step 1 : Look at the information you have been given, and the information you still need.

Here you are given the volume, temperature and pressure. You are required to work out the mass of N_2.

Step 2 : Check that all the units are S.I. units

Pressure: $93{,}3 \times 10^3$ Pa
Volume: 65×10^{-3} m^3
Temperature: $(273 + 25)$ K
Gas Constant: $8{,}31$

Step 3 : Write out the Ideal Gas formula

$$pV = nRT$$

Step 4 : Solve for the required quantity using symbols

$$n = \frac{pV}{RT}$$

Step 5 : Solve by substituting numbers into the equation to solve for 'n'.

$$n = \frac{99{,}3 \times 10^3 \times 65 \times 10^{-3}}{8{,}31 \times (273 + 25)}$$

Step 6 : Convert the number of moles to number of grams

$$m = n \times M$$
$$m = 2{,}61 \times 28$$
$$m = 73{,}0g$$

Step 7 : Theory Question
When the temperature decreases the intensity of collisions with the walls of the airbag and between particles decreases. Therefore pressure decreases.

?

Exercise: The ideal gas equation

1. An unknown gas has pressure, volume and temperature of 0.9 atm, 8 L and 120°C respectively. How many moles of gas are present?
2. 6 g of chlorine (Cl_2) occupies a volume of 0.002 m^3 at a temperature of 26°C. What is the pressure of the gas under these conditions?
3. An average pair of human lungs contains about 3.5 L of air after inhalation and about 3.0 L after exhalation. Assuming that air in your lungs is at 37°C and 1.0 atm, determine the number of moles of air in a typical breath.
4. A learner is asked to calculate the answer to the problem below:
   Calculate the pressure exerted by 1.5 moles of nitrogen gas in a container with a volume of 20 dm^3 at a temperature of 37°C.
   The learner writes the solution as follows:
   V = 20 dm^3
   n = 1.5 mol
   R = 8.3 J.K^{-1}.mol^{-1}
   T = 37 + 273 = 310 K
   p = nRT, therefore
   $$p = \frac{nRV}{T}$$
   $$= \frac{1.5 \times 8.3 \times 20}{310}$$
   = 0.8 kPa
   (a) Identify 2 mistakes the learner has made in the calculation.
   (b) Are the units of the final answer correct?
   (c) Rewrite the solution, correcting the mistakes to arrive at the right answer.

8.7 Molar volume of gases

It is possible to calculate the volume of a mole of gas at STP using what we now know about gases.

1. *Write down the ideal gas equation*

pV = nRT, therefore V = $\frac{nRT}{p}$

2. *Record the values that you know, making sure that they are in SI units*

   You know that the gas is under STP conditions. These are as follows:

   p = 101.3 kPa = 101300 Pa

   n = 1 mole

   R = 8.3 J.K^{-1}.mol^{-1}

   T = 273 K

3. *Substitute these values into the original equation.*

$$V = \frac{nRT}{p}$$

$$V = \frac{1mol \times 8.3J.K^{-1}.mol^{-1} \times 273K}{101300Pa}$$

4. *Calculate the volume of 1 mole of gas under these conditions*

   The volume of 1 mole of gas at STP is 22.4 × 10^{-3} m^3 = 22.4 dm^3.

8.8 Ideal gases and non-ideal gas behaviour

In looking at the behaviour of gases to arrive at the Ideal Gas Law, we have limited our examination to a small range of temperature and pressure. Most gases do obey these laws most of the time, and are called **ideal gases**, but there are deviations at **high pressures** and **low temperatures**. So what is happening at these two extremes?

Earlier when we discussed the kinetic theory of gases, we made a number of assumptions about the behaviour of gases. We now need to look at two of these again because they affect how gases behave either when pressures are high or when temperatures are low.

1. *Molecules do occupy volume*

   This means that when pressures are very high and the molecules are compressed, their volume becomes significant. This means that the total volume available for the gas molecules to move is reduced and collisions become more frequent. This causes the pressure of the gas to be *higher* than what would normally have been predicted by Boyle's law (figure 8.6).

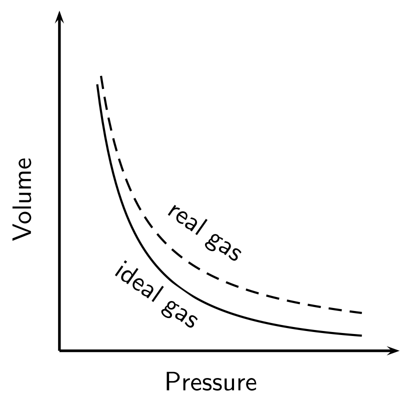


Figure 8.6: Gases deviate from ideal gas behaviour at high pressure.

2. *Forces of attraction do exist between molecules*

   At low temperatures, when the speed of the molecules decreases and they move closer together, the intermolecular forces become more apparent. As the attraction between molecules increases, their movement decreases and there are fewer collisions between them. The pressure of the gas at low temperatures is therefore lower than what would have been expected for an ideal gas (figure 8.7. If the temperature is low enough or the pressure high enough, a real gas will **liquify**.

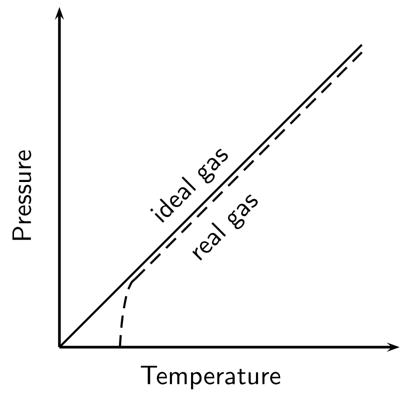


Figure 8.7: Gases deviate from ideal gas behaviour at low temperatures

8.9 Summary

- The **kinetic theory of matter** helps to explain the behaviour of gases under different conditions.
- An **ideal gas** is one that obeys all the assumptions of the kinetic theory.
- A **real gas** behaves like an ideal gas, except at high pressures and low temperatures. Under these conditions, the forces between molecules become significant and the gas will liquify.
- **Boyle's law** states that the pressure of a fixed quantity of gas is inversely proportional to its volume, as long as the temperature stays the same. In other words, pV = k or

$$\mathrm{p_1V_1 = p_2V_2}.$$

- **Charles's law** states that the volume of an enclosed sample of gas is directly proportional to its temperature, as long as the pressure stays the same. In other words,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- The **temperature** of a fixed mass of gas is directly proportional to its pressure, if the volume is constant. In other words,

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

- In the above equations, temperature must be written in **Kelvin**. Temperature in degrees Celsius (temperature = t) can be converted to temperature in Kelvin (temperature = T) using the following equation:

$$T = t + 273$$

- Combining Boyle's law and the relationship between the temperature and pressure of a gas, gives the **general gas equation**, which applies as long as the amount of gas remains constant. The general gas equation is pV = kT, or

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

- Because the mass of gas is not always constant, another equation is needed for these situations. The **ideal gas equation** can be written as

$$pV = nRT$$

where n is the number of moles of gas and R is the universal gas constant, which is 8.3 $J.K^{-1}.mol^{-1}$. In this equation, **SI units** must be used. Volume (m^3), pressure (Pa) and temperature (K).

- The **volume** of one mole of gas under STP is 22.4 dm^3. This is called the **molar gas volume**.

s

?

Exercise: Summary exercise

1. For each of the following, say whether the statement is **true** or **false**. If the statement is false, rewrite the statement correctly.
   (a) Real gases behave like ideal gases, except at low pressures and low temperatures.
   (b) The volume of a given mass of gas is inversely proportional to the pressure it exerts.
   (c) The temperature of a fixed mass of gas is directly proportional to its pressure, regardless of the volume of the gas.
2. For each of the following multiple choice questions, choose the **one correct answer**.
   (a) Which one of the following properties of a fixed quantity of a gas must be kept constant during a Boyle's law investigation?
      i. density
      ii. pressure
      iii. temperature
      iv. volume
      (*IEB 2003 Paper 2*)
   (b) Three containers of EQUAL VOLUME are filled with EQUAL MASSES of helium, nitrogen and carbon dioxide gas respectively. The gases in the three containers are all at the same TEMPERATURE. Which one of the following statements is correct regarding the pressure of the gases?
      i. All three gases will be at the same pressure
      ii. The helium will be at the greatest pressure
      iii. The nitrogen will be at the greatest pressure
      iv. The carbon dioxide will be at the greatest pressure
      (*IEB 2004 Paper 2*)
   (c) One mole of an ideal gas is stored at a temperature T (in Kelvin) in a rigid gas tank. If the average speed of the gas particles is doubled, what is the new Kelvin temperature of the gas?
      i. 4T
      ii. 2T
      iii. $\sqrt{2}$T

iv. 0.5 T

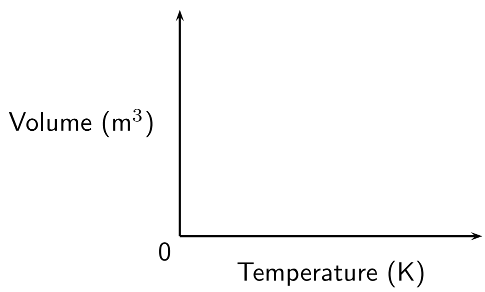
(*IEB 2002 Paper 2*)

(d) The ideal gas equation is given by $\mathbf{pV = nRT}$. Which one of the following conditions is true according to Avogadro's hypothesis?

a	$p \propto 1/V$	(T = constant)
b	$V \propto T$	(p = constant)
c	$V \propto n$	(p, T = constant)
d	$p \propto T$	(n = constant)

(*DoE Exemplar paper 2, 2007*)

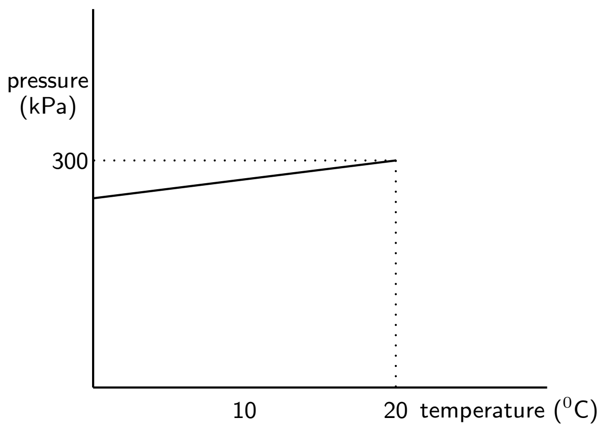
3. Use your knowledge of the gas laws to explain the following statements.
   (a) It is dangerous to put an aerosol can near heat.
   (b) A pressure vessel that is poorly designed and made can be a serious safety hazard (a pressure vessel is a closed, rigi container that is used to hold gases at a pressure that is higher than the normal air pressure).
   (c) The volume of a car tyre increases after a trip on a hot road.
4. Copy the following set of labelled axes and answer the questions that follow:



   (a) On the axes, **using a solid line**, draw the graph that would be obtained for a fixed mass of an ideal gas if the pressure is kept constant.
   (b) If the gradient of the above graph is measured to be 0.008 $m^3.K^{-1}$, calculate the pressure that 0.3 mol of this gas would exert.

   (**IEB 2002 Paper 2**)
5. Two gas cylinders, A and B, have a volume of 0.15 m^3 and 0.20 m^3 respectively. Cylinder A contains 1.25 mol He gas at pressure p and cylinder B contains 2.45 mol He gas at standard pressure. The ratio of the Kelvin temperatures A:B is 1.80:1.00. Calculate the pressure of the gas (in kPa) in cylinder A.

   (*IEB 2002 Paper 2*)
6. A learner investigates the relationship between the Celsius temperature and the pressure of a fixed amount of helium gas in a 500 cm^3 closed container. From the results of the investigation, she draws the graph below:



   (a) Under the conditions of this investigation, helium gas behaves like an ideal gas. Explain briefly why this is so.

(b) From the shape of the graph, the learner concludes that the pressure of the helium gas is directly proportional to the Celcius temperature. Is her conclusion correct? Briefly explain your answer.

(c) Calculate the pressure of the helium gas at 0 °C.

(d) Calculate the mass of helium gas in the container.

(*IEB 2003 Paper 2*)

7. One of the cylinders of a motor car engine, before compression contains 450 cm^3 of a mixture of air and petrol in the gaseous phase, at a temperature of 30°C and a pressure of 100 kPa. If the volume of the cylinder after compression decreases to one tenth of the original volume, and the temperature of the gas mixture rises to 140°C, calculate the pressure now exerted by the gas mixture.

8. In an experiment to determine the relationship between pressure and temperature of a fixed mass of gas, a group of learners obtained the following results:

Pressure (kPa)	101	120	130.5	138
Temperature (^{0}C)	0	50	80	100
Total gas volume (cm^3)	250	250	250	250

(a) Draw a straight-line graph of pressure (on the dependent, y-axis) versus temperature (on the independent, x-axis) on a piece of graph paper. Plot the points. Give your graph a suitable heading.
A straight-line graph passing through the origin is essential to obtain a mathematical relationship between pressure and temperature.

(b) Extrapolate (extend) your graph and determine the temperature (in ^{0}C) at which the graph will pass through the temperature axis.

(c) Write down, in words, the relationship between pressure and Kelvin temperature.

(d) From your graph, determine the pressure (in kPa) at 173 K. Indicate on your graph how you obtained this value.

(e) How would the gradient of the graph be affected (if at all) if a larger mass of the gas is used? Write down ONLY **increases**, **decreases** or **stays the same**.

(*DoE Exemplar Paper 2, 2007*)

Chapter 9

Organic Molecules - Grade 12

9.1 What is organic chemistry?

Organic chemistry is the branch of chemistry that deals with **organic molecules**. An organic molecule is one which contains **carbon**, and these molecules can range in size from simple molecules to complex structures containing thousands of atoms! Although the main element in organic compounds is carbon, other elements such as hydrogen (H), oxygen (O), nitrogen (N), sulfur (S) and phosphorus (P) are also common in these molecules.

Until the early nineteenth century, chemists had managed to make many simple compounds in the laboratory, but were still unable to produce the complex molecules that they found in living organisms. It was around this time that a Swedish chemist called **Jons Jakob Berzelius** suggested that compounds found only in living organisms (the organic compounds) should be grouped separately from those found in the non-living world (the inorganic compounds). He also suggested that the laws that governed how organic compounds formed, were different from those for inorganic compounds. From this, the idea developed that there was a 'vital force' in organic compounds. In other words, scientists believed that organic compounds would not follow the normal physical and chemical laws that applied to other inorganic compounds because the very 'force of life' made them different.

This idea of a mystical 'vital force' in organic compounds was weakened when scientists began to manufacture organic compounds in the laboratory from non-living materials. One of the first to do this was **Friedrich Wohler** in 1828, who successfully prepared urea, an organic compound in the urine of animals which, until that point, had only been found in animals. A few years later a student of Wohler's, **Hermann Kolbe**, made the organic compound *acetic acid* from inorganic compounds. By this stage it was acknowledged that organic compounds are governed by exactly the same laws that apply to inorganic compounds. The properties of organic compounds are not due to a 'vital force' but to the unique properties of the carbon atom itself.

Organic compounds are very important in daily life. They make up a big part of our own bodies, they are in the food we eat and in the clothes we wear. Organic compounds are also used to make products such as medicines, plastics, washing powders, dyes, along with a list of other items.

9.2 Sources of carbon

The main source of the carbon in organic compounds is **carbon dioxide** in the air. Plants use sunlight to convert carbon dioxide into organic compounds through the process of **photosynthesis**. Plants are therefore able to make their own organic compounds through photosynthesis, while animals feed on plants or plant products so that they gain the organic compounds that they need to survive.

Another important source of carbon is **fossil fuels** such as coal, petroleum and natural gas. This is because fossil fuels are themselves formed from the decaying remains of dead organisms (refer to chapter 21 for more information on fossil fuels).

9.3 Unique properties of carbon

Carbon has a number of unique properties which influence how it behaves and how it bonds with other atoms:

- Carbon has *four valence electrons* which means that each carbon atom can form bonds with four other atoms. Because of this, long *chain structures* can form. These chains can either be *unbranched* (figure 9.1) or *branched* (figure 9.2). Because of the number of bonds that carbon can form with other atoms, organic compounds can be very complex.

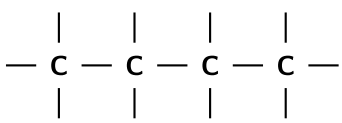


Figure 9.1: An unbranched carbon chain

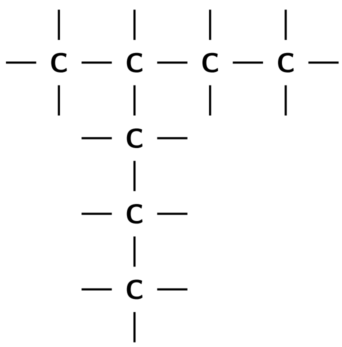


Figure 9.2: A branched carbon chain

- Because of its position on the Periodic Table, most of the bonds that carbon forms with other atoms are *covalent*. Think for example of a C-C bond. The difference in electronegativity between the two atoms is zero, so this is a pure covalent bond. In the case of a C-H bond, the difference in electronegativity between carbon (2.5) and hydrogen (2.1) is so small that C-H bonds are almost purely covalent. The result of this is that most organic compounds are non-polar. This affects some of the properties of organic compounds.

9.4 Representing organic compounds

There are a number of ways to represent organic compounds. It is useful to know all of these so that you can recognise a molecule however it is shown. There are three main ways of representing a compound. We will use the example of a molecule called 2-methylpropane to help explain the difference between each.

9.4.1 Molecular formula

The molecular formula of a compound shows how many atoms of each type are in a molecule. The number of each atom is written as a subscript after the atomic symbol. The molecular formula of 2-methylpropane is:

$$C_4H_{10}$$

9.4.2 Structural formula

The structural formula of an organic compound shows every bond between every atom in the molecule. Each bond is represented by a line. The structural formula of 2-methylpropane is shown in figure 9.3.

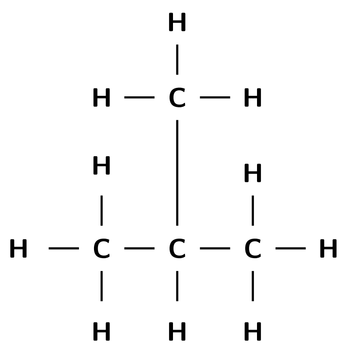


Figure 9.3: The structural formula of 2-methylpropane

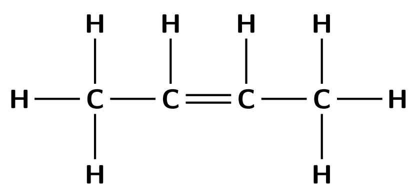
9.4.3 Condensed structural formula

When a compound is represented using its condensed structural formula, each carbon atom and the hydrogen atoms that are bonded directly to it are listed as a molecular formula, followed by a similar molecular formula for the neighbouring carbon atom. Branched groups are shown in brackets after the carbon atom to which they are bonded. The condensed structural formula below shows that in 2-methylpropane, there is a branched chain attached to the second carbon atom of the main chain. You can check this by looking at the structural formula in figure **??**.

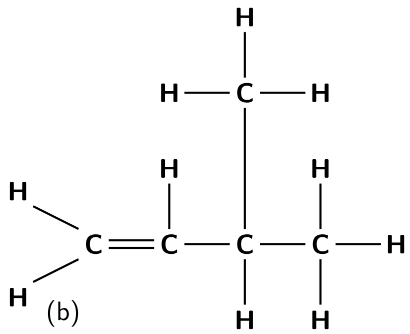
$$CH_3CH(CH_3)CH_3$$

Exercise: Representing organic compounds

1. For each of the following organic compounds, give the **condensed structural formula** and the **molecular formula**.



(a)



(b)

2. For each of the following, give the **structural formula** and the **molecular formula**.
   (a) $CH_3CH_2CH_3$
   (b) $CH_3CH_2CH(CH_3)CH_3$
   (c) C_2H_6
3. Give two possible structural formulae for the compound with a molecular formula of C_4H_{10}.

9.5 Isomerism in organic compounds

It is possible for two organic compounds to have the same *molecular formula* but a different *structural formula*. Look for example at the two organic compounds that are shown in figure 9.4.

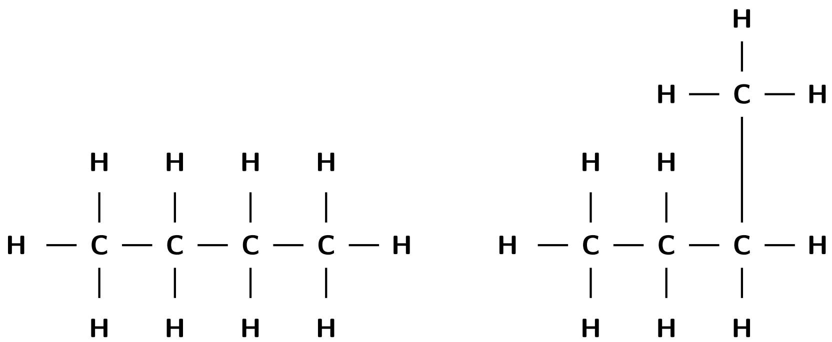
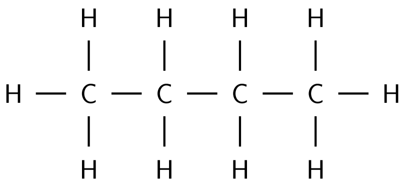
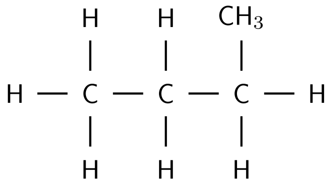
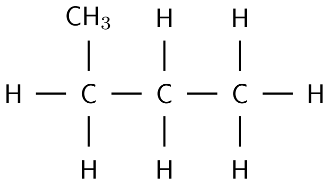


Figure 9.4: Isomers of a 4-carbon organic compound

If you were to count the number of carbon and hydrogen atoms in each compound, you would find that they are the same. They both have the same molecular formula (C_4H_{10}), but their structure is different and so are their properties. Such compounds are called **isomers**.

Definition: Isomer
In chemistry, isomers are molecules with the same molecular formula and often with the same kinds of chemical bonds between atoms, but in which the atoms are arranged differently.

Exercise: Isomers
Match the organic compound in Column A with its isomer Column B:

Column A	**Column B**
$CH_3CH(CH_3)OH$	$CH_3CH(CH_3)CH_3$
	
	C_3H_7OH

9.6 Functional groups

All organic compounds have a particular bond or group of atoms which we call its **functional group**. This group is important in determining how a compound will react.

Definition: Functional group
In organic chemistry, a functional group is a specific group of atoms within molecules, that are responsible for the characteristic chemical reactions of those molecules. The same functional group will undergo the same or similar chemical reaction(s) regardless of the size of the molecule it is a part of.

In one group of organic compounds called the **hydrocarbons**, the single, double and triple bonds of the alkanes, elkenes and alkynes are examples of functional groups. In another group, the alcohols, an oxygen and a hydrogen atom that are bonded to each other form the functional group for those compounds. All alcohols will contain an oxygen and a hydrogen atom bonded together in some part of the molecule.

Table 9.1 summarises some of the common functional groups. We will look at these in more detail later in this chapter.

9.7 The Hydrocarbons

Let us first look at a group of organic compounds known as the **hydrocarbons**. These molecules only contain carbon and hydrogen. The hydrocarbons that we are going to look at are called **aliphatic compounds**. The aliphatic compounds are divided into *acyclic compounds* (chain structures) and *cyclic compounds* (ring structures). The chain structures are further divided into structures that contain only *single bonds* (**alkanes**), those that contain *double bonds* (**alkenes**) and those that contain *triple bonds* (**alkynes**). Cyclic compounds include structures such as the *benzene ring*. Figure 9.5 summarises the classification of the hydrocarbons.

Hydrocarbons that contain only single bonds are called **saturated** hydrocarbons because each carbon atom is bonded to as many hydrogen atoms as possible. Figure 9.6 shows a molecule of ethane which is a saturated hydrocarbon.

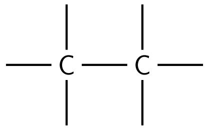
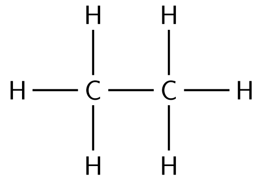
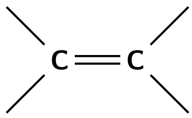
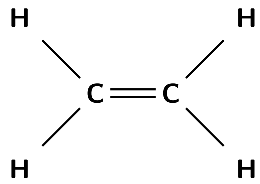
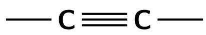
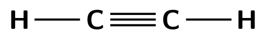
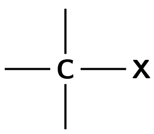
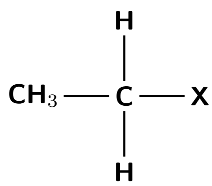
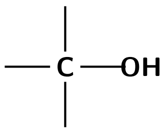
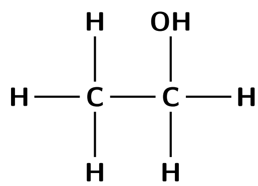
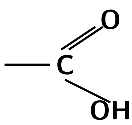
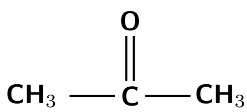
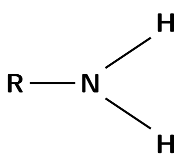
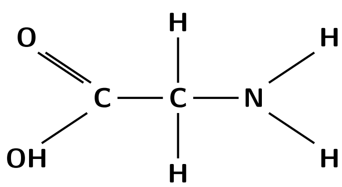
Name of group	Functional group	Example	Diagram
Alk*ane*		Ethane	
Alk*ene*		Ethene	
Alk*yne*		Ethyne (acetylene)	
Halo-alkane	 (X=F,Cl,Br,I)	Chloroethane	
Alcoh*ol*/ alkan*ol*		Ethanol	
Carboxylic acid		ethanoic acid	
Amine		Glycine	

Table 9.1: Some functional groups of organic compounds

Hydrocarbons that contain double or triple bonds are called **unsaturated** hydrocarbons because they don't contain as many hydrogen atoms as possible. Figure 9.7 shows a molecule of ethene which is an unsaturated hydrocarbon. If you compare the number of carbon and hydrogen atoms

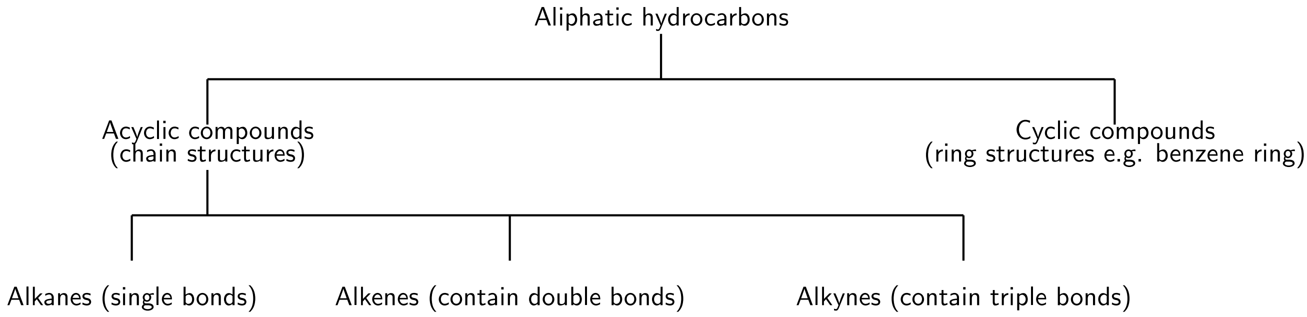


Figure 9.5: The classification of the aliphatic hydrocarbons

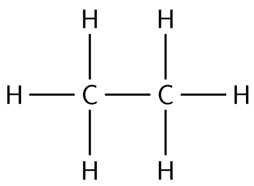


Figure 9.6: A saturated hydrocarbon

in a molecule of ethane and a molecule of ethene, you will see that the number of hydrogen atoms in ethene is *less* than the number of hydrogen atoms in ethane despite the fact that they both contain two carbon atoms. In order for an unsaturated compound to become saturated, a double bond has to be broken, and another two hydrogen atoms added for each double bond that is replaced by a single bond.

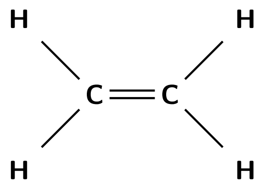


Figure 9.7: An unsaturated hydrocarbon

Interesting Fact

Fat that occurs naturally in living matter such as animals and plants is used as food for human consumption and contains varying proportions of saturated and unsaturated fat. Foods that contain a high proportion of saturated fat are butter, ghee, suet, tallow, lard, coconut oil, cottonseed oil, and palm kernel oil, dairy products (especially cream and cheese), meat, and some prepared foods. Diets high in saturated fat are correlated with an increased incidence of atherosclerosis and coronary heart disease according to a number of studies. Vegetable oils contain unsaturated fats and can be hardened to form margarine by adding hydrogen on to some of the carbon=carbon double bonds using a nickel catalyst. The process is called hydrogenation

We will now go on to look at each of the hydrocarbon groups in more detail. These groups are the alkanes, the alkenes and the alkynes.

9.7.1 The Alkanes

The alkanes are hydrocarbons that only contain *single covalent bonds* between their carbon atoms. This means that they are *saturated* compounds and are quite unreactive. The simplest alkane has only one carbon atom and is called **methane**. This molecule is shown in figure 9.8.

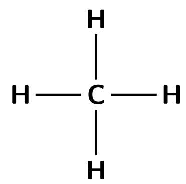
(a)  **(b)** CH_4

Figure 9.8: The structural (a) and molecular formula (b) for methane

The second alkane in the series has two carbon atoms and is called **ethane**. This is shown in figure 9.9.

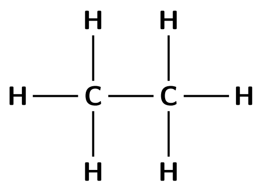
(a)  **(b)** C_2H_6

Figure 9.9: The structural (a) and molecular formula (b) for ethane

The third alkane in the series has three carbon atoms and is called **propane** (Figure 9.10).

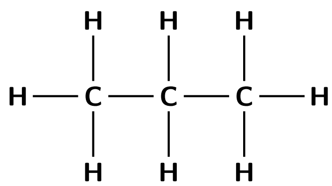
(a)  **(b)** C_3H_8

Figure 9.10: The structural (a) and molecular formula (b) for propane

When you look at the molecular formula for each of the alkanes, you should notice a pattern developing. For each carbon atom that is added to the molecule, two hydrogen atoms are added. In other words, each molecule differs from the one before it by CH_2. This is called a *homologous series*. The alkanes have the general formula C_nH_{2n+2}.

The alkanes are the most important source of fuel in the world and are used extensively in the chemical industry. Some are gases (e.g. methane and ethane), while others are liquid fuels (e.g. octane, an important component of petrol).

Interesting Fact

Some fungi use alkanes as a source of carbon and energy. One fungus *Amorphotheca resinae* prefers the alkanes used in aviation fuel, and this can cause problems for aircraft in tropical areas!

9.7.2 Naming the alkanes

In order to give compounds a name, certain rules must be followed. When naming organic compounds, the IUPAC (International Union of Pure and Applied Chemistry) nomenclature is used. We will first look at some of the steps that need to be followed when naming a compound, and then try to apply these rules to some specific examples.

1. STEP 1: Recognise the *functional group* in the compound. This will determine the suffix (the 'end') of the name. For example, if the compound is an alkane, the suffix will be -ane; if the compound is an alkene the suffix will be -ene; if the compound is an alcohol the suffix will be -ol, and so on.

2. STEP 2: Find the longest continuous carbon chain (it won't always be a *straight* chain) and count the number of carbon atoms in this chain. This number will determine the prefix (the 'beginning') of the compound's name. These prefixes are shown in table 9.2. So, for example, an alkane that has 3 carbon atoms will have the suffix *prop* and the compound's name will be *propane*.

Carbon atoms	prefix
1	meth(ane)
2	eth(ane)
3	prop(ane)
4	but(ane)
5	pent(ane)
6	hex(ane)
7	hept(ane)
8	oct(ane)
9	non(ane)
10	dec(ane)

Table 9.2: The prefix of a compound's name is determined by the number of carbon atoms in the longest chain

3. STEP 3: Number the carbons in the longest carbon chain (Important: If there is a double or triple bond, you need to start numbering so that the bond is at the carbon with the lowest number.

4. STEP 4: Look for any branched groups and name them. Also give them a number to show their position on the carbon chain. If there are no branched groups, this step can be left out.

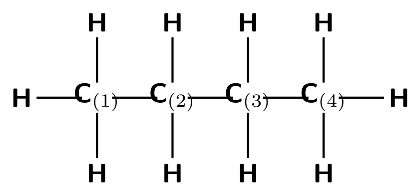
5. STEP 5: Combine the elements of the name into a single word in the following order: branched groups; prefix; name ending according to the functional group and its position along the longest carbon chain.

Worked Example 38: Naming the alkanes

Question: Give the IUPAC name for the following compound:
Note: The numbers attached to the carbon atoms would not normally be shown. The atoms have been numbered to help you to name the compound.

Answer
Step 1 : Identify the functional group
The compound is a hydrocarbon with single bonds between the carbon atoms. It is an alkane and will have a suffix of -ane.

Step 2 : Find the longest carbon chain



There are four carbon atoms in the longest chain. The prefix of the compound will be 'but'.

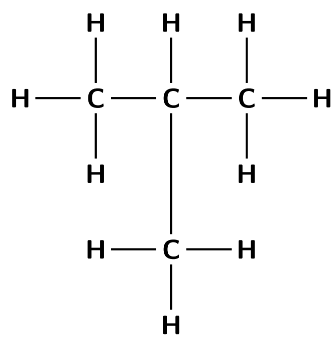
Step 3 : Number the carbons in the longest chain
In this case, it is easy. The carbons are numbered from left to right, from one to four.

Step 4 : Look for any branched groups, name them and give their position on the carbon chain
There are no branched groups in this compound.

Step 5 : Combine the elements of the name into a single word
The name of the compound is **butane**.

Worked Example 39: Naming the alkanes

Question: Give the IUPAC name for the following compound:

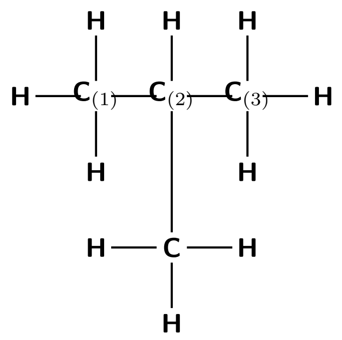


Answer
Step 1 : Identify the functional group
The compound is an alkane and will have the suffix -ane.

Step 2 : Find the longest carbon chain
There are three carbons in the longest chain. The prefix for this compound is -prop.

Step 3 : Number the carbons in the carbon chain
If we start at the carbon on the left, we can number the atoms as shown below:

Step 4 : Look for any branched groups, name them and give their position on the carbon chain
There is a branched group attached to the second carbon atom. This group has the formula CH_3 which is methane. However, because it is not part of the main chain, it is given the suffix -yl (i.e. methyl). The position of the methyl group comes just



before its name (see next step).

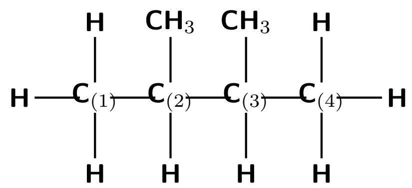
Step 5 : Combine the elements of the compound's name into a single word in the order of branched groups; prefix; name ending according to the functional group.
The compound's name is **2-methylpropane**.

Worked Example 40: Naming the alkanes

Question: Give the IUPAC name for the following compound:

$$CH_3CH(CH_3)CH(CH_3)CH_3$$

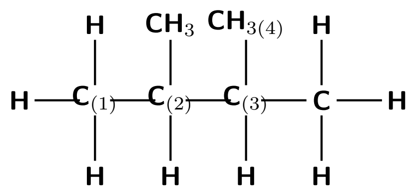
(Remember that the side groups are shown in brackets after the carbon atom to which they are attached.)
Answer
Step 1 : Draw the compound from its condensed structural formula
The structural formula of the compound is:



Step 2 : Identify the functional group
The compound is an alkane and will have the suffix -ane.

Step 3 : Find the longest carbon chain
There are four carbons in the longest chain. The prefix for this compound is -but.

Step 4 : Number the carbons in the carbon chain
If we start at the carbon on the left, carbon atoms are numbered as shown in the diagram above. A second way that the carbons could be numbered is:



Step 5 : Look for any branched groups, name them and give their position on the carbon chain

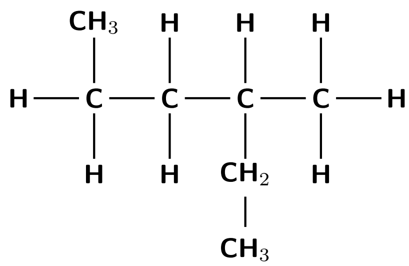
There are two methyl groups attached to the main chain. The first one is attached to the second carbon atom and the second methyl group is attached to the third carbon atom. Notice that in this example it does not matter how you have chosen to number the carbons in the main chain; the methyl groups are still attached to the second and third carbons and so the naming of the compound is not affected.

Step 6 : Combine the elements of the compound's name into a single word in the order of branched groups; prefix; name ending according to the functional group.

The compound's name is **2,3-dimethyl-butane**.

Worked Example 41: Naming the alkanes

Question: Give the IUPAC name for the following compound:



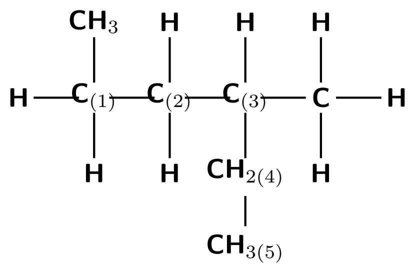
Answer

Step 1 : Identify the functional group

The compound is an alkane and will have the suffix -ane.

Step 2 : Find the longest carbon chain and number the carbons in the longest chain.

There are five carbons in the longest chain if they are numbered as shown below. The prefix for the compound is -pent.



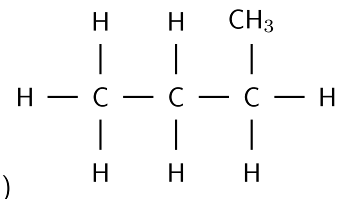
Step 3 : Look for any branched groups, name them and give their position on the carbon chain

There are two methyl groups attached to the main chain. The first one is attached to the first carbon atom and the second methyl group is attached to the third carbon atom.

Step 4 : Combine the elements of the compound's name into a single word in the order of branched groups; prefix; name ending according to the functional group.

The compound's name is **1,3-dimethyl-pentane**.

Exercise: Naming the alkanes

1. Give the structural formula for each of the following:
   (a) Octane
   (b) $CH_3CH_2CH_3$
   (c) $CH_3CH(CH_3)CH_3$
   (d) 3-ethyl-pentane
2. Give the IUPAC name for each of the following organic compounds.
   (a) 
   (b) $CH_3CH_2CH(CH_3)CH_2CH_3$
   (c) $CH_3CH(CH_3)CH_2CH(CH_3)CH_3$

9.7.3 Properties of the alkanes

We have already mentioned that the alkanes are relatively unreactive because of their stable C-C and C-H bonds. The boiling point and melting point of these molecules is determined by their molecular structure, and their surface area. The more carbon atoms there are in an alkane, the greater the surface area and therefore the higher the boiling point. The melting point also increases as the number of carbon atoms in the molecule increases. This can be seen in the data in table 9.3.

Formula	Name	Melting point (^{0}C)	Boiling point (^{0}C)	Phase at room temperature
CH_4	methane	-183	-162	gas
C_2H_6	ethane	-182	-88	gas
C_3H_8	propane	-187	-45	gas
C_4H_{10}	butane	-138	-0.5	gas
C_5H_{12}	pentane	-130	36	liquid
C_6H_{14}	hexane	-95	69	liquid
$C_{17}H_{36}$	heptadecane	22	302	solid

Table 9.3: Properties of some of the alkanes

You will also notice that, when the molecular mass of the alkanes is low (i.e. there are few carbon atoms), the organic compounds are *gases* because the intermolecular forces are weak. As the number of carbon atoms and the molecular mass increases, the compounds are more likely to be liquids or solids because the intermolecular forces are stronger.

9.7.4 Reactions of the alkanes

There are three types of reactions that can occur in saturated compounds such as the alkanes.

1. **Substitution reactions**
   Substitution reactions involve the removal of a hydrogen atom which is replaced by an atom of another element, such as a halogen (F, Cl, Br or I) (figure 9.11). The product is called a **halo-alkane**. Since alkanes are not very reactive, heat or light are needed for this

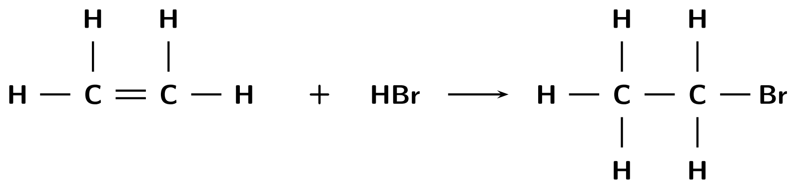


Figure 9.11: A substitution reaction

reaction to take place.

e.g. $CH_2{=}CH_2$ + HBr → CH_3-CH_2-Br (halo-alkane)

Halo-alkanes (also sometimes called *alkyl halides*) that contain methane and chlorine are substances that can be used as anaesthetics during operations. One example is trichloromethane, also known as 'chloroform' (figure 9.12).

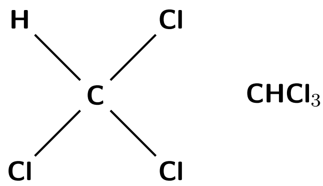


Figure 9.12: Trichloromethane

2. **Elimination reactions**

   Saturated compounds can also undergo elimination reactions to become unsaturated (figure 9.13). In the example below, an atom of hydrogen and chlorine are eliminated from the original compound to form an unsaturated halo-alkene.

   e.g. $CH_2Cl - CH_2Cl \rightarrow CH_2 = CHCl + HCl$

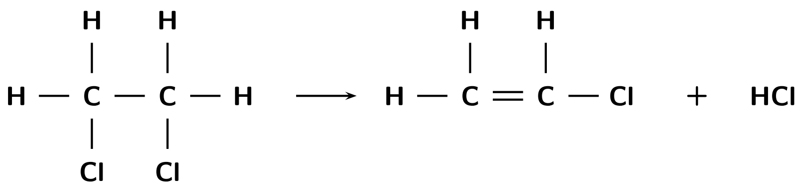


Figure 9.13: An elimination reaction

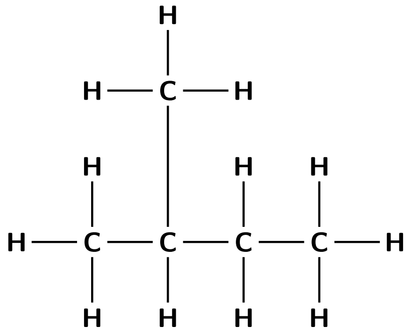
3. **Oxidation reactions**

   When alkanes are burnt in air, they react with the oxygen in air and heat is produced. This is called an oxidation or combustion reaction. Carbon dioxide and water are given off as products. Heat is also released during the reaction. The burning of alkanes provides most of the energy that is used by man.

   e.g. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + heat$

Exercise: The Alkanes

1. Give the IUPAC name for each of the following alkanes:
   (a) C_6H_{14}



   (b)
   (c) CH_3CH_3
2. Give the structural formula for each of the following compounds:
   (a) octane
   (b) 3-methyl-hexane
3. Methane is one of the simplest alkanes and yet it is an important fuel source. Methane occurs naturally in wetlands, natural gas and permafrost. However, methane can also be produced when organic wastes (e.g. animal manure and decaying material) are broken down by bacteria under conditions that are anaerobic (there is no oxygen). The simplified reaction is shown below:

$$\text{Organic matter} \rightarrow \text{Simple organic acids} \rightarrow \text{Biogas}$$

   The organic matter could be carbohydrates, proteins or fats which are broken down by acid-forming bacteria into simple organic acids such as acetic acid or formic acid. Methane-forming bacteria then convert these acids into biogases such as methane and ammonia.
   The production of methane in this way is very important because methane can be used as a fuel source. One of the advantages of methane over other fuels like coal, is that it produces more energy but with lower carbon dioxide emissions. The problem however, is that methane itself is a greenhouse gas and has a much higher global warming potential than carbon dioxide. So, producing methane may in fact have an even more dangerous impact on the environment.
   (a) What is the structural formula of methane?
   (b) Write an equation to show the reaction that takes place when methane is burned as a fuel.
   (c) Explain what is meant by the statement that methane 'has a greater global warming potential than carbon dioxide'.
4. Chlorine and ethane react to form chloroethane and hydrogen chloride.
   (a) Write a balanced chemical equation for this reaction, using molecular formulae.
   (b) Give the structural formula of chloroethane.
   (c) What type of reaction has taken place in this example?
5. Petrol (C_8H_{18}) is in fact not pure C_8H_{18} but a mixture of various *alkanes*. The 'octane rating' of petrol refers to the percentage of the petrol which is C_8H_{18}. For example, 93 octane fuel contains 93% C_8H_{18} and 7% other alkanes. The *isomer* of C_8H_{18} referred to in the 'octane rating' is in fact not octane but 2,2,4-trimethylpentane.
   (a) Write an unbalanced equation for the chemical reaction which takes place when petrol (C_8H_{18}) burns in excess oxygen.
   (b) Write the general formula of the *alkanes*.
   (c) Define the term *structural isomer*.
   (d) Use the information given in this question and your knowledge of naming organic compounds to deduce and draw the full structural formula for 2,2,4-trimethylpentane. (IEB pg 25)

9.7.5 The alkenes

In the alkenes, there is at least one double bond between two carbon atoms. This means that they are **unsaturated** and are *more reactive* than the alkanes. The simplest alkene is ethene (also known as ethylene), which is shown in figure 9.14.

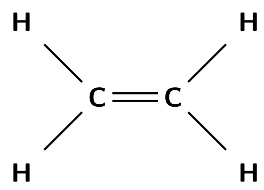
(a)  (b) CH_2CH_2 (c) C_2H_4

Figure 9.14: The (a) structural, (b) condensed structural and (c) molecular structure representations of ethene

As with the alkanes, the elkenes also form a homologous series. They have the general formula C_nH_{2n}. The second alkene in the series would therefore be C_3H_6. This molecule is known as propene (figure 9.15). Note that if an alkene has two double bonds, it is called a **diene** and if it has three double bonds it is called a **triene**.

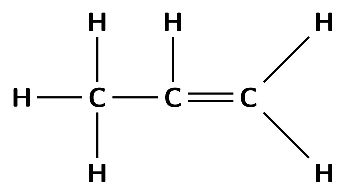
(a)  (b) CH_3CHCH_2 (c) C_3H_6

Figure 9.15: The (a) structural, (b) condensed structural and (c) molecular structure representations of propene

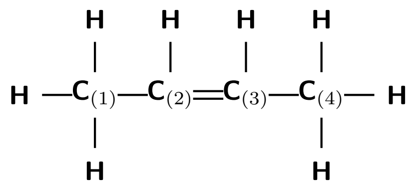
The elkenes have a variety of uses. Ethylene for example is a hormone in plants that stimulates the ripening of fruits and the opening of flowers. Propene is an important compound in the petrochemicals industry. It is used as a monomer to make polypropylene and is also used as a fuel gas for other industrial processes.

9.7.6 Naming the alkenes

Similar rules will apply in naming the alkenes, as for the alkanes.

Worked Example 42: Naming the alkenes

Question: Give the IUPAC name for the following compound:



Answer

Step 1 : Identify the functional group

The compound is an alkene and will have the suffix -ene.

Step 2 : Find the longest carbon chain

There are four carbon atoms in the longest chain and so the prefix for this compound will be 'but'.

Step 3 : Number the carbon atoms

Remember that when there is a double or triple bond, the carbon atoms must be numbered so that the double or triple bond is at the lowest numbered carbon. In this case, it doesn't matter whether we number the carbons from the left to right, or from the right to left. The double bond will still fall between C_2 and C_3. The position of the bond will come just before the suffix in the compound's name.

Step 4 : Look for any branched groups, name them and give their position on the carbon chain

There are no branched groups in this molecule.

Step 5 : Name the compound

The name of this compound is **but-2-ene**.

Worked Example 43: Naming the alkenes

Question: Draw the structural formula for the organic compound 3-methyl-butene

Answer

Step 1 : Identify the functional group

The suffix -ene means that this compound is an alkene and there must be a double bond in the molecule. There is no number immediately before the suffix which means that the double bond must be at the first carbon in the chain.

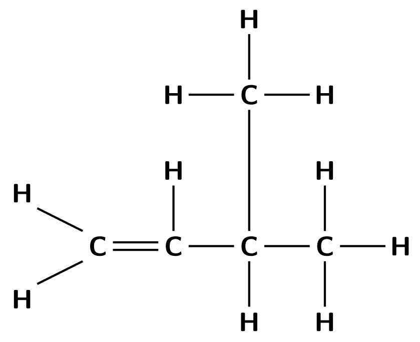
Step 2 : Determine the number of carbons in the longest chain

The prefix for the compound is 'but' so there must be four carbons in the longest chain.

Step 3 : Look for any branched groups

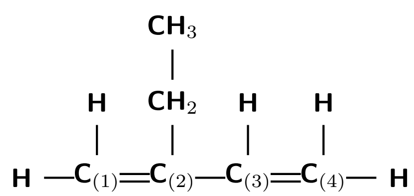
There is a methyl group at the third carbon atom in the chain.

Step 4 : Combine this information to draw the structural formula for this molecule.



Worked Example 44: Naming the alkenes

Question: Give the IUPAC name for the following compound:



Answer

Step 1 : Identify the functional group

The compound is an alkene and will have the suffix -ene. There is a double bond between the first and second carbons and also between the third and forth carbons. The organic compound is therefore a 'diene'.

Step 2 : Find the longest carbon chain and number the carbon atoms

There are four carbon atoms in the longest chain and so the prefix for this compound will be 'but'. The carbon atoms are numbered 1 to 4 in the diagram above. Remember that the main carbon chain must contain both the double bonds.

Step 3 : Look for any branched groups, name them and give their position on the carbon chain

There is a methyl group on the first carbon and an ethyl group on the second carbon.

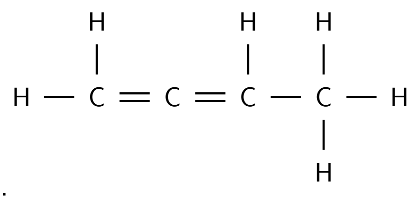
Step 4 : Name the compound

The name of this compound is **1-methyl,2-ethyl-1,3 diene**.

Exercise: Naming the alkenes

Give the IUPAC name for each of the following alkenes:

1. C_5H_{10}
2. $CH_3CHCHCH_3$

3. 

9.7.7 The properties of the alkenes

The properties of the alkenes are very similar to those of the alkanes, except that the alkenes are more reactive because they are unsaturated. As with the alkanes, compounds that have four or less carbon atoms are gases at room temperature, while those with five or more carbon atoms are liquids.

9.7.8 Reactions of the alkenes

Alkenes can undergo **addition reactions** because they are unsaturated. They readily react with hydrogen, water and the halogens. The double bond is broken and a single, saturated bond is formed. A new group is then added to one or both of the carbon atoms that previously made up the double bond. The following are some examples:

1. **Hydrogenation reactions**

   A catalyst such as platinum is normally needed for these reactions

   $CH_2 = CH_2 + H_2 \rightarrow CH_3 - CH_3$ (figure 9.16)

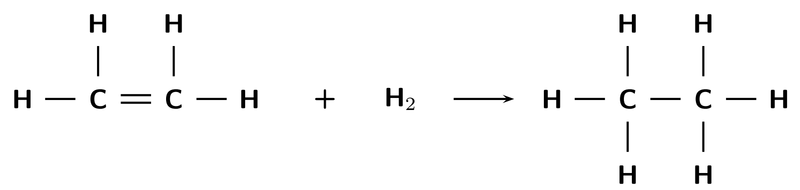


Figure 9.16: A hydrogenation reaction

2. **Halogenation reactions**

   $CH_2 = CH_2 + HBr \rightarrow CH_3 - CH_2 - Br$ (figure 9.17)

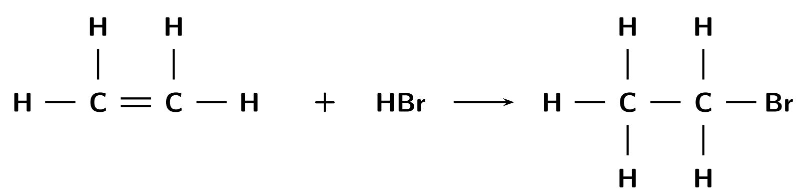


Figure 9.17: A halogenation reaction

3. **The formation of alcohols**

   $CH_2 = CH_2 + H_2O \rightarrow CH_3 - CH_2 - OH$ (figure 9.18)

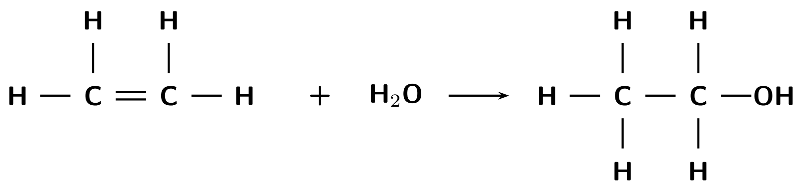
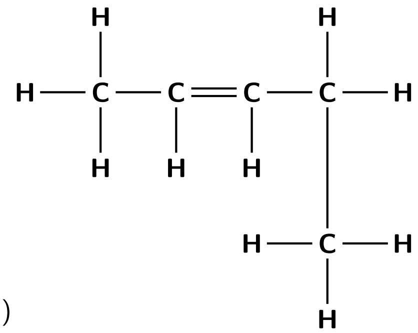


Figure 9.18: The formation of an alcohol

Exercise: The Alkenes

1. Give the IUPAC name for each of the following organic compounds:

   (a) 

   (b) CH_3CHCH_2

2. Refer to the data table below which shows the melting point and boiling point for a number of different organic compounds.

Formula	Name	Melting point (^{0}C)	Boiling point (^{0}C)
C_4H_{10}	Butane	-138	-0.5
C_5H_{12}	Pentane	-130	36
C_6H_{14}	Hexane	-95	69
C_4H_8	Butene	-185	-6
C_5H_{10}	Pentene	-138	30
C_6H_{12}	Hexene	-140	63

   (a) At room temperature (approx. 25^{0}C), which of the organic compounds in the table are:
      i. gases
      ii. liquids
   (b) In the alkanes...
      i. Describe what happens to the melting point and boiling point as the number of carbon atoms in the compound increases.
      ii. Explain why this is the case.
   (c) If you look at an alkane and an alkene that have the same number of carbon atoms...
      i. How do their melting points and boiling points compare?
      ii. Can you explain why their melting points and boiling points are different?
   (d) Which of the compounds, hexane or hexene, is more reactive? Explain your answer.

3. The following reaction takes place:

$$CH_3CHCH_2 + H_2 \rightarrow CH_3CH_2CH_3$$

   (a) Give the name of the organic compound in the reactants.
   (b) What is the name of the product?
   (c) What type of reaction is this?

(d) Which compound in the reaction is a saturated hydrocarbon?

9.7.9 The Alkynes

In the alkynes, there is at least one triple bond between two of the carbon atoms. They are unsaturated compounds and are therefore highly reactive. Their general formula is C_nH_{2n-2}. The simplest alkyne is ethyne (figure 9.19), also known as acetylene. Many of the alkynes are used to synthesise other chemical products.

$$H - C \equiv C - H$$

Figure 9.19: Ethyne (acetylene)

The raw materials that are needed to make acetylene are calcium carbonate and coal. Acetylene can be produced through the following reactions:

$$CaCO_3 \rightarrow CaO$$
$$CaO + 3C \rightarrow CaC_2 + CO$$
$$CaC_2 + 2H_2O \rightarrow Ca(OH)_2 + C_2H_2$$

An important use of acetylene is in oxyacetylene gas welding. The fuel gas burns with oxygen in a torch. An incredibly high heat is produced, and this is enough to melt metal.

9.7.10 Naming the alkynes

The same rules will apply as for the alkanes and alkenes, except that the suffix of the name will now be -yne.

Worked Example 45: Naming the alkynes

Question: Give the IUPAC name for the following compound:

$$\begin{array}{l} CH_3 - \underset{\displaystyle \overset{|}{CH_3}}{CH} - CH_2 - C \equiv C - CH_3 \end{array}$$

Answer
Step 1 : Identify the functional group

There is a triple bond between two of the carbon atoms, so this compound is an alkyne. The suffix will be -yne. The triple bond is at the second carbon, so the suffix will in fact be 2-yne.

Step 2 : Find the longest carbon chain and give the compound the correct prefix
If we count the carbons in a straight line, there are six. The prefix of the compound's name will be 'hex'.

Step 3 : Number the carbons in the longest chain
In this example, you will need to number the carbons from right to left so that the triple bond is between carbon atoms with the lowest numbers.

Step 4 : Look for any branched groups, name them and show the number of the carbon atom to which the group is attached
There is a methyl (CH_3) group attached to the fifth carbon (remember we have numbered the carbon atoms from right to left).

Step 5 : Combine the elements of the name into a single word in the following order: branched groups; prefix; name ending according to the functional group and its position along the longest carbon chain.
If we follow this order, the name of the compound is **5-methyl-hex-2-yne**.

Exercise: The alkynes
Give the IUPAC name for each of the following organic compounds.

1.
$$\begin{array}{ccccccccc} & & H & & CH_3 & & & & \\ & & | & & | & & & & \\ H & — & C & — & C & — & C \equiv C & — & H \\ & & | & & | & & & & \\ & & H & & H & & & & \end{array}$$

2. C_2H_2
3. CH_3CH_2CCH

9.8 The Alcohols

An alcohol is any organic compound where there is a *hydroxyl* functional group (-OH) bound to a carbon atom. The general formula for a simple alcohol is $C_nH_{2n+1}OH$.

The simplest and most commonly used alcohols are methanol and ethanol (figure 9.20).

The alcohols have a number of different uses:

- methylated spirits (surgical spirits) is a form of ethanol where methanol has been added
- ethanol is used in alcoholic drinks
- ethanol is used as an industrial solvent

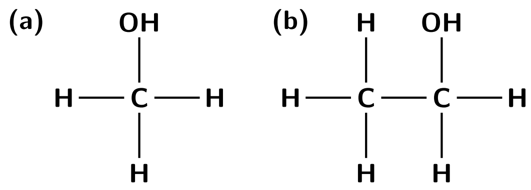


Figure 9.20: (a) methanol and (b) ethanol

- methanol and ethanol can both be used as a fuel and they burn more cleanly than gasoline or diesel (refer to chapter 21 for more information on biofuels as an alternative energy resource.)
- ethanol is used as a solvent in medical drugs, perfumes and vegetable essences
- ethanol is an antiseptic

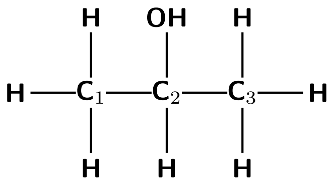
'Fermentation' refers to the conversion of sugar to alcohol using yeast (a fungus). The process of fermentation produces items such as wine, beer and yoghurt. To make wine, grape juice is fermented to produce alcohol. This reaction is shown below:

$$C_6H_{12}O_6 \rightarrow 2CO_2 + 2C_2H_5OH + energy$$

Ethanol is a diuretic. In humans, ethanol reduces the secretion of a hormone called antidiuretic hormone (ADH). The role of ADH is to control the amount of water that the body retains. When this hormone is not secreted in the right quantities, it can cause dehyration because too much water is lost from the body in the urine. This is why people who drink too much alcohol can become dehydrated, and experience symptoms such as headaches, dry mouth, and lethargy. Part of the reason for the headaches is that dehydration causes the brain to shrink away from the skull slightly. The effects of drinking too much alcohol can be reduced by drinking lots of water.

9.8.1 Naming the alcohols

The rules used to name the alcohols are similar to those already discussed for the other compounds, except that the suffix of the name will be different because the compound is an alcohol.



Worked Example 46: Naming alcohols 1

Question: Give the IUPAC name for the following organic compound

Answer

Step 1 : Identify the functional group

The compound has an -OH (hydroxyl) functional group and is therefore an alcohol. The compound will have the suffix -ol.

Step 2 : Find the longest carbon chain

There are three carbons in the longest chain. The prefix for this compound will be 'prop'. Since there are only single bonds between the carbon atoms, the suffix becomes 'propan' (similar to the alkane 'propane').

Step 3 : Number the carbons in the carbon chain

In this case, it doesn't matter whether you start numbering from the left or right. The hydroxyl group will still be attached to the middle carbon atom, numbered '2'.

Step 4 : Look for any branched groups, name them and give their position on the carbon chain.

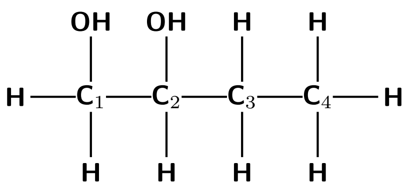
There are no branched groups in this compound, but you still need to indicate the position of the hydroxyl group on the second carbon. The suffix will be -2-ol because the hydroxyl group is attached to the second carbon.

Step 5 : Combine the elements of the compound's name into a single word in the order of branched groups; prefix; name ending according to the functional group.

The compound's name is propan-2-ol.

Worked Example 47: Naming alcohols 2

Question: Give the IUPAC name for the following compound:



Answer

Step 1 : Identify the functional group

The compound has an -OH (hydroxyl) functional group and is therefore an alcohol. There are *two* hydroxyl groups in the compound, so the suffix will be -diol.

Step 2 : Find the longest carbon chain

There are four carbons in the longest chain. The prefix for this compound will be 'butan'.

Step 3 : Number the carbons in the carbon chain

The carbons will be numbered from left to right so that the two hydroxyl groups are attached to carbon atoms with the lowest numbers.

Step 4 : Look for any branched groups, name them and give their position on the carbon chain.

There are no branched groups in this compound, but you still need to indicate the position of the hydroxyl groups on the first and second carbon atoms. The suffix will therefore become 1,2-diol.

Step 5 : Combine the elements of the compound's name into a single word in the order of branched groups; prefix; name ending according to the functional group.

The compound's name is butan-1,2-diol.

Exercise: Naming the alcohols

1. Give the structural formula of each of the following organic compounds:
   (a) pentan-3-ol
   (b) butan-2,3-diol
   (c) 2-methyl-propanol
2. Give the IUPAC name for each of the following:
   (a) $CH_3CH_2CH(OH)CH_3$
   (b)

$$\begin{array}{c}
 & & H & & & & & & \\
 & & | & & & & & & \\
CH_3 & — & C & — & CH_2 & — & CH_2 & — & CH_3 \\
 & & | & & & & & & \\
 & & OH & & & & & &
\end{array}$$

9.8.2 Physical and chemical properties of the alcohols

The hydroxyl group affects the **solubility** of the alcohols. The hydroxyl group generally makes the alcohol molecule *polar* and therefore more likely to be **soluble** in water. However, the carbon chain resists solubility, so there are two opposing trends in the alcohols. Alcohols with shorter carbon chains are usually more soluble than those with longer carbon chains.

Alcohols tend to have higher **boiling points** than the hydrocarbons because of the strong hydrogen bond between hydrogen and oxygen atoms.

Alcohols can show either **acidic** or **basic** properties because of the hydroxyl group. They also undergo oxidation reactions to form aldehydes, ketones and carboxylic acids.

Activity :: Case Study : The uses of the alcohols

Read the extract below and then answer the questions that follow:

The alcohols are a very important group of organic compounds, and they have a variety of uses. Our most common use of the word 'alcohol' is with reference to alcoholic drinks. The alcohol in drinks is in fact **ethanol**. But ethanol has many more uses apart from alcoholic drinks! When ethanol burns in air, it produces carbon dioxide, water and energy and can therefore be used as a fuel on its own, or in mixtures with petrol. Because ethanol can be produced through fermentation, this is a useful way for countries without an oil industry to reduce imports of petrol. Ethanol is also used as a solvent in many perfumes and cosmetics.

Methanol can also be used as a fuel, or as a petrol additive to improve combustion. Most methanol is used as an industrial feedstock, in other words it is used to make other things such as methanal (formaldehyde), ethanoic acid and methyl esters. In most cases, these are then turned into other products.

Propan-2-ol is another important alcohol, which is used in a variety of applications as a solvent.

Questions

1. Give the structural formula for propan-2-ol.
2. Write a balanced chemical equation for the combustion reaction of ethanol.
3. Explain why the alcohols are good solvents.

9.9 Carboxylic Acids

Carboxylic acids are organic acids that are characterised by having a carboxyl group, which has the formula -(C=O)-OH, or more commonly written as -COOH. In a carboxyl group, an oxygen atom is double-bonded to a carbon atom, which is also bonded to a hydroxyl group. The simplest carboxylic acid, methanoic acid, is shown in figure 9.21.

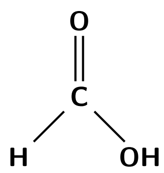


Figure 9.21: Methanoic acid

Carboxylic acids are widespread in nature. Methanoic acid (also known as *formic acid*) has the formula HCOOH and is found in insect stings. Ethanoic acid (CH_3COOH), or *acetic acid*, is the main component of vinegar. More complex organic acids also have a variety of different functions. Benzoic acid (C_6H_5COOH) for example, is used as a food preservative.

A certain type of ant, called formicine ants, manufacture and secrete formic acid, which is used to defend themselves against other organisms that might try to eat them.

9.9.1 Physical Properties

Carboxylic acids are **weak acids**, in other words they only dissociate partially. Why does the carboxyl group have acidic properties? In the carboxyl group, the hydrogen tends to separate itself from the oxygen atom. In other words, the carboxyl group becomes a source of positively-charged hydrogen ions (H^+). This is shown in figure 9.22.

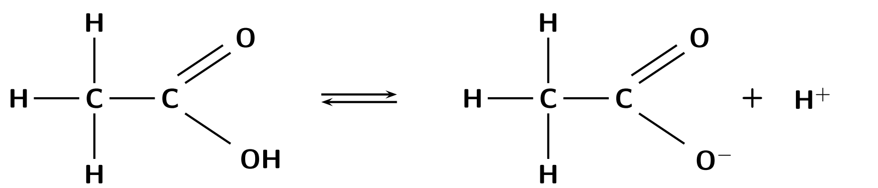


Figure 9.22: The dissociation of a carboxylic acid

?

Exercise: Carboxylic acids

1. Refer to the table below which gives information about a number of carboxylic acids, and then answer the questions that follow.

Formula	Common name	Source	IUPAC name	melting point (^{0}C)	boiling point (^{0}C)
	formic acid	ants	methanoic acid	8.4	101
CH_3CO_2H		vinegar	ethanoic acid	16.6	118
	propionic acid	milk	propanoic acid	-20.8	141
$CH_3(CH_2)_2CO_2H$	butyric acid	butter		-5.5	164
	valeric acid	valerian root	pentanoic acid	-34.5	186
$CH_3(CH_2)_4CO_2H$	caproic acid	goats		-4	205
	enanthic acid	vines		-7.5	223
$CH_3(CH_2)_6CO_2H$	caprylic acid	goats		16.3	239

   (a) Fill in the missing spaces in the table by writing the formula, common name or IUPAC name.
   (b) Draw the structural formula for butyric acid.
   (c) Give the molecular formula for caprylic acid.
   (d) Draw a graph to show the relationship between molecular mass (on the x-axis) and boiling point (on the y-axis)

i. Describe the trend you see.
ii. Suggest a reason for this trend.

9.9.2 Derivatives of carboxylic acids: The esters

When an alcohol reacts with a carboxylic acid, an **ester** is formed. Most esters have a characteristic and pleasant smell. In the reaction, the hydrogen atom from the hydroxyl group, and an OH from the carboxlic acid, form a molecule of water. A new bond is formed between what remains of the alcohol and acid. The name of the ester is a combination of the names of the alcohol and carboxylic acid. The suffix for an ester is -oate. An example is shown in figure 9.23.

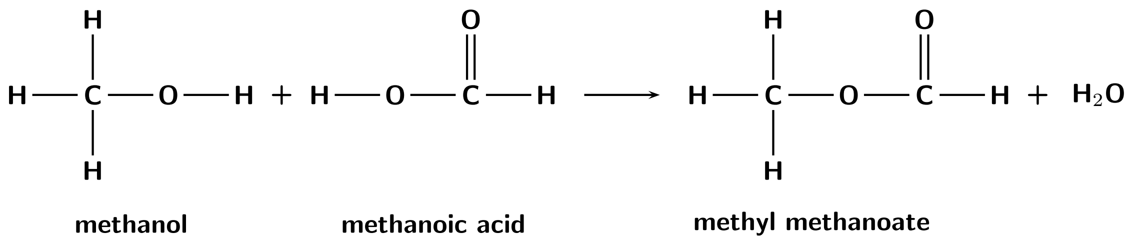


Figure 9.23: The formation of an ester from an alcohol and carboxylic acid

9.10 The Amino Group

The amino group has the formula -NH_2 and consists of a nitrogen atom that is bonded to two hydrogen atoms, and to the carbon skeleton. Organic compounds that contain this functional group are called **amines**. One example is *glycine*. Glycine belongs to a group of organic compounds called *amino acids*, which are the building blocks of proteins.

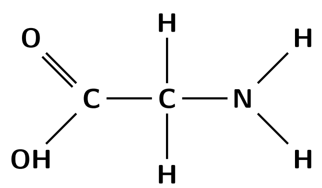


Figure 9.24: A molecule of glycine

9.11 The Carbonyl Group

The carbonyl group (-CO) consists of a carbon atom that is joined to an oxygen by a double bond. If the functional group is on the *end* of the carbon chain, the organic compound is called a **ketone**. The simplest ketone is *acetone*, which contains three carbon atoms. A ketone has the ending 'one' in its IUPAC name.

?

Exercise: Carboxylic acids, esters, amines and ketones

1. Look at the list of organic compounds in the table below:

Organic compound	Type of compound
$CH_3CH_2CH_2COOH$	
NH_2CH_2COOH	
propyl ethanoate	
CH_3CHO	

   (a) Complete the table by identifying each compound as either a carboxylic acid, ester, amine or ketone.
   (b) Give the name of the compounds that have been written as condensed structural formulae.

2. A chemical reaction takes place and ethyl methanoate is formed.
   (a) What type of organic compound is ethyl methanoate?
   (b) Name the two reactants in this chemical reaction.
   (c) Give the structural formula of ethyl methanoate.

9.12 Summary

- **Organic chemistry** is the branch of chemistry that deals with organic molecules. An **organic molecule** is one that contains carbon.
- All **living organisms** contain carbon. Plants use sunlight to convert carbon dioxide in the air into organic compounds through the process of **photosynthesis**. Animals and other organisms then feed on plants to obtain their own organic compounds. **Fossil fuels** are another important source of carbon.
- It is the **unique properties of the carbon atom** that give organic compounds certain properties.
- The carbon atom has **four valence electrons**, so it can bond with many other atoms, often resulting in long chain structures. It also forms mostly **covalent bonds** with the atoms that it bonds to, meaning that most organic molecules are **non-polar**.
- An organic compound can be represented in different ways, using its **molecular formula**, **structural formula** or **condensed structural formula**.
- If two compounds are **isomers**, it means that they have the same molecular formulae but different structural formulae.
- A **functional group** is a particular group of atoms within a molecule, which give it certain reaction characteristics. Organic compounds can be grouped according to their functional group.
- The **hydrocarbons** are organic compounds that contain only carbon and hydrogen. They can be further divided into the alkanes, alkenes and alkynes, based on the type of bonds between the carbon atoms.
- The **alkanes** have only **single bonds** between their carbon atoms and are unreactive.
- The **alkenes** have at least one **double bond** between two of their carbon atoms. They are more reactive than the alkanes.
- The **alkynes** have at least one **triple bond** between two of their carbon atoms. They are the most reactive of the three groups.
- A hydrocarbon is said to be **saturated** if it contains the maximum possible number of hydrogen atoms for that molecule. The alkanes are all saturated compounds.

- A hydrocarbon is **unsaturated** if it does not contain the maximum number of hydrogen atoms for that molecule. The alkenes and alkynes are examples of unsaturated molecules. If a double or triple bond is broken, more hydrogen atoms can be added to the molecule.
- There are three types of reactions that occur in the alkanes: **substitution**, **elimination** and **oxidation** reactions.
- The alkenes undergo **addition** reactions because they are unsaturated.
- Organic compounds are **named** according to their functional group and its position in the molecule, the number of carbon atoms in the molecule and the position of any double and triple bonds. The IUPAC rules for nomenclature are used in the naming of organic molecules.
- Many of the **properties** of the hydrocarbons are determined by their **molecular structure**, the **bonds** between atoms and molecules, and their **surface area**.
- The **melting point** and **boiling point** of the hydrocarbons increases as their number of carbon atoms increases.
- The **molecular mass** of the hydrocarbons determines whether they will be in the gaseous, liquid or solid phase at certain temperatures.
- An **alcohol** is an organic compound that contains a **hydroxyl group** (OH).
- The alcohols have a number of different uses including their use as a solvent, for medicinal purposes and in alcoholic drinks.
- The alcohols share a number of **properties** because of the hydroxyl group. The hydroxyl group affects the **solubility** of the alcohols. Those with shorter carbon chains are generally more soluble, and those with longer chains are less soluble. The strong hydrogen bond between the hydrogen and oxygen atoms in the hydroxyl group gives alcohols a higher melting point and boiling point than other organic compounds. The hydroxyl group also gives the alcohols both acidic and basic properties.
- The **carboxylic acids** are organic acids that contain a **carboxyl group** with the formula COOH. In a carboxyl group, an oxygen atom is double-bonded to a carbon atom, which is also bonded to a hydroxyl group.
- The carboxylic acids have weak **acidic properties** because the hydrogen atom is able to dissociate from the carboxyl group.
- An **ester** is formed when an alcohol reacts with a carboxylic acid.
- The **amines** are organic compounds that contain an **amino** functional group, which has the formula NH_2. Some amines belong to the **amino acid** group, which are the building blocks of proteins.
- The **ketones** are a group of compounds that contain a **carbonyl group**, which consists of an oxygen atom that is double-bonded to a carbon atom. In a ketone, the carbonyl group is on the end of the carbon chain.

?

Exercise: Summary exercise

1. Give **one word** for each of the following descriptions:
   (a) The group of hydrocarbons to which 2-methyl-propene belongs.
   (b) The name of the functional group that gives alcohols their properties.
   (c) The group of organic compounds that have acidic properties.
   (d) The name of the organic compound that is found in vinegar.
   (e) The name of the organic compound that is found in alcoholic beverages.

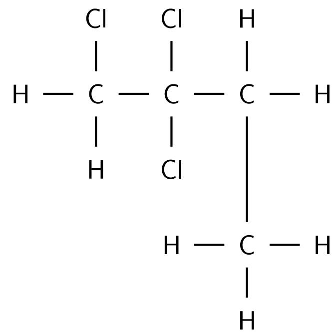
2. In each of the following questions, choose the **one correct answer** from the list provided.

   (a) When 1-propanol is oxidised by acidified potassium permanganate, the possible product formed is...

      i. propane
      ii. propanoic acid
      iii. methyl propanol
      iv. propyl methanoate

      (IEB 2004)

   (b) What is the IUPAC name for the compound represented by the following structural formula?



      i. 1,2,2-trichlorobutane
      ii. 1-chloro-2,2-dichlorobutane
      iii. 1,2,2-trichloro-3-methylpropane
      iv. 1-chloro-2,2-dichloro-3-methylpropane

      (IEB 2003)

3. Write balanced equations for the following reactions:

   (a) Ethene reacts with bromine
   (b) Ethyne gas burns in an excess of oxygen
   (c) Ethanoic acid ionises in water

4. The table below gives the boiling point of ten organic compounds.

	Compound	Formula	Boiling Point (^{0}C)
1	methane	CH_4	-164
2	ethane	C_2H_6	-88
3	propane	C_3H_8	-42
4	butane	C_4H_{10}	0
5	pentane	C_5H_{12}	36
6	methanol	CH_3OH	65
7	ethanol	C_2H_5OH	78
8	propan-1-ol	C_3H_7OH	98
9	propan-1,2-diol	$CH_3CHOHCH_2OH$	189
10	propan-1,2,3-triol	$CH_2OHCHOHCH_2OH$	290

   The following questions refer to the compounds shown in the above table.

   (a) To which homologous series do the following compounds belong?

      i. Compounds 1,2 and 3
      ii. Compounds 6,7 and 8

   (b) Which of the above compounds are gases at room temperature?
   (c) What causes the trend of increasing boiling points of compounds 1 to 5?
   (d) Despite the fact that the length of the carbon chain in compounds 8,9 and 10 is the same, the boiling point of propan-1,2,3-triol is much higher than the boiling point of propan-1-ol. What is responsible for this large difference in boiling point?
   (e) Give the IUPAC name and the structural formula of an isomer of butane.

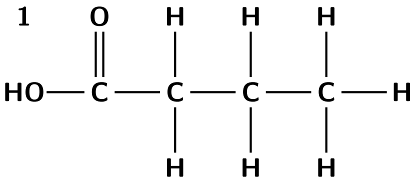
(f) Which **one** of the above substances is used as a reactant in the preparation of the ester ethylmethanoate?

(g) Using structural formulae, write an equation for the reaction which produces ethylmethanoate.

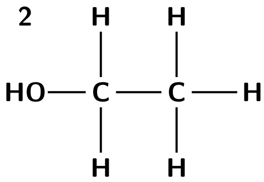
(*IEB 2004*)

5. Refer to the numbered diagrams below and then answer the questions that follow.

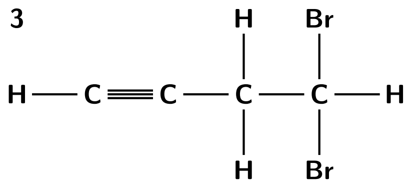
1



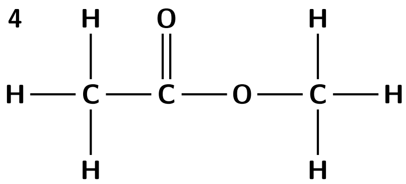
2



3



4



(a) Which one of the above compounds is produced from the fermentation of starches and sugars in plant matter?
   i. compound 1
   ii. compound 2
   iii. compound 3
   iv. compound 4

(b) To which one of the following homologous series does compound 1 belong?
   i. esters
   ii. alcohols
   iii. aldehydes
   iv. carboxylic acids

(c) The correct IUPAC name for compound 3 is...
   i. 1,1-dibromo-3-butyne
   ii. 4,4-dibromo-1-butyne
   iii. 2,4-dibromo-1-butyne
   iv. 4,4-dibromo-1-propyne

(d) What is the correct IUPAC name for compound 4?
   i. propanoic acid
   ii. ethylmethanoate
   iii. methylethanoate
   iv. methylpropanoate

IEB 2005

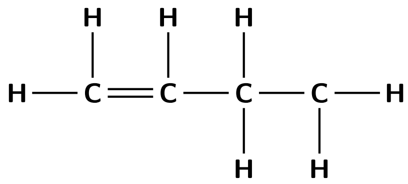
6. Answer the following questions:

(a) What is a homologous series?

(b) A mixture of ethanoic acid and methanol is warmed in the presence of concentrated sulphuric acid.
   i. Using structural formulae, give an equation for the reaction which takes place.

ii. What is the IUPAC name of the organic compound formed in this reaction?

(c) Consider the following unsaturated hydrocarbon:



i. Give the IUPAC name for this compound.

ii. Give the balanced equation for the combustion of this compound in excess oxygen.

(IEB Paper 2, 2003)

7. Consider the organic compounds labelled A to E.

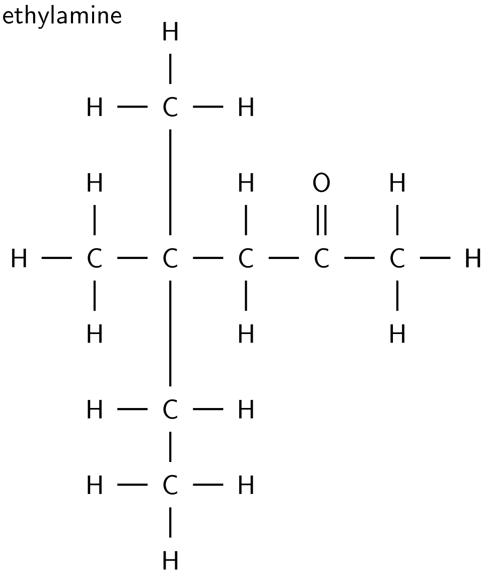
A. $CH_3CH_2CH_2CH_2CH_2CH_3$

B. C_6H_6

C. CH_3-Cl

D. Methylamine

E



(a) Write a balanced chemical equation for the preparation of compound C using an alkane as one of the reactants.

(b) Write down the IUPAC name for compound E.

(c) Write down the structural formula of an isomer of compound A that has only FOUR carbon atoms in the longest chain.

(d) Write down the structural formula for compound B.

Chapter 10

Organic Macromolecules - Grade 12

As its name suggests, a macromolecule is a large molecule that forms when lots of smaller molecules are joined together. In this chapter, we will be taking a closer look at the structure and properties of different macromolecules, and at how they form.

10.1 Polymers

Some macromolecules are made up of lots of repeating structural units called **monomers**. To put it more simply, a monomer is like a building block. When lots of similar monomers are joined together by covalent bonds, they form a **polymer**. In an **organic polymer**, the monomers would be joined by the *carbon* atoms of the polymer 'backbone'. A polymer can also be **inorganic**, in which case there may be atoms such as *silicon* in the place of carbon atoms. The key feature that makes a polymer different from other macromolecules, is the repetition of identical or similar monomers in the polymer chain. The examples shown below will help to make these concepts clearer.

Definition: Polymer
Polymer is a term used to describe large molecules consisting of repeating structural units, or monomers, connected by covalent chemical bonds.

1. **Polyethene**

   Chapter 9 looked at the structure of a group of hydrocarbons called the *alkenes*. One example is the molecule **ethene**. The structural formula of ethene is is shown in figure 10.1. When lots of ethene molecules bond together, a polymer called **polyethene** is formed. Ethene is the *monomer* which, when joined to other ethene molecules, forms the *polymer* **polyethene**. Polyethene is a cheap plastic that is used to make plastic bags and bottles.

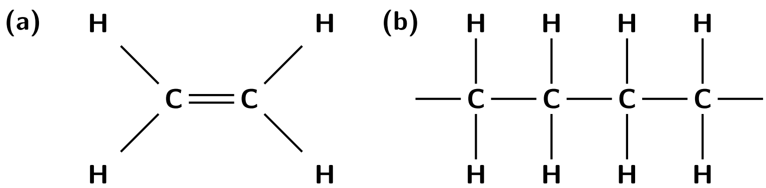


Figure 10.1: (a) Ethene monomer and (b) polyethene polymer

A polymer may be a chain of thousands of monomers, and so it is impossible to draw the entire polymer. Rather, the structure of a polymer can be condensed and represented as shown in figure 10.2. The monomer is enclosed in brackets and the 'n' represents the number of ethene molecules in the polymer, where 'n' is any whole number. What this shows is that the ethene monomer is repeated an indefinite number of times in a molecule of polyethene.

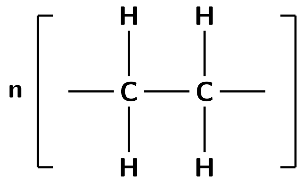


Figure 10.2: A simplified representation of a polyethene molecule

2. **Polypropene**

   Another example of a polymer is *polypropene* (fig 10.3). Polypropene is also a plastic, but is stronger than polyethene and is used to make crates, fibres and ropes. In this polymer, the monomer is the alkene called **propene**.

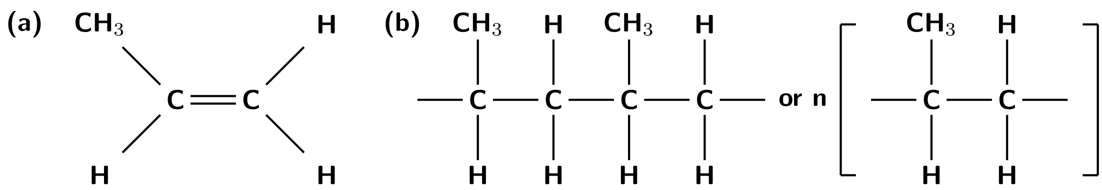


Figure 10.3: (a) Propene monomer and (b) polypropene polymer

10.2 How do polymers form?

Polymers are formed through a process called **polymerisation**, where monomer molecules react together to form a polymer chain. Two types of polymerisation reactions are **addition polymerisation** and **condensation polymerisation**.

Definition: Polymerisation
In chemistry, polymerisation is a process of bonding monomers, or *single units* together through a variety of reaction mechanisms to form longer chains called polymers.

10.2.1 Addition polymerisation

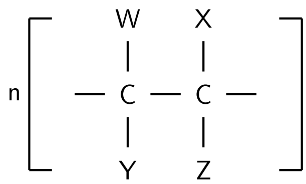
In this type of reaction, monomer molecules are added to a growing polymer chain one at a time. No small molecules are eliminated in the process. An example of this type of reaction is the formation of *polyethene* from *ethene* (fig 10.1). When molecules of ethene are joined to each other, the only thing that changes is that the double bond between the carbon atoms in each ethene monomer is replaced by a single bond so that a new carbon-carbon bond can be formed with the next monomer in the chain. In other words, the monomer is an *unsaturated* compound which, after an addition reaction, becomes a *saturated* compound.

Extension: Initiation, propagation and termination

There are three stages in the process of addition polymerisation. **Initiation** refers to a chemical reaction that triggers off another reaction. In other words, initiation is the starting point of the polymerisation reaction. **Chain propagation** is the part where monomers are continually added to form a longer and longer polymer chain. During chain propagation, it is the reactive end groups of the polymer chain that react in each propagation step, to add a new monomer to the chain. Once a monomer has been added, the reactive part of the polymer is now in this last monomer unit so that propagation will continue. **Termination** refers to a chemical reaction that destroys the reactive part of the polymer chain so that propagation stops.

Worked Example 48: Polymerisation reactions

Question: A polymerisation reaction takes place and the following polymer is formed:



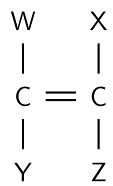
Note: W, X, Y and Z could represent a number of different atoms or combinations of atoms e.g. H, F, Cl or CH_3.

1. Give the structural formula of the monomer of this polymer.
2. To what group of organic compounds does this monomer belong?
3. What type of polymerisation reaction has taken place to join these monomers to form the polymer?

Answer

Step 1 : Look at the structure of the repeating unit in the polymer to determine the monomer.

The monomer is:



Step 2 : Look at the atoms and bonds in the monomer to determine which group of organic compounds it belongs to.

The monomer has a double bond between two carbon atoms. The monomer must be an alkene.

Step 3 : Determine the type of polymerisation reaction.

In this example, unsaturated monomers combine to form a saturated polymer. No atoms are lost or gained for the bonds between monomers to form. They are simply added to each other. This is an addition reaction.

10.2.2 Condensation polymerisation

In this type of reaction, two monomer molecules form a covalent bond and a small molecule such as water is lost in the bonding process. Nearly all biological reactions are of this type. **Polyester** and **nylon** are examples of polymers that form through condensation polymerisation.

1. **Polyester**

   Polyesters are a group of polymers that contain the **ester** functional group in their main chain. Although there are many forms of polyesters, the term *polyester* usually refers to polyethylene terephthalate (PET). PET is made from ethylene glycol (an alcohol) and terephthalic acid (an acid). In the reaction, a hydrogen atom is lost from the alcohol, and a hydroxyl group is lost from the carboxylic acid. Together these form one water molecule which is lost during condensation reactions. A new bond is formed between an oxygen and a carbon atom. This bond is called an **ester linkage**. The reaction is shown in figure 10.4.

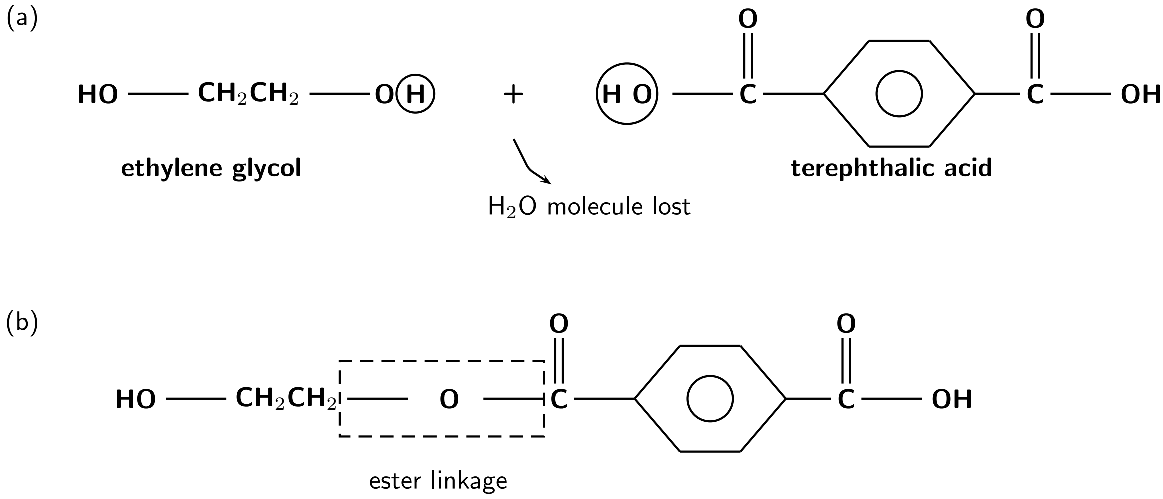


Figure 10.4: An acid and an alcohol monomer react (a) to form a molecule of the polyester 'polyethylene terephthalate' (b).

   Polyesters have a number of characteristics which make them very useful. They are resistant to stretching and shrinking, they are easily washed and dry quickly, and they are resistant to mildew. It is for these reasons that polyesters are being used more and more in **textiles**. Polyesters are stretched out into fibres and can then be made into fabric and articles of clothing. In the home, polyesters are used to make clothing, carpets, curtains, sheets, pillows and upholstery.

Interesting Fact

Polyester is not just a textile. Polyethylene terephthalate is in fact a plastic which can also be used to make plastic drink bottles. Many drink bottles are recycled by being reheated and turned into polyester fibres. This type of recycling helps to reduce disposal problems.

2. **Nylon**

   Nylon was the first polymer to be commercially successful. Nylon replaced silk, and was used to make parachutes during World War 2. Nylon is very strong and resistant, and is used in fishing line, shoes, toothbrush bristles, guitar strings and machine parts to name just a few. Nylon is formed from the reaction of an amine (1,6-diaminohexane) and an acid monomer (adipic acid) (figure 10.5). The bond that forms between the two monomers is called an **amide linkage**. An amide linkage forms between a nitrogen atom in the amine monomer and the carbonyl group in the carboxylic acid.

(a)
(b)

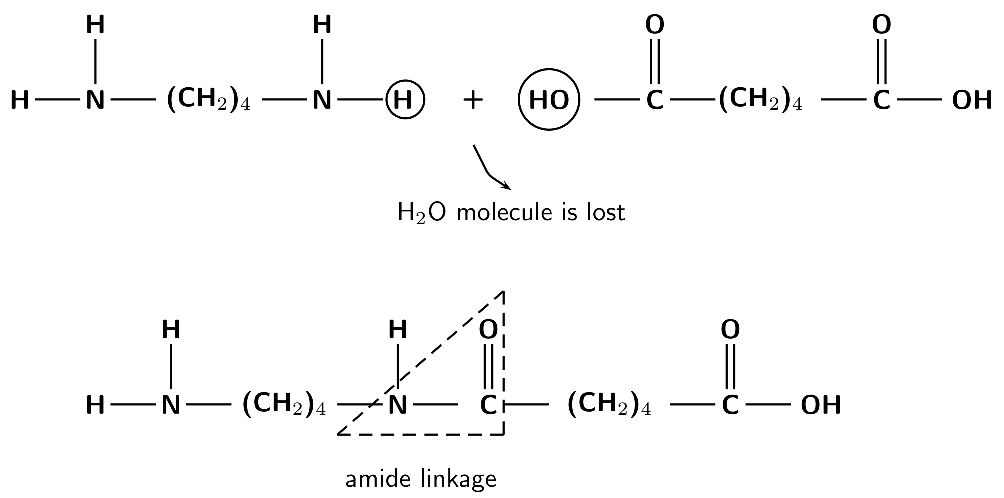


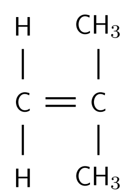
Figure 10.5: An amine and an acid monomer (a) combine to form a section of a nylon polymer (b).

Interesting Fact

Nylon was first introduced around 1939 and was in high demand to make stockings. However, as World War 2 progressed, nylon was used more and more to make parachutes, and so stockings became more difficult to buy. After the war, when manufacturers were able to shift their focus from parachutes back to stockings, a number of riots took place as women queued to get stockings. In one of the worst disturbances, 40 000 women queued up for 13 000 pairs of stockings, which led to fights breaking out!

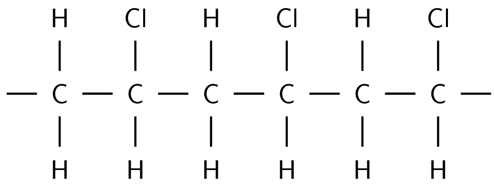
Exercise: Polymers

1. The following monomer is a reactant in a polymerisation reaction:



   (a) What is the IUPAC name of this monomer?
   (b) Give the structural formula of the polymer that is formed in this polymerisation reaction.
   (c) Is the reaction an addition or condensation reaction?

2. The polymer below is the product of a polymerisation reaction.



   (a) Give the structural formula of the monomer in this polymer.
   (b) What is the name of the monomer?
   (c) Draw the abbreviated structural formula for the polymer.
   (d) Has this polymer been formed through an addition or condensation polymerisation reaction?

3. A condensation reaction takes place between methanol and methanoic acid.
   (a) Give the structural formula for...
      i. methanol
      ii. methanoic acid
      iii. the product of the reaction
   (b) What is the name of the product? (Hint: The product is an ester)

10.3 The chemical properties of polymers

The attractive forces between polymer chains play a large part in determining a polymer's properties. Because polymer chains are so long, these interchain forces are very important. It is usually the side groups on the polymer that determine what types of intermolecular forces will exist. The greater the strength of the intermolecular forces, the greater will be the tensile strength and melting point of the polymer. Below are some examples:

- **Hydrogen bonds between adjacent chains**

  Polymers that contain amide or carbonyl groups can form *hydrogen bonds* between adjacent chains. The positive hydrogen atoms in the N-H groups of one chain are strongly attracted to the oxygen atoms in the C=O groups on another. Polymers that contain *urea* linkages would fall into this category. The structural formula for urea is shown in figure 10.6. Polymers that contain urea linkages have high tensile strength and a high melting point.

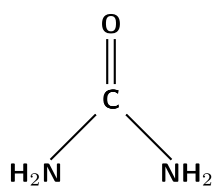


Figure 10.6: The structural formula for urea

- **Dipole-dipole bonds between adjacent chains**

  Polyesters have *dipole-dipole bonding* between their polymer chains. Dipole bonding is not as strong as hydrogen bonding, so a polyester's melting point and strength are lower than those of the polymers where there are hydrogen bonds between the chains. However, the weaker bonds between the chains means that polyesters have greater flexibility. The greater the flexibility of a polymer, the more likely it is to be moulded or stretched into fibres.

- **Weak van der Waal's forces**

  Other molecules such as ethene do not have a permanent dipole and so the attractive forces between polyethene chains arise from weak *van der Waals* forces. Polyethene therefore has a lower melting point than many other polymers.

10.4 Types of polymers

There are many different types of polymers. Some are organic, while others are inorganic. Organic polymers can be broadly grouped into either synthetic/semi-synthetic (artificial) or biological (natural) polymers. We are going to take a look at two groups of organic polymers: *plastics*, which are usually synthetic or semi-synthetic and *biological macromolecules* which are natural polymers. Both of these groups of polymers play a very important role in our lives.

10.5 Plastics

In today's world, we can hardly imagine life without plastic. From cellphones to food packaging, fishing line to plumbing pipes, compact discs to electronic equipment, plastics have become a very important part of our daily lives. "Plastics" cover a range of synthetic and semi-synthetic organic polymers. Their name comes from the fact that they are 'malleable', in other words their shape can be changed and moulded.

Definition: Plastic
Plastic covers a range of synthetic or semisynthetic organic polymers. Plastics may contain other substances to improve their performance. Their name comes from the fact that many of them are malleable, in other words they have the property of plasticity.

It was only in the nineteenth century that it was discovered that plastics could be made by chemically changing natural polymers. For centuries before this, only natural organic polymers had been used. Examples of natural organic polymers include *waxes* from plants, *cellulose* (a plant polymer used in fibres and ropes) and *natural rubber* from rubber trees. But in many cases, these natural organic polymers didn't have the characteristics that were needed for them to be used in specific ways. Natural rubber for example, is sensitive to temperature and becomes sticky and smelly in hot weather and brittle in cold weather.

In 1834 two inventors, Friedrich Ludersdorf of Germany and Nathaniel Hayward of the US, independently discovered that adding sulfur to raw rubber helped to stop the material from becoming sticky. After this, Charles Goodyear discovered that heating this modified rubber made it more resistant to abrasion, more elastic and much less sensitive to temperature. What these inventors had done was to improve the properties of a natural polymer so that it could be used in new ways. An important use of rubber now is in vehicle tyres, where these properties of rubber are critically important.

Interesting Fact

The first true plastic (i.e. one that was not based on any material found in nature) was *Bakelite*, a cheap, strong and durable plastic. Some of these plastics are still used for example in electronic circuit boards, where their properties of insulation and heat resistance are very important.

10.5.1 The uses of plastics

There is such a variety of different plastics available, each having their own specific properties and uses. The following are just a few examples.

- **Polystyrene**

  Polystyrene (figure 15.2) is a common plastic that is used in model kits, disposable eating utensils and a variety of other products. In the polystyrene polymer, the monomer is *styrene*, a liquid hydrocarbon that is manufactured from petroleum.

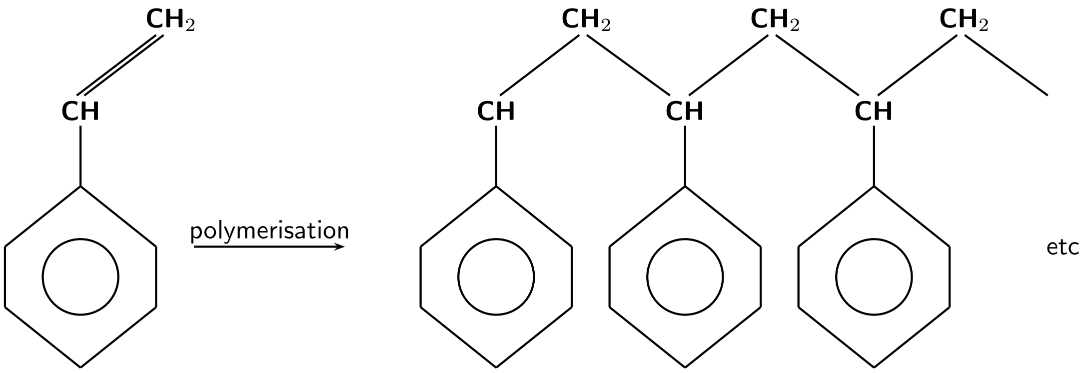


Figure 10.7: The polymerisation of a styrene monomer to form a polystyrene polymer

- **Polyvinylchloride (PVC)**

  Polyvinyl chloride (PVC) (figure 10.8) is used in plumbing, gutters, electronic equipment, wires and food packaging. The side chains of PVC contain chlorine atoms, which give it its particular characteristics.

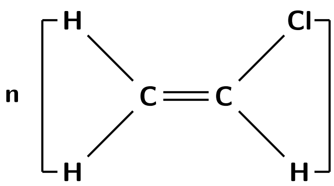


Figure 10.8: Polyvinyl chloride

Interesting Fact

Many vinyl products have other chemicals added to them to give them particular properties. Some of these chemicals, called additives, can leach out of the vinyl products. In PVC, *plasticizers* are used to make PVC more flexible. Because many baby toys are made from PVC, there is concern that some of these products may leach into the mouths of the babies that are chewing on them. In the USA, most companies have stopped making PVC toys. There are also concerns that some of the plasticizers added to PVC may cause a number of health conditions including cancer.

- **Synthetic rubber**

  Another plastic that was critical to the World War 2 effort was *synthetic rubber*, which was produced in a variety of forms. Not only were worldwide natural rubber supplies limited, but most rubber-producing areas were under Japanese control. Rubber was needed for tyres and parts of war machinery. After the war, synthetic rubber also played an important part in the space race and nuclear arms race.

- **Polyethene/polyethylene (PE)**

  Polyethylene (figure 10.1) was discovered in 1933. It is a cheap, flexible and durable plastic and is used to make films and packaging materials, containers and car fittings. One of the most well known polyethylene products is 'Tupperware', the sealable food containers designed by Earl Tupper and promoted through a network of housewives!

- **Polytetrafluoroethylene (PTFE)**

  Polytetrafluoroethylene (figure 10.9) is more commonly known as 'Teflon' and is most well known for its use in non-stick frying pans. Teflon is also used to make the breathable fabric Gore-Tex.

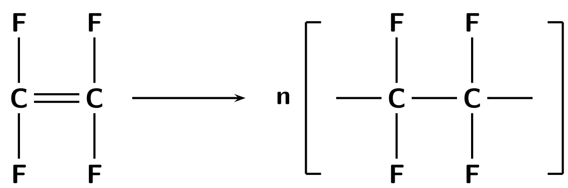


Figure 10.9: A tetra fluoroethylene monomer and polytetrafluoroethylene polymer

Table 10.1 summarises the formulae, properties and uses of some of the most common plastics.

Table 10.1: A summary of the formulae, properties and uses of some common plastics

Name	**Formula**	**Monomer**	**Properties**	**Uses**
Polyethene (low density)	-$(CH_2\text{-}CH_2)_n$-	$CH_2{=}CH_2$	soft, waxy solid	film wrap and plastic bags
Polyethene (high density)	-$(CH_2\text{-}CH_2)_n$-	$CH_2{=}CH_2$	rigid	electrical insulation, bottles and toys
Polypropene	-$[CH_2\text{-}CH(CH_3)]_n$-	$CH_2{=}CHCH_3$	different grades: some are soft and others hard	carpets and upholstery
Polyvinylchloride (PVC)	-$(CH_2\text{-}CHCl)_n$-	$CH_2{=}CHCl$	strong, rigid	pipes, flooring
Polystyrene	-$[CH_2\text{-}CH(C_6H_5)]_n$	$CH_2{=}CHC_6H_5$	hard, rigid	toys, packaging
Polytetrafluoroethylene	-$(CF_2\text{-}CF_2)_n$-	$CF_2{=}CF_2$	resistant, smooth, solid	non-stick surfaces, electrical insulation

?

Exercise: Plastics

1. It is possible for macromolecules to be composed of more than one type of repeating monomer. The resulting polymer is called a **copolymer**. Varying the monomers that combine to form a polymer, is one way of controlling the properties of the resulting material. Refer to the table below which shows a number of different copolymers of rubber, and answer the questions that follow:

Monomer A	Monomer B	Copolymer	Uses
H_2C=CHCl	H_2C=CCl_2	Saran	films and fibres
H_2C=CHC_6H_5	H_2C=C-CH=CH_2	SBR (styrene butadiene rubber)	tyres
H_2C=CHCN	H_2C=C-CH=CH_2	Nitrile rubber	adhesives and hoses
H_2C=$C(CH_3)_2$	H_2C=C-CH=CH_2	Butyl rubber	inner tubes
F_2C=$CF(CF_3)$	H_2C=CHF	Viton	gaskets

(a) Give the structural formula for each of the monomers of nitrile rubber.
(b) Give the structural formula of the copolymer viton.
(c) In what ways would you expect the properties of SBR to be different from nitrile rubber?
(d) Suggest a reason why the properties of these polymers are different.

2. In your home, find as many examples of different types of plastics that you can. Bring them to school and show them to your group. Together, use your examples to complete the following table:

Object	Type of plastic	Properties	Uses

10.5.2 Thermoplastics and thermosetting plastics

A **thermoplastic** is a plastic that can be melted to a liquid when it is heated and freezes to a brittle, glassy state when it is cooled enough. These properties of thermoplastics are mostly due to the fact that the forces between chains are weak. This also means that these plastics can be easily stretched or moulded into any shape. Examples of thermoplastics include nylon, polystyrene, polyethylene, polypropylene and PVC. Thermoplastics are more easily recyclable than some other plastics.

Thermosetting plastics differ from thermoplastics because once they have been formed, they cannot be remelted or remoulded. Examples include bakelite, vulcanised rubber, melanine (used to make furniture), and many glues. Thermosetting plastics are generally stronger than thermoplastics and are better suited to being used in situations where there are high temperatures. They are not able to be recycled. Thermosetting plastics have strong covalent bonds between chains and this makes them very strong.

Activity :: Case Study : Biodegradable plastics
Read the article below and then answer the questions that follow.

Our whole world seems to be wrapped in plastic. Almost every product we buy, most of the food we eat and many of the liquids we drink come encased in plastic. Plastic packaging provides excellent protection for the product, it is cheap to manufacture and seems to last forever. Lasting forever, however, is proving to be a major environmental problem. Another problem is that traditional plastics are manufactured from non-renewable resources - oil, coal and natural gas. In an effort to overcome these problems, researchers and engineers have been trying to develop biodegradable plastics that are made from renewable resources, such as plants.

The term biodegradable means that a substance can be broken down into simpler substances by the activities of living organisms, and therefore is unlikely to remain in the environment. The reason most plastics are not biodegradable is because their long polymer molecules are too large and too tightly bonded together to be broken apart and used by decomposer organisms. However, plastics based on natural plant polymers that come from wheat or corn starch have molecules that can be more easily broken down by microbes.

Starch is a natural polymer. It is a white, granular carbohydrate produced by plants during photosynthesis and it serves as the plant's energy store. Many plants contain large amounts of starch. Starch can be processed directly into a bioplastic but, because it is soluble in water, articles made from starch will swell and deform when exposed to moisture, and this limits its use. This problem can be overcome by changing starch into a different polymer. First, starch is harvested from corn, wheat or potatoes, then microorganisms transform it into lactic acid, a monomer. Finally, the lactic acid is chemically treated to cause the molecules of lactic acid to link up into long chains or polymers, which bond together to form a plastic called polylactide (PLA).

PLA can be used for products such as plant pots and disposable nappies. It has been commercially available in some countries since 1990, and certain blends have proved successful in medical implants, sutures and drug delivery systems because they are able to dissolve away over time. However, because PLA is much more expensive than normal plastics, it has not become as popular as one would have hoped.

Questions

1. In your own words, explain what is meant by a 'biodegradable plastic'.
2. Using your knowledge of chemical bonding, explain why some polymers are biodegradable and others are not.
3. Explain why lactic acid is a more useful monomer than starch, when making a biodegradable plastic.
4. If you were a consumer (shopper), would you choose to buy a biodegradable plastic rather than another? Explain your answer.
5. What do you think could be done to make biodegradable plastics more popular with consumers?

10.5.3 Plastics and the environment

Although plastics have had a huge impact globally, there is also an environmental price that has to be paid for their use. The following are just some of the ways in which plastics can cause damage to the environment.

1. **Waste disposal**

   Plastics are not easily broken down by micro-organisms and therefore most are not easily biodegradeable. This leads to waste dispoal problems.

2. **Air pollution**

   When plastics burn, they can produce toxic gases such as carbon monoxide, hydrogen cyanide and hydrogen chloride (particularly from PVC and other plastics that contain chlorine and nitrogen).

3. **Recycling**

   It is very difficult to recycle plastics because each type of plastic has different properties and so different recycling methods may be needed for each plastic. However, attempts are being made to find ways of recycling plastics more effectively. Some plastics can be remelted and re-used, while others can be ground up and used as a filler. However, one of the problems with recycling plastics is that they have to be sorted according to plastic *type*. This process is difficult to do with machinery, and therefore needs a lot of labour. Alternatively, plastics should be re-used. In many countries, including South Africa, shoppers must now pay for plastic bags. This encourages people to collect and re-use the bags they already have.

Activity :: Case Study : Plastic pollution in South Africa

Read the following extract, taken from 'Planet Ark' (September 2003), and then answer the questions that follow.

> South Africa launches a programme this week to exterminate its "national flower" - the millions of used plastic bags that litter the landscape.
> Beginning on Friday, plastic shopping bags used in the country must be both thicker and more recyclable, a move officials hope will stop people from simply tossing them away. "Government has targeted plastic bags because they are the most visible kind of waste," said Phindile Makwakwa, spokeswoman for the Department of Environmental Affairs and Tourism. "But this is mostly about changing people's mindsets about the environment."
> South Africa is awash in plastic pollution. Plastic bags are such a common eyesore that they are dubbed "roadside daisies" and referred to as the national flower. Bill Naude of the Plastics Federation of South Africa said the country used about eight billion plastic bags annually, a figure which could drop by 50 percent if the new law works.

It is difficult sometimes to imagine exactly how much waste is produced in our country every year. Where does all of this go to? You are going to do some simple calculations to try to estimate the volume of plastic packets that is produced in South Africa every year.

1. Take a plastic shopping packet and squash it into a tight ball.
   (a) Measure the approximate length, breadth and depth of your squashed plastic bag.
   (b) Calculate the approximate volume that is occupied by the packet.
   (c) Now calculate the approximate volume of your classroom by measuring its length, breadth and height.
   (d) Calculate the number of squashed plastic packets that would fit into a classroom of this volume.
   (e) If South Africa produces an average of 8 billion plastic bags each year, how many clasrooms would be filled if all of these bags were thrown away and not re-used?
2. What has South Africa done to try to reduce the number of plastic bags that are produced?
3. Do you think this has helped the situation?
4. What can *you* do to reduce the amount of plastic that you throw away?

10.6 Biological Macromolecules

A *biological macromolecule* is one that is found in living organisms. Biological macromolecules include molecules such as carbohydrates, proteins and nucleic acids. Lipids are also biological macromolecules. They are essential for all forms of life to survive.

Definition: Biological macromolecule
A biological macromolecule is a polymer that occurs naturally in living organisms. These molecules are essential to the survival of life.

10.6.1 Carbohydrates

Carbohydrates include the sugars and their polymers. One key characteristic of the carbohydrates is that they contain only the elements carbon, hydrogen and oxygen. In the carbohydrate monomers, every carbon except one has a hydroxyl group attached to it, and the remaining carbon atom is double bonded to an oxygen atom to form a carbonyl group. One of the most important monomers in the carbohydrates is **glucose** (figure 10.10). The glucose molecule can exist in an open-chain (acyclic) and ring (cyclic) form.

(a)

(b)

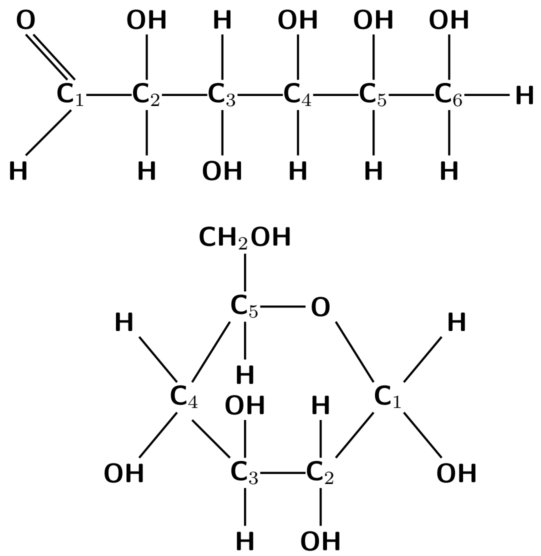


Figure 10.10: The open chain (a) and cyclic (b) structure of a glucose molecule

Glucose is produced during **photosynthesis**, which takes place in plants. During photosynthesis, sunlight (solar energy), water and carbon dioxide are involved in a chemical reaction that produces glucose and oxygen. This glucose is stored in various ways in the plant.

The photosynthesis reaction is as follows:

$$6CO_2 + 6H_2O + \text{sunlight} \rightarrow C_6H_{12}O_6 + 6O_2$$

Glucose is an important source of **energy** for both the plant itself, and also for the other animals and organisms that may feed on it. Glucose plays a critical role in **cellular respiration**, which is a chemical reaction that occurs in the cells of all living organisms. During this reaction, glucose and oxygen react to produce carbon dioxide, water and ATP energy. ATP is a type of energy that can be used by the body's cells so that they can function normally. The purpose of *eating* then, is to obtain glucose which the body can then convert into the ATP energy it needs to be able to survive.

The reaction for cellular respiration is as follows:

$$6C_6H_{12}O_6 + 60_2 \rightarrow 6CO_2 + 6H_2O + ATP \text{ (cell energy)}$$

We don't often eat glucose in its simple form. More often, we eat complex carbohydrates that our bodies have to break down into individual glucose molecules before they can be used in cellular respiration. These complex carbohydrates are polymers, which form through condensation polymerisation reactions (figure 10.11). *Starch* and *cellulose* are two example of carbohydrates that are polymers composed of glucose monomers.

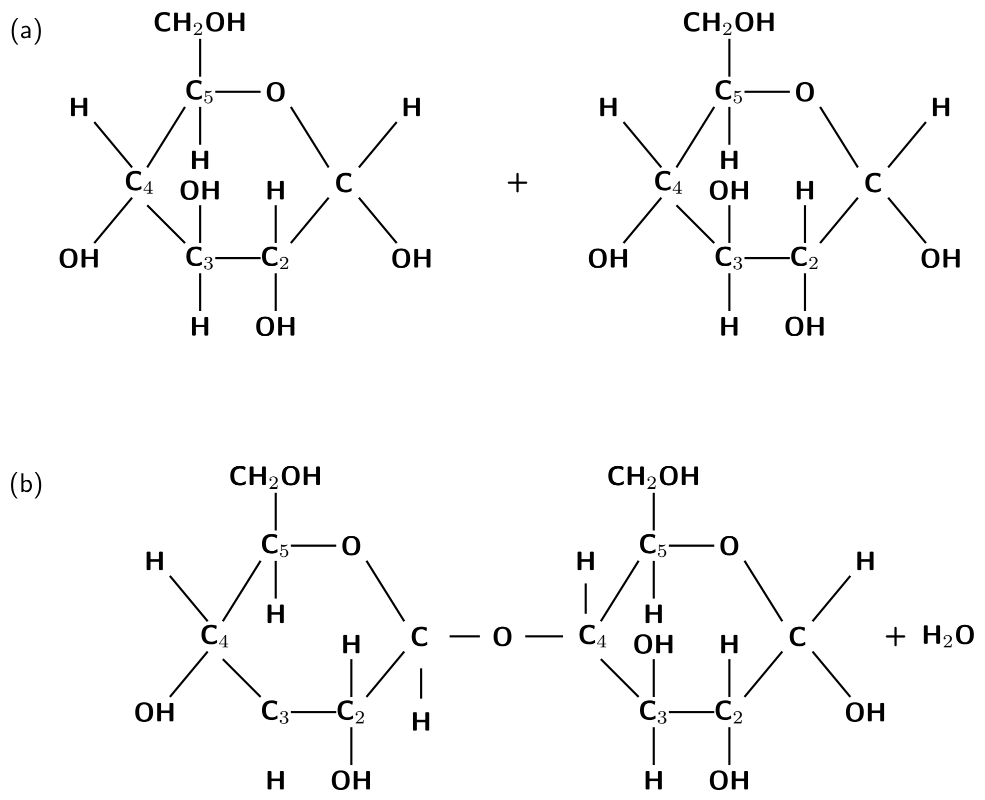


Figure 10.11: Two glucose monomers (a) undergo a condensation reaction to produce a section of a carbohydrate polymer (b). One molecule of water is produced for every two monomers that react.

- **Starch**

  Starch is used by plants to store excess glucose, and consists of long chains of glucose monomers. Potatoes are made up almost entirely of starch. This is why potatoes are such a good source of energy. Animals are also able to store glucose, but in this case it is stored as a compound called **glycogen**, rather than as starch.

- **Cellulose**

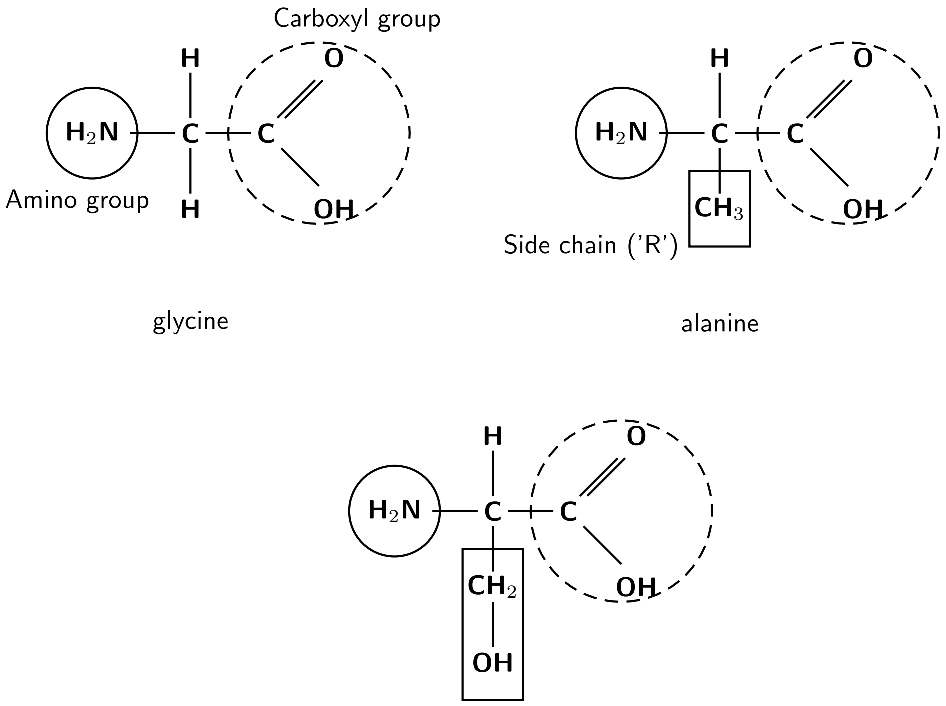
  Cellulose is also made up of chains of glucose molecules, but the bonding between the polymers is slightly different from that in starch. Cellulose is found in the cell walls of plants and is used by plants as a building material.

Interesting Fact

It is very difficult for animals to digest the cellulose in plants that they may have been feeding on. However, fungi and some protozoa are able to break down cellulose. Many animals, including termites and cows, use these organisms to break cellulose down into glucose, which they can then use more easily.

10.6.2 Proteins

Proteins are an incredibly important part of any cell, and they carry out a number of functions such as support, storage and transport within the body. The monomers of proteins are called **amino acids**. An amino acid is an organic molecule that contains a carboxyl and an amino group, as well as a carbon side chain. The carbon side chain varies from one amino acid to the next, and is sometimes simply represented by the letter 'R' in a molecule's structural formula. Figure 10.12 shows some examples of different amino acids.



glycine

alanine

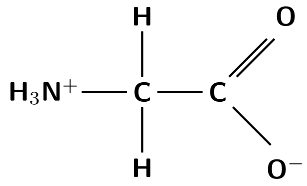
serine

Figure 10.12: Three amino acids: glycine, alanine and serine

Although each of these amino acids has the same basic structure, their side chains ('R' groups) are different. In the amino acid glycine, the side chain consists only of a hydrogen atom, while alanine has a *methyl* side chain. The 'R' group in serine is CH_2 - OH. Amongst other things, the side chains affect whether the amino acid is *hydrophilic* (attracted to water) or *hydrophobic* (repelled by water). If the side chain is *polar*, then the amino acid is hydrophilic, but if the side chain is *non-polar* then the amino acid is hydrophobic. Glycine and alanine both have non-polar side chains, while serine has a polar side chain.

Extension: Charged regions in an amino acid

In an amino acid, the amino group acts as a base because the nitrogen atom has a pair of unpaired electrons which it can use to bond to a hydrogen ion. The amino group therefore attracts the hydrogen ion from the carboxyl group, and ends up having a charge of $+1$. The carboxyl group from which the hydrogen ion has been taken then has a charge of -1. The amino acid glycine can therefore also be represented as shown in the figure below.



glycine

When two amino acid monomers are close together, they may be joined to each other by **peptide bonds** (figure 10.13) to form a **polypeptide** chain. . The reaction is a condensation reaction. Polypeptides can vary in length from a few amino acids to a thousand or more. The polpeptide chains are then joined to each other in different ways to form a **protein**. It is the sequence of the amino acids in the polymer that gives a protein its particular properties.

The sequence of the amino acids in the chain is known as the protein's **primary structure**. As the chain grows in size, it begins to twist, curl and fold upon itself. The different parts of the polypeptide are held together by hydrogen bonds, which form between hydrogen atoms in one part of the chain and oxygen or nitrogen atoms in another part of the chain. This is known as the **secondary structure** of the protein. Sometimes, in this coiled helical structure, bonds may form between the side chains (R groups) of the amino acids. This results in even more irregular contortions of the protein. This is called the **tertiary structure** of the protein.

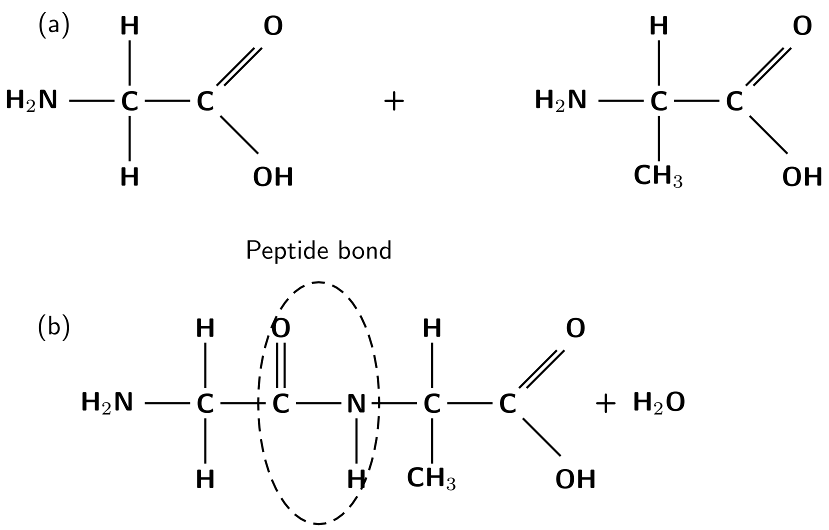


Figure 10.13: Two amino acids (glycine and alanine) combine to form part of a polypeptide chain. The amino acids are joined by a peptide bond between a carbon atom of one amino acid and a nitrogen atom of the other amino acid.

There are twenty different amino acids that exist. All cells, both plant and animal, build their proteins from only twenty amino acids. At first, this seems like a very small number, especially considering the huge number of different proteins that exist. However, if you consider that most proteins are made up of polypeptide chains that contain at least 100 amino acids, you will start to realise the endless possible combinations of amino acids that are available.

The functions of proteins

Proteins have a number of functions in living organisms.

- *Structural proteins* such as collagen in animal connective tissue and keratin in hair, horns and feather quills, all provide support.
- *Storage proteins* such as albumin in egg white provide a source of energy. Plants store proteins in their seeds to provide energy for the new growing plant.
- *Transport proteins* transport other substances in the body. Haemoglobin in the blood for example, is a protein that contains iron. Haemoglobin has an affinity (attraction) for oxygen and so this is how oxygen is transported around the body in the blood.
- *Hormonal proteins* coordinate the body's activities. Insulin for example, is a hormonal protein that controls the sugar levels in the blood.
- *Enzymes* are chemical catalysts and speed up chemical reactions. Digestive enzymes such as salivary amylase in your saliva, help to break down polymers in food. Enzymes play an important role in all cellular reactions such as respiration, photosynthesis and many others.

Activity :: Research Project : Macromolecules in our daily diet

1. In order to keep our bodies healthy, it is important that we eat a balanced diet with the right amounts of carbohydrates, proteins and fats. Fats are an important source of energy, they provide insulation for the body, and they also provide a protective layer around many vital organs. Our bodies also need certain essential vitamins and minerals. Most food packaging has a label that provides this information.
   Choose a number of different food items that you eat. Look at the food label for each, and then complete the following table:

Food	Carbohydrates (%)	Proteins (%)	Fats (%)

   (a) Which food type contains the largest proportion of protein?
   (b) Which food type contains the largest proportion of carbohydrates?
   (c) Which of the food types you have listed would you consider to be the 'healthiest'? Give a reason for your answer.

2. In an effort to lose weight, many people choose to *diet*. There are many diets on offer, each of which is based on particular theories about how to lose weight most effectively. Look at the list of diets below:
   - Vegetarian diet
   - Low fat diet
   - Atkin's diet

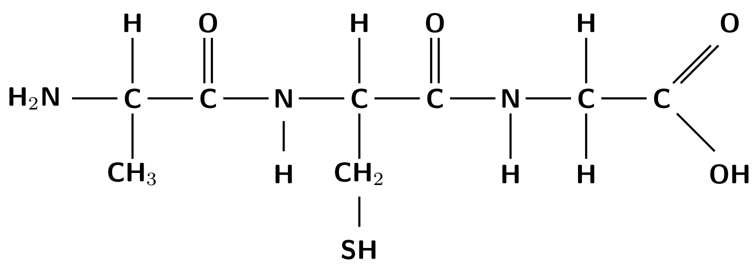
- Weight Watchers

For each of these diets, answer the following questions:

(a) What theory of weight loss does each type of diet propose?
(b) What are the *benefits* of the diet?
(c) What are the potential *problems* with the diet?

Exercise: Carbohydrates and proteins

1. Give the structural formula for each of the following:
   (a) A polymer chain, consisting of three glucose molecules.
   (b) A polypeptide chain, consisting of two molecules of alanine and one molecule of serine.
2. Write balanced equations to show the polymerisation reactions that produce the polymers described above.
3. The following polypeptide is the end product of a polymerisation reaction:



   (a) Give the structural formula of the monomers that make up the polypeptide.
   (b) On the structural formula of the first monomer, label the amino group and the carboxyl group.
   (c) What is the chemical formula for the carbon side chain in the second monomer?
   (d) Name the bond that forms between the monomers of the polypeptide.

10.6.3 Nucleic Acids

You will remember that we mentioned earlier that each protein is different because of its unique sequence of amino acids. But what controls how the amino acids arrange themselves to form the specific proteins that are needed by an organism? This task is for the **gene**. A gene contains DNA (deoxyribonucleic acid) which is a polymer that belongs to a class of compounds called the **nucleic acids**. DNA is the genetic material that organisms inherit from their parents. It is DNA that provides the genetic coding that is needed to form the specific proteins that an organism needs. Another nucleic acid is RNA (ribonucleic acid).

The DNA polymer is made up of monomers called **nucleotides**. Each nucleotide has three parts: a sugar, a phosphate and a nitrogenous base. The diagram in figure 10.14 may help you to understand this better.

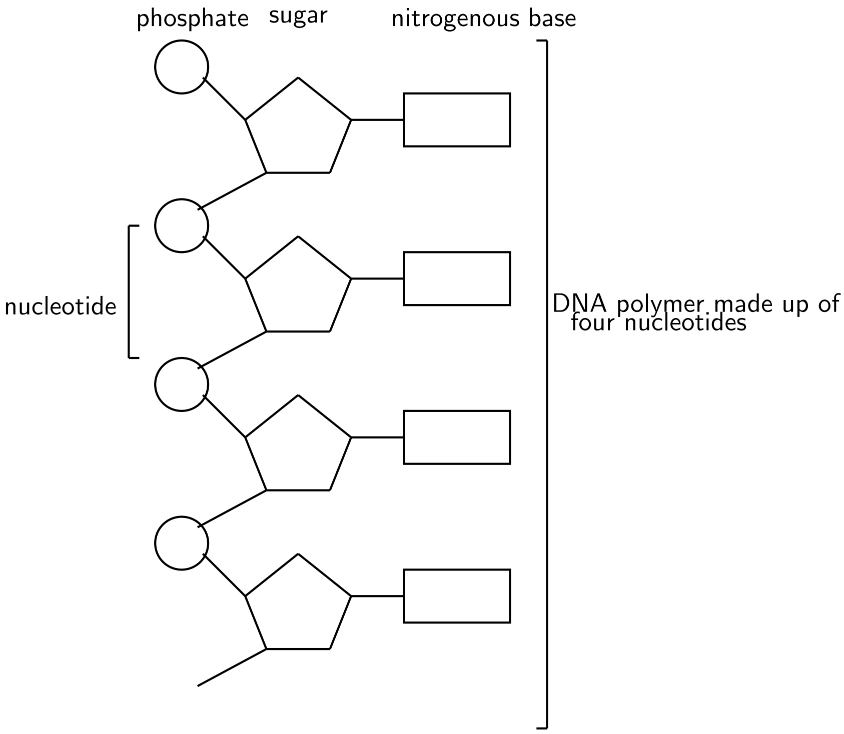


Figure 10.14: Nucleotide monomers make up the DNA polymer

There are five different nitrogenous bases: adenine (A), guanine (G), cytosine (C), thymine (T) and uracil (U). It is the sequence of the nitrogenous bases in a DNA polymer that will determine the genetic code for that organism. Three consecutive nitrogenous bases provide the coding for one amino acid. So, for example, if the nitrogenous bases on three nucleotides are *uracil*, *cytosine* and *uracil* (in that order), one **serine** amino acid will become part of the polypeptide chain. The polypeptide chain is built up in this way until it is long enough (and with the right amino acid sequence) to be a protein. Since proteins control much of what happens in living organisms, it is easy to see how important nucleic acids are as the starting point of this process.

A single defect in even one nucleotide, can be devastating to an organism. One example of this is a disease called **sickle cell anaemia**. Because of one wrong nucletide in the genetic code, the body produces a protein called **sickle haemoglobin**. Haemoglobin is the protein in red blood cells that helps to transport oxygen around the body. When sickle haemoglobin is produced, the red blood cells change shape. This process damages the red blood cell membrane, and can cause the cells to become stuck in blood vessels. This then means that the red blood cells, whcih are carrying oxygen, can't get to the tissues where they are needed. This can cause serious organ damage. Individuals who have sickle cell anaemia generally have a lower life expectancy.

Table 10.2 shows some other examples of genetic coding for different amino acids.

Exercise: Nucleic acids

Table 10.2: Nitrogenouse base sequences and the corresponding amino acid

Nitrogenous base sequence	Amino acid
UUU	Phenylalanine
CUU	Leucine
UCU	Serine
UAU	Tyrosine
UGU	Cysteine
GUU	Valine
GCU	Alanine
GGU	Glycine

1. For each of the following, say whether the statement is **true** or **false**. If the statement is *false*, give a reason for your answer.
   (a) Deoxyribonucleic acid (DNA) is an example of a *polymer* and a nucleotide is an example of a *monomer*.
   (b) Thymine and uracil are examples of nucleotides.
   (c) A person's DNA will determine what proteins their body will produce, and therefore what characteristics they will have.
   (d) An amino acid is a protein monomer.
   (e) A polypeptide that consists of five amino acids, will also contain five nucleotides.
2. For each of the following sequences of nitrogenous bases, write down the amino acid/s that will be part of the polypeptide chain.
   (a) UUU
   (b) UCUUUU
   (c) GGUUAUGUUGGU
3. A polypeptide chain consists of three amino acids. The sequence of nitrogenous bases in the nucleotides of the DNA is GCUGGUGCU. Give the structural formula of the polypeptide.

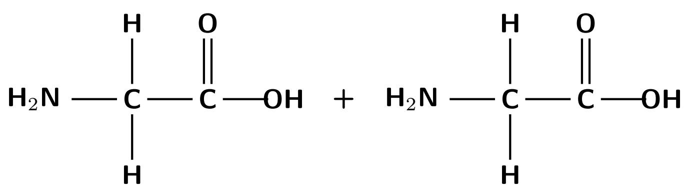
10.7 Summary

- A **polymer** is a macromolecule that is made up of many repeating structural units called **monomers** which are joined by covalent bonds.
- Polymers that contain carbon atoms in the main chain are called **organic polymers**.
- Organic polymers can be divided into **natural organic polymers** (e.g. natural rubber) or **synthetic organic polymers** (e.g. polystyrene).
- The polymer **polyethene** for example, is made up of many ethene monomers that have been joined into a polymer chain.
- Polymers form through a process called **polymerisation**.
- Two examples of polymerisation reactions are **addition** and **condensation** reactions.
- An **addition reaction** occurs when unsaturated monomers (e.g. alkenes) are added to each other one by one. The breaking of a double bond between carbon atoms in the monomer, means that a bond can form with the next monomer. The polymer **polyethene** is formed through an addition reaction.

- In a **condensation reaction**, a molecule of water is released as a product of the reaction. The water molecule is made up of atoms that have been lost from each of the monomers. Polyesters and nylon are polymers that are produced through a condensation reaction.
- The **chemical properties** of polymers (e.g. tensile strength and melting point) are determined by the types of atoms in the polymer, and by the strength of the bonds between adjacent polymer chains. The stronger the bonds, the greater the strength of the polymer, and the higher its melting point.
- One group of synthetic organic polymers, are the **plastics**.
- **Polystyrene** is a plastic that is made up of styrene monomers. Polystyrene is used a lot in packaging.
- **Polyvinyl chloride** (PVC) consists of vinyl chloride monomers. PVC is used to make pipes and flooring.
- **Polyethene**, or **polyethylene**, is made from ethene monomers. Polyethene is used to make film wrapping, plastic bags, electrical insulation and bottles.
- **Polytetrafluoroethylene** is used in non-stick frying pans and electrical insulation.
- A **thermoplastic** can be heated and melted to a liquid. It freezes to a brittle, glassy state when cooled very quickly. Examples of thermoplastics are polyethene and PVC.
- A **thermoset** plastic cannot be melted or re-shaped once formed. Examples of thermoset plastics are vulcanised rubber and melanine.
- It is not easy to **recycle** all plastics, and so they create environmental problems.
- Some of these **environmental problems** include issues of waste disposal, air pollution and recycling.
- A **biological macromolecule** is a polymer that occurs naturally in living organisms.
- Examples of biological macromolecules include **carbohydrates** and **proteins**, both of which are essential for life to survive.
- Carbohydrates include the **sugars** and their polymers, and are an important source of **energy** in living organisms.
- **Glucose** is a carbohydrate monomer. Glucose is the molecule that is needed for **photosynthesis** in plants.
- The glucose monomer is also a building block for carbohydrate polymers such as **starch**, **glycogen** and **cellulose**.
- **Proteins** have a number of important functions. These include their roles in structures, transport, storage, hormonal proteins and enzymes.
- A protein consists of monomers called **amino acids**, which are joined by **peptide bonds**.
- A protein has a **primary**, **secondary** and **tertiary** structure.
- An amino acid is an organic molecule, made up of a **carboxyl** and an **amino** group, as well as a carbon **side chain** of varying lengths.
- It is the **sequence** of amino acids that determines the nature of the protein.
- It is the **DNA** of an organism that determines the order in which amino acids combine to make a protein.
- DNA is a **nucleic acid**. DNA is a polymer, and is made up of monomers called **nucleotides**.
- Each nucleotide consists of a **sugar**, a **phosphate** and a **nitrogenous base**. It is the sequence of the nitrogenous bases that provides the 'code' for the arrangement of the amino acids in a protein.

Exercise: Summary exercise

1. Give one word for each of the following descriptions:
   (a) A chain of monomers joined by covalent bonds.
   (b) A polymerisation reaction that produces a molecule of water for every two monomers that bond.
   (c) The bond that forms between an alcohol and a carboxylic acid monomer during a polymerisation reaction.
   (d) The name given to a protein monomer.
   (e) A six-carbon sugar monomer.
   (f) The monomer of DNA, which determines the sequence of amino acids that will make up a protein.
2. For each of the following questions, choose the one correct answer from the list provided.
   (a) A polymer is made up of monomers, each of which has the formula $CH_2{=}CHCN$. The formula of the polymer is:
      i. $-(CH_2{=}CHCN)_n-$
      ii. $-(CH_2\text{-}CHCN)_n-$
      iii. $-(CH\text{-}CHCN)_n-$
      iv. $-(CH_3\text{-}CHCN)_n-$
   (b) A polymer has the formula $-[CO(CH_2)_4CO\text{-}NH(CH_2)6NH]_n-$. Which of the following statements is **true**?
      i. The polymer is the product of an addition reaction.
      ii. The polymer is a polyester.
      iii. The polymer contains an amide linkage.
      iv. The polymer contains an ester linkage.
   (c) Glucose...
      i. is a monomer that is produced during cellular respiration
      ii. is a sugar polymer
      iii. is the monomer of starch
      iv. is a polymer produced during photosynthesis
3. The following monomers are involved in a polymerisation reaction:



   (a) Give the structural formula of the polymer that is produced.
   (b) Is the reaction an addition or condensation reaction?
   (c) To what group of organic compounds do the two monomers belong?
   (d) What is the name of the monomers?
   (e) What type of bond forms between the monomers in the final polymer?
4. The table below shows the melting point for three plastics. Suggest a reason why the melting point of PVC is *higher* than the melting point for polyethene, but *lower* than that for polyester.

Plastic	**Melting point** (^{0}C)
Polyethene	105 - 115
PVC	212
Polyester	260

5. An amino acid has the formula $H_2NCH(CH_2CH_2SCH_3)COOH$.
   (a) Give the structural formula of this amino acid.
   (b) What is the chemical formula of the carbon side chain in this molecule?
   (c) Are there any peptide bonds in this molecule? Give a reason for your answer.

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