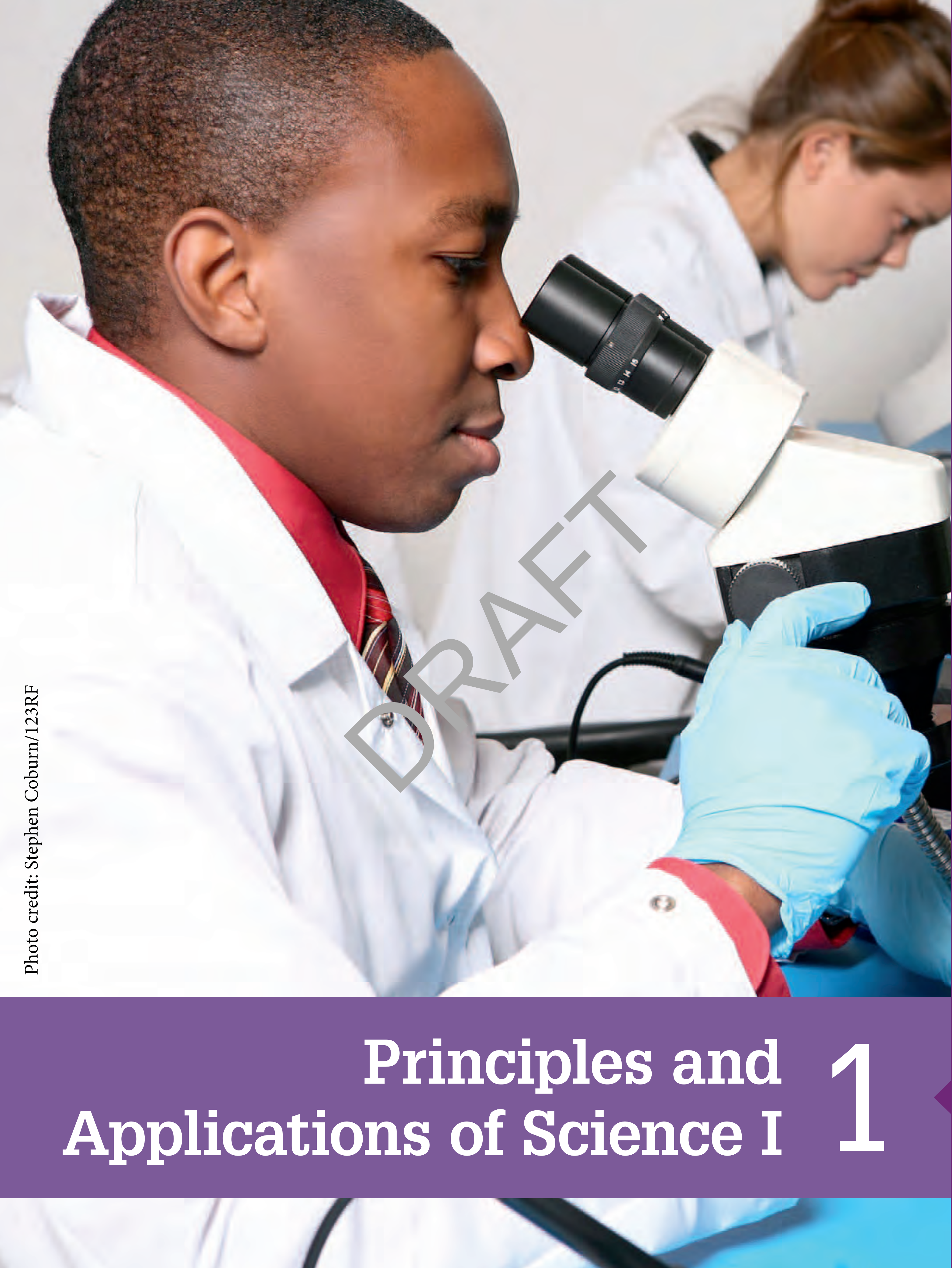


Photo credit: Stephen Coburn/123RF

Principles and Applications of Science I

1

Getting to know your unit

All scientists and technicians need to understand core science concepts. Chemists need to understand atoms and electronic structure to predict how a range of chemical substances will react to make useful products. Medical professionals need to understand the structure and workings of cells when they think about how the body stays healthy as well as when diagnosing and treating illness.

Scientists working in the communication industry need a good understanding of waves.

Assessment

You will be assessed through a 90-minute written exam worth 90 marks, which is set and marked by Pearson.

How you will be assessed

The external paper for this unit will be split into three sections, each worth 30 marks.

- **Section A** – Biology (Cell structure and function, Cell specialisation, Tissue structure and function)
- **Section B** – Chemistry (Structure and bonding in applications of science, Production and uses of substances in relation to properties)
- **Section C** – Physics (Working with waves, Waves in communication, Use of electromagnetic waves in communication)

The paper will contain a range of question types, including multiple choice, calculations, short answer and open response. These question types by their very nature generally assess discrete knowledge and understanding of content in this unit.

You need to be able to apply and synthesise knowledge from this unit. The questions on the paper will be contextualised in order for you to show you can do this.

There will be two opportunities each year to sit this paper: January and May/June.

Throughout this chapter, you will find assessment practices that will help you prepare for the exam. Completing each of these will give you an insight into the types of questions that will be asked and, importantly, how to answer them.

Unit 1 has four Assessment Outcomes (AO) which will be included in the external examination. These are:

- **AO1** Demonstrate knowledge of scientific facts, terms definitions and scientific formulae
  - Command words: give, label, name, state
  - Marks: ranges 12 to 18 marks
- **AO2**: demonstrate understanding of scientific concepts, procedures, processes and techniques and their application
  - Command words: calculate, compare, discuss, draw, explain, state, write
  - Marks: ranges from 30 to 45 marks
- **AO3**: Analyse, interpret and evaluate scientific information to make judgements and reach conclusions
  - Command words: calculate, compare, comment complete, describe, discuss, explain, state
  - Marks: ranges from 18 to 24 marks
- **AO4** Make connections, use and integrate different scientific concepts, procedures, processes or techniques
  - Command words: compare, comment, discuss, explain
  - Marks: ranges from 9 to 12 marks

Here are some of the command words. The rest are found in the specifications.

Command word	Definition – what it is asking you to do
Analyse	Identify several relevant facts of a topic, demonstrate how they are linked and then explain the importance of each, often in relation to the other facts.
Compare	Identify the main factors of two or more items and point out their similarities and differences. You may need to say which are the most best or most important. The word *Contrast* is very similar.
Comment	Requires the synthesis of a number of variables from data/information to form a judgement. More than 2 factors need to be synthesised.
Define	State the meaning of something, using clear and relevant facts.
Describe	Give a full account of all the information, including all the relevant details of any features, of a topic.
Discuss	Write about the topic in detail, taking into account different ideas and opinions.
Evaluate	Bring all the relevant information you have on a topic together and make a judgment on it (for example on its success or importance). Your judgment should be clearly supported by the information you have gathered.
Explain	Make an idea, situation or problem clear to your reader, by describing it in detail, including any relevant data or facts.

Getting started

Scientists working in a hospital laboratory use a range of core scientific principles. Write a list of core scientific principles you think they might need and why they are useful. Remember these may be to do with physics, chemistry or biology. When you have completed this unit, see if you can add any more principles to your list.

A Periodicity and properties of elements

A1 Structure and bonding in applications in science

The electronic structure of atoms

You should already know about the structure of an atom. The nucleus contains positive protons and neutral nutrons. Surrounding the nucleus are energy shells containing negative electrons. You should also know that protons and neutrons both have a relative mass of 1 and that the relative mass of an electron us almost 0.

Lab technicians need to understand the electronic structure of atoms. They can use this knowledge to predict how chemical substances will behave and react.

The protons and the neutrons are found in the nucleus at the centre of an atom. The electrons are in shells or energy levels surrounding the nucleus. Each shell can hold electrons up to a maximum number. When the first shell is full electrons then go into the second shell and so on. The maximum number of electrons in each shell is shown in Table 1.1.

Table 1.1: Maximum number of electrons for each electron shell shell

Electron shell	Maximum number of electrons
1	2
2	8
3	18
4	32
5	50

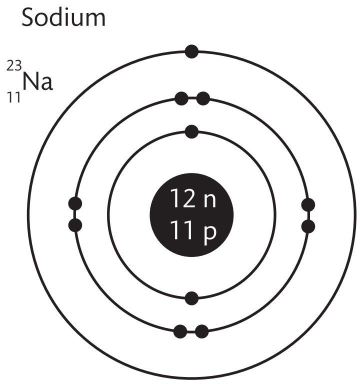


Figure 1.1: Simple atomic structure of sodium

A sodium atom containing 11 electrons has an electron arrangement of 2, 8, 1.

This can be represented by a simple Bohr diagram, as shown in Figure 1.1.

This is the simple version of electron structure you will have seen at Key Stage 4.

Under Bohr's theory, an electron's shells can be imagined as orbiting circles around the nucleus.

However, it is more complicated than this. Electrons within each shell will not have the same amount of energy and so the energy levels or shells are broken down into sub-shells called **orbitals**. These are called s, p, d and f orbitals. The orbitals have different energy states.

Key term

Orbitals – regions where there is a 95% probability of locating an electron. An orbital can hold a maximum of two electrons.

The Aufbau principle states that electrons fill the orbital with the lowest available energy state in relation to the proximity to the nucleus before filling orbitals with higher energy states. This gives the most stable **electron configuration** possible.

Electrons have the same charge and so repel each other so if there is more than one orbital in an energy level (sub-shell) they will fill them singly until all the orbitals in that sub-shell have an electron in them and then they will pair up.

Figure 1.2 shows the energy levels of the shells, sub-shells and orbitals for an atom.

Key terms

Electron configuration – the distribution of electrons in an atom or molecule.

Spin – electrons have two possible states, 'spin up' and 'spin down'. In an orbital, each electron will be in a different 'spin state'.

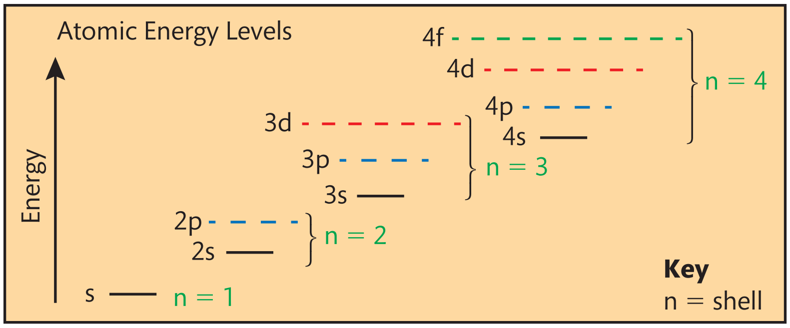


Figure 1.2: Energy levels of the shells subshells and orbitals for an atom

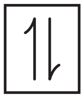
Step-by-step: Electron structures

8 Steps

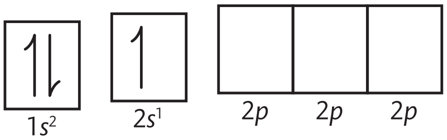
When writing out electron structures, you should follow these rules.
Half arrows are used to represent each electron in the orbital. They are drawn facing up and down as each electron in an orbital will have a different **spin**.

1 The electrons sit in orbitals within the shell. Each orbital can hold up to two electrons.

2 The first shell can hold two electrons in an *s*-type orbital.



3 The second shell consists of one *s*-type orbital and three *p*-type orbitals. This diagram represents lithium.

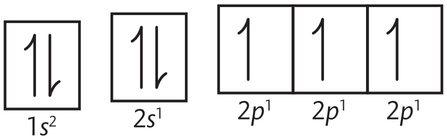


4 The third shell consists of one *s*-type orbital, three *p*-type orbitals and five *d*-type orbitals.

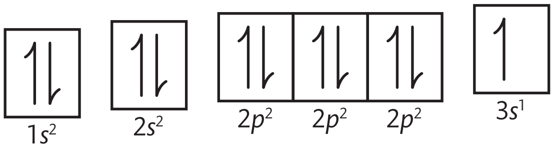
5 Electrons fill the lowest energy level orbitals first.

6 Where there are several orbitals of exactly the same energy, for example, the three *2p* orbitals in the second shell, then the electrons will occupy different orbitals wherever possible.

7 So the electronic structure of nitrogen (which has 7 electrons) is:



8 and the electronic structure of a sodium atom (which has 11 electrons) becomes:



Assessment practice 1.1

Copy out the following table and complete the electronic structures for the elements. Three have been done for you.

Element	Number of electrons	Electron structure
hydrogen	1	$1s^1$
helium		
lithium		
boron		
carbon	6	$1s^2\ 2s^2\ 2p^2$
oxygen	8	$1s^2\ 2s^2\ 2p^4$
magnesium		
chlorine		
calcium		

PAUSE POINT Try explaining what you have learned so far.

Hint Close the book and write out all the key concepts you have learned so far. What do you know about electronic structure? Could you draw the electronic structure for calcium?

Extend What is new compared to what you learned at level 2 about electronic structure?

One of the tasks of a lab technician is to make up solutions ready for experiments or for making products. Different types of compounds dissolve in different types of solvents depending on what type of bonding is in the compound. The lab technician must know what type of compound they are using in order to use the correct solvent.

Ionic bonding

Noble gases (elements in group 0 of the periodic table) have a stable electronic configuration. They have full outer shells. This means they do not react easily and most do not react at all. Elements in the other groups do not have full outer shells. This means that they react to gain stable electronic configurations.

Key term

Ionic bonding – electrostatic attraction between two oppositely charged ions.

Ionic bonding occurs when an atom of an element loses one or more electron and donates it to an atom of a different element. The atom that loses electrons becomes positively charged and the atom that gains electron(s) become negatively charged because of the imbalance of protons and electrons.

For example, the bonding in sodium chloride is ionic. This means that the sodium atom loses the electron in its outer shell to become the positively charged sodium ion, Na^+, with the same electron configuration as neon. Chlorine gains an electron to become the negatively charged chloride ion, Cl^-, with the same electron configuration as argon. This means that both the sodium ion and the chloride ion have a full outer shell and become stable. The positive charge on the sodium ion and the negative charge on the chloride ion are attracted.

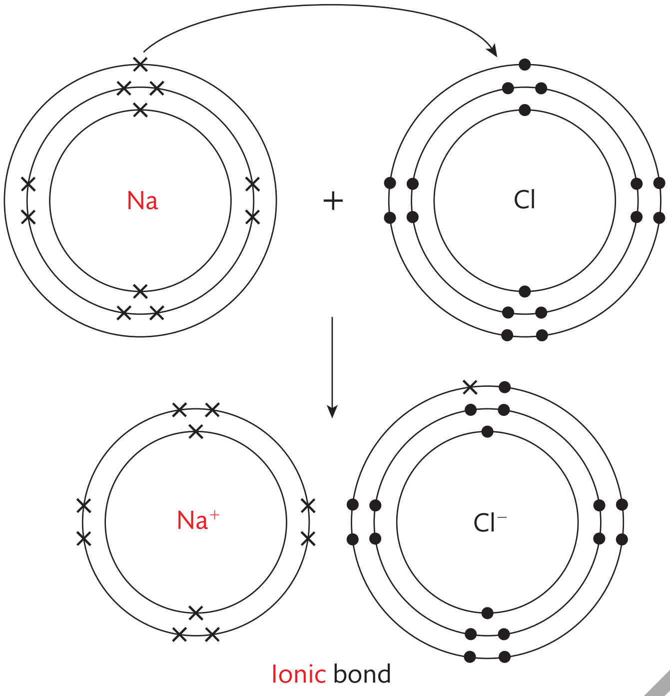


Figure 1.3: Electron transfer and bonding in sodium chloride

Figure 1.3 shows bonding using a dot and cross diagram. The dots and crosses represent electrons in the shells. The electrons fill up the shells singly and then in pairs.

Ions containing more than one element can also be formed. For example, in sodium hydroxide, Na$^+$ bonds with the hydroxide ion (OH)$^-$.

The opposite charges on the ions are what hold them together. This is **electrostatic attraction**.

The opposite charged ions in sodium chloride form a **giant ionic lattice** (see Figure 1.4) where the ions are arranged in a regular pattern.

Key terms

Electrostatic attraction – the force experienced by oppositely charged particles. It holds the particles strongly together.

Giant ionic lattice – a regular arrangement of positive ions and negative ions, for example, in NaCl.

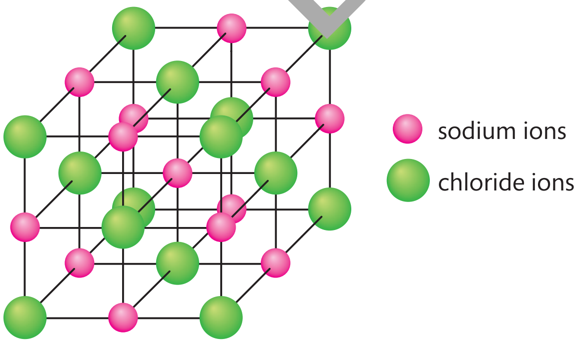


Figure 1.4: Lattice structure of sodium chloride

The strength of the electrostatic force and, therefore, of the ionic bond is dependent on the ionic charge and the ionic radii of the ions. The more electrons a positive ion has, the more shells it will have. If an ion has more shells, then its radius will be bigger than an ion with fewer shells.

The electrostatic force is stronger when the ionic charge is higher. However, the force becomes weaker if the ionic radii are bigger. This is because, when the ionic radius is bigger, the ionic charge is spread over a larger surface area.

Covalent bonding

Covalent bonding usually occurs between atoms of two non-metals. A covalent bond forms when an electron is shared between the atoms. These electrons come from the top energy level of the atoms.

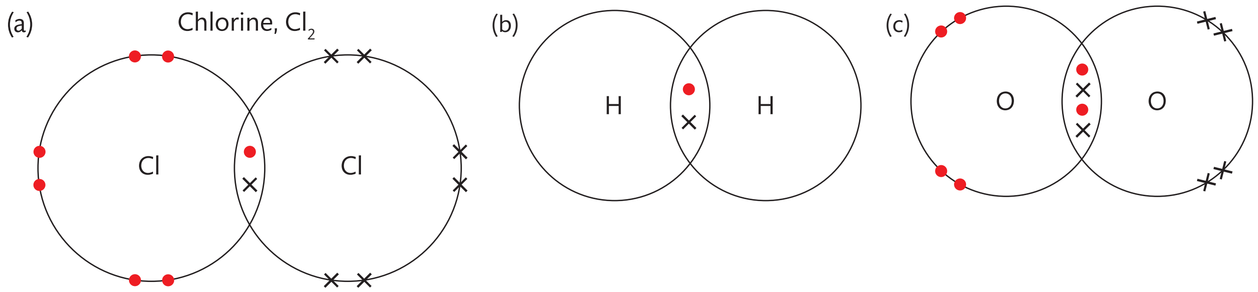


Figure 1.5: Covalent bonding in (a) a chlorine molecule (b) a hydrogen molecule (c) oxygen moelcule

A chlorine molecule has a covalent bond (see Figure 1.5). The highest shell in each chlorine atom contains seven electrons. One electron from the highest shell in each atom is shared to give each chlorine atom the electron configuration of argon with a stable full outer shell.

Dative covalent bonding

In some covalent molecules, both sharing electrons come from one atom. This is called a dative (coordinate) covalent bond. (see Figure 1.6).

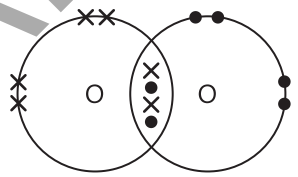


Figure 1.6: Covalent bonds can be formed when atoms share more than one pair of electrons, e.g. oxygen atoms. The double bonds between the oxygen are formed by two shared pairs of electrons.

If three pairs of electrons are shared, then a triple covalent bond is formed. A triple bond is present in a nitrogen molecule (see Figure 1.7).

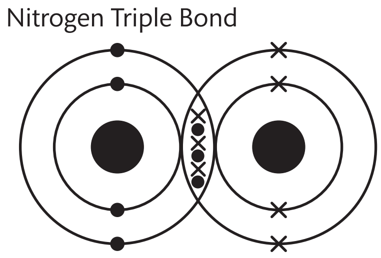


Figure 1.7: Bonding in a nitrogen molecule

Key term

Lone pair – a non-binding pair of electrons.

An ammonium ion contains a dative bond (see Figure 1.8). When ammonia reacts with hydrochloric acid, a hydrogen ion from the acid is transferred to the ammonia molecule. A **lone pair** of electrons on the nitrogen atom forms a dative covalent bond with the hydrogen ion.

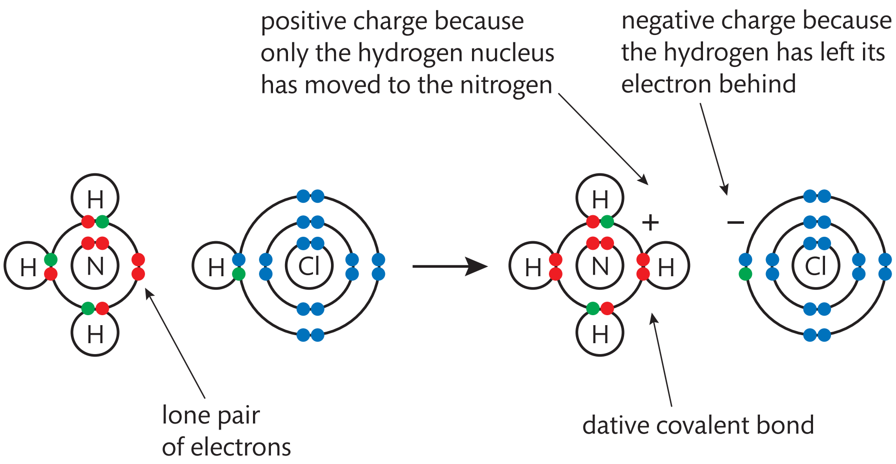


Image credit: after http://www.chemguide.co.uk/atoms/bonding/dative.html, Jim Clark

Figure 1.8: Dative bond formation in reaction between ammonia and hydrochloric acid

Single bonds have a greater length than double bonds and double bonds have a greater length than triple bonds. The shorter the length of the bond, the stronger the bond is. Therefore, triple bonds are stronger than double or single bonds. A single bond between carbon atoms has a length of 154 pm and a bond energy of 347 kJ mol^{-1}. A double bond between carbon atoms has a length of 134 pm and a bond energy of 612 kJ mol^{-1}. A triple bond between atoms has a bond length of 120 pm and a bond energy of 820 kJ mol^{-1}.

Covalent bonding in organic molecules

Carbon makes four covalent bonds so it forms many compounds which are called **organic compounds**.

Key term

Organic compound – a compound that contains one or more carbons in a carbon chain.

Methane has the formula CH_4. Each carbon atom bonds covalently with four hydrogen atoms. The carbon gains the stable electron structure of neon and hydrogen gains the stable electron structure of helium.

These four bonds mean that methane is not a flat molecule. It has a tetrahedral structure (see Figure 1.9). This is because the bonds are as separated from each other as possible, because the negative electron pairs repel each other, with each bond angle being 109.5°. If you were to build a model of a methane molecule, it would have a 3D shape with a hydrogen pointing down towards you, one pointing down away from you, one pointing down to the side and one pointing up, all connected to the carbon in the centre.

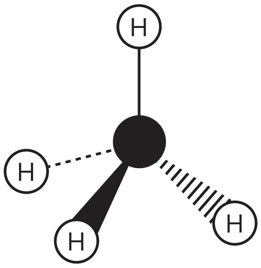


Figure 1.9: Tetrahedral structure of methane

Step by step: Building models of organic compounds

3 Steps

1 Use molecular model kits to build models of the following organic compounds.

- methane CH_4
- ethane CH_3CH_3
- propane $CH_3CH_2CH_3$.

2 Write down what you notice about the structure of these molecules.

3 Look at one of the carbons in each molecule and the atoms bonded to it. Write down what you notice about the shape.

Organic compounds with three or more carbons in a chain cannot be linear because of the tetrahedral structure around each central carbon (see Figure 1.10).

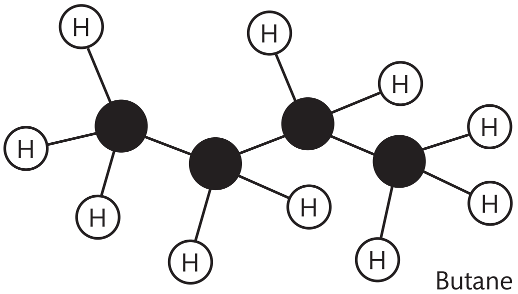


Figure 1.10: A butane model

Metallic bonding

Key term

Delocalised electron – electrons that are free to move. They are present in metals and are not associated with a single atom or covalent bond.

Metals are giant structures of atoms held together by metallic bonds. The metal structure is a regular lattice (see Figure 1.11).

Metallic bonding is caused because the electrons in the highest energy level of a metal atom has the ability to become **delocalised**. They are free to move through the metal in a 'sea' of electrons. This gives the metal nuclei a positive charge which is attracted to the negative charge on the delocalised electrons. There is a very strong force of attraction between the positive metal nuclei and the negative delocalised electrons. However, the forces in metallic bonding are not as strong as in covalent or ionic bonding.

Figure 1.11: Metallic structure

The metal structure is a lattice of positive ions with electrons flowing between these ions.

PAUSE POINT What have you learned about bonding?

Hint Describe the differences between ionic, covalent and metallic bonding.

Extend Give two examples of elements, compounds or molecules with each type of bond.

The **electronegativity** of two atoms will determine what type of bond will form between them.

Key term

Electronegativity– This is the tendency of an atom to attract a bonding pair of electrons.

Atoms that have similar electronegativities form covalent bonds.

There is a strong electrostatic attraction between the two nuclei and the shared pair(s) of electrons between them. This is the covalent bond. Both chlorine atoms have the same electronegativity, and so the electrons are equally shared. The chlorine molecule is **non-polar** (see Figure 1.6). Hydrogen only has one shell containing one electron. This electron from each hydrogen is shared to give each atom the electronic configuration of helium. Oxygen only has 6 electrons in its highest energy shell. Each oxygen atom shares two of its electrons with another oxygen atom, giving both 8 electrons in their outer shell. This makes the atoms in the oxygen molecule stable.

[redacted]

Figure 1.12 : Non-polar covalent bond

In most covalent compounds, the bonding is **polar** covalent (see Figure 1.7). The shared electrons are attracted more to one nucleus in the molecule than the other. The atom with the higher electronegativity will attract the electrons more strongly. This gives the atom a slight negative charge. The other atom in the molecule will have a slight positive charge.

Key terms

Non-polar molecule – a molecule where the electrons are distributed evenly throughout the molecule.

Polar molecule – a molecule with partial positive charge in one part of the molecule and similar negative charge in another part due to an uneven electron distribution.

[redacted]

Figure 1.13: Polar covalent bond

As the difference in electronegativity between the atoms increases, the bond will become more polar. See Figure 1.4.

[redacted]

Figure 1.14: Electronegativity spectrum

The electronegativities of some of the common elements you will use are shown in Table 1.2.

Table 1.2: Electronegativities of elements

Element	Electronegativity
Fluorine	3.98
Oxygen	3.44
Nitrogen	3.04
Carbon	2.55
Chlorine	3.16
Hydrogen	2.20
Lithium	0.98
Sodium	0.82

Key terms

Intermolecular forces – the attraction or repulsion between neighbouring molecules.

Dipole – separation of charges within a covalent molecule.

Intermolecular forces

Intermolecular forces also affect how chemical substances behave. A laboratory technician must know where these are present and understand how they will affect the behaviour and reactions of chemical substances they are working with.

London dispersion forces

One type of intermolecular force is called London dispersion forces (also called **temporary dipole** – induced dipole forces). They are weak forces present between non-polar covalent molecules. They are less than 1% of the force of a covalent bond (see Figure 1.15).

When the electron distribution in a molecule becomes non-symmetrical (i.e. there are more electrons at one end of the molecule than the other) then one end of the molecule can become more positive and one end can become more negative. This causes a temporary dipole. The positive and negative charge in the dipole can disturb the electrons in a nearby molecule, repelling the electrons and so causing (inducing) a dipole in that molecule. The molecule with the temporary dipole and the molecule with the induced dipole attract each other and pull the molecules together. The forces are temporary because the electrons are constantly moving so electron density in any part of a molecule is constantly changing. Larger molecules have more electrons which can move further so more temporary dipoles can form, meaning the force is bigger.

more electrons → more movement → bigger dipoles → stronger attraction

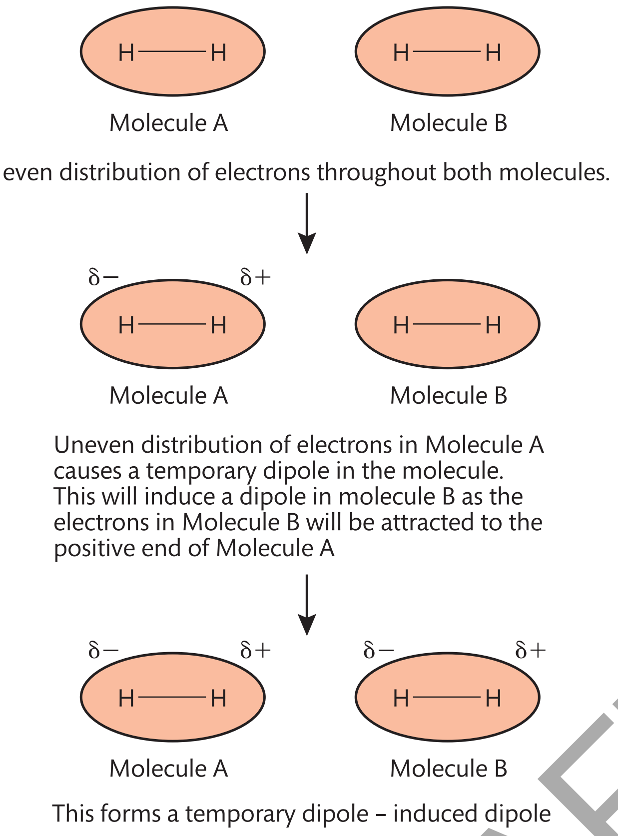


Figure 1.15: London Dispersion forces

London dispersion forces are the only forces that exist between noble gases and non-polar molecules.

Assessment practice 1.2

Pentane (C_5H_{12}) boils at 309 K and ethane (C_2H_6) boils at 185 K. This means that pentane is a liquid at room temperature (293 K) and ethane is a gas.

Explain why pentane is a liquid at room temperature but ethane is a gas.

Dipole-dipole forces

Another form of **van der Waals forces** are dipole-dipole forces. These are permanent forces between polar molecules (see Figure 1.16). Polar molecules have a permanent negative end and a permanent positive end. These oppositely charged end attract each other. Dipole-dipole forces are slightly stronger than London dispersion forces but are still weak in comparison to a covalent bond. The force is about 1% the strength of a covalent bond. Molecules that have permanent dipole-dipole forces include hydrogen chloride, HCl, and iodine monochloride, ICl. In both cases, the chlorine atom in the molecule is slightly negative. The hydrogen and iodine atoms are slightly positive.

Key term

Van der Waals forces – All intermolecular attractions are van der Waals forces.

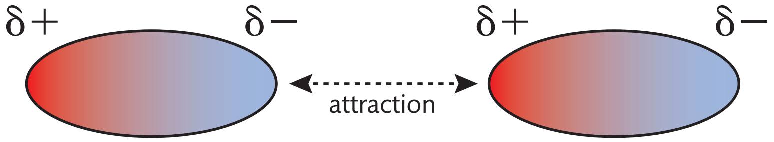


Figure 1.16: Dipole-dipole forces

There are dipole-dipole forces between molecules of iodine monochloride (ICl).

Hydrogen bonding

The strongest form of intermolecular force is a hydrogen bond. These are a special type of dipole-dipole bond and are forces that are about 10% the strength of a covalent bond. Hydrogen bonds will form when compounds have hydrogen directly bonded to fluorine, oxygen or nitrogen. This is because there is a large difference in electronegativity between hydrogen and and of these three atoms. This large difference means that very polar bonds are formed so the molecules have permanent dipoles. When two of these molecules are close together, there will be an attraction between the positive end of one and the lone pair of electrons of the other. This is a hydrogen bond.

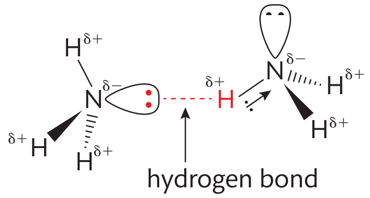


Figure 1.17: Hydrogen bond in ammonia

This is different to other dipole-dipole forces because there are inner bonding electrons. The single electron in the hydrogen atom is drawn to the nitrogen (see Figure 1.17), oxygen or fluorine atom. There are no non-bonding electrons shielding the nucleus of the hydrogen. The hydrogen proton is strongly attracted to the lone pair of electrons on the nitrogen atom of another molecule.

Discussion

Hydrogen bonding in water is the reason why water has such unusual properties, e.g. solid water is less dense than liquid water, it has a higher boiling point than expected, it is a good solvent for many chemical substances.

Research how hydrogen bonding is caused in a water molecule. Work in pairs to list properties of water due to the hydrogen bonding. In groups, explain the properties to other pairs of learners.

PAUSE POINT Try to describe all the different types of intermolecular forces to a partner.

Hint Draw a table showing the different types of intermolecular bonding and their properties.

Extend Explain how each type of intermolecular bond affects the properties of the molecules.

Quantities used in chemical reactions

Balancing equations

All chemical reactions can be written as a balanced equation using the chemical formulae for the reactants and the products involved in the reaction. Symbols for elements can be found in the periodic table. The numbers in the formulae show how many atoms of each element there are. You can use the periodic table to predict whether the compound is covalent or ionic. The group numbers will show you how many electrons the atom needs to lose or gain or share to form a bond.

The equation must balance like a maths equation. There should be the same number and types of atoms on both sides of the equation.

Step by step: Writing a balanced equation

5 Steps

1 Write the equation as a word equation including all the reactants and all the products.

2 Write out the formulae for each substance in the reaction.
Note that gaseous elements (except those in group 0) like hydrogen and oxygen are diatomic (molecules with two atoms) so they must be written as H_2 and O_2. Metal elements and the noble gases are monatomic (one atom)

3 Write out the number of each element on both sides.

4 Make the number of each atom equal on each side. Remember that you cannot change the formula of the compounds. To increase the number of atoms of a particular element, you must place a number in front of the compound it is in. This will affect the number of atoms of all the other elements in the compound.

5 Check that there is the same number of atoms of each element on both sides.

Worked Example

1 Write a balanced equation for the following reaction.

Ethanol + oxygen ⟶ carbon dioxide +water

Step 1: Write out the formulae for each substance in the reaction.

$C_2H_5OH + O_2 \longrightarrow CO_2 + H_2O$

Step 2: Write out the number of each element on both sides.

left-hand side	right-hand side
C 2	C 1
H 6	H 2
O 3	O 3

Step 3: Make the number of each atom equal on each side.

In this case, start by putting a 2 in front of the carbon dioxide to equal out the carbons. This will also add two more oxygens to the right-hand side

$C_2H_5OH + O_2 \longrightarrow 2CO_2 + H_2O$

left-hand side	right-hand side
C 2	C ~~1~~ 2
H 6	H 2
O 3	O ~~3~~ 5

Put a 3 in front of the water to balance the hydrogens. Remember to add to the oxygens again

$C_2H_5OH + O_2 \longrightarrow 2CO_2 + 3H_2O$

left-hand side	right-hand side
C 2	C ~~1~~ 2
H 6	H ~~2~~ 6
O 3	O ~~3~~ ~~5~~ 7

The carbons and hydrogens are now equal on both sides so you must multiply the oxygens on the left-hand side to finish balancing the equation

$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$

left-hand side	right-hand side
C 2	C ~~1~~ 2
H 6	H ~~2~~ 6
O ~~3~~ 7	O ~~3~~ ~~5~~ 7

This equation is now balanced.

PAUSE POINT Write a balanced equation for the following reaction:
butanol (C_4H_9OH) + water → carbon dioxide + water.

Hint Remember you can only change the number of moles of each substance, you cannot change the formula.

Extend Now write a balanced equation for:
magnesium carbonate + hydrochloric acid → magnesium chloride + water + carbon dioxide.

Assessment practice 1.3

Write balanced equations for the following reactions.

1. methane (CH_4) + oxygen → carbon dioxide + water
2. calcium carbonate ($CaCo_3$) + hydrochloric acid (HCl) → calcium chloride + carbon dioxide + water
3. calcium hydroxide (CaOH) + hydrochloric acid → calcium chloride + water

Moles, molar masses and molarities

Chemical equations allow you to work out the masses of the reactants you need to use in order to get a specific mass of product. Chemists never use one molecule of a substance because that would be too small. Even 0.1 g of hydrochloric acid will contain millions of molecules of the acid. These numbers are very big and difficult to work with so chemists use a quantity called a **mole** with the symbol mol.

Do not let the idea of a mole confuse you. It is just a number. One mole of a chemical means there are 6.023×10^{23} particles (Avogadro's constant).

6.023×10^{23} is a number is standard form. This is a simple way of showing a very large number. The 10^3 is how you would write 1000 in standard form. The 10^3 means that if you write the number out in full, it will have 3 zeroes at the end. So 6.023×10^{23} is a simple way to write 6023 with 20 zeroes at the end.

A mole is the amount of a substance which has the same number of particles as there are atoms in 12 g of carbon-12.

So one mole of carbon dioxide has the same number of particles as one mole of gold. The **molar mass** of a substance is equal to the mass of one mole of a substance.

It is useful to be able to convert masses into moles and moles into masses.

Mass (g) = molar mass × number of moles

- The relative atomic mass (*Ar*) of an element on the periodic table tells you how much mass there is in one mole of the element. The relative atomic mass is the average mass of an atom of an element compared to one twelfth of the mass of an atom of carbon–12. The relative atomic mass of hydrogen is 1.0. The relative atomic mass of oxygen is 16.0.
- The relative formula mass is the sum of all the relative atomic masses of all the atoms in the empirical formula (simplest formula) of a compound (*Mr*).
  The relative formula mass of water, H_2O, is $(1 \times 2) + 16 = 18$
  Relative atomic and formula masses do not have any units as they are only relative to carbon–12.

Key terms

Mole – a unit of substance equivalent to the number of atoms in 12 g of carbon-12. 1 mole of a compound has a mass equal to its relative atomic mass expressed in grams.

Molar mass – the mass of one mole of a substance.

Assessment practice 1.4

What is the relative formula mass for these molecules?

1 CO_2

2 NaOH

3 H_2SO_4

4 $Ca(OH)_2$

5 Fe_2O_3

The following worked examples show how to convert masses to moles.

Worked Example

1 What is the number of moles in 136.5 g of potassium?

Number of moles of an element = mass/*Ar*

For potassium A_r =39

$$\text{Number of moles} = \frac{136.5}{39}$$

$$= 3.5 \text{ moles}$$

2 What is the number of moles in 20 g of sodium hydroxide, NaOH?

Number of moles = mass/M_r

For sodium hydroxide M_r = 23 + 16 + 1 = 40

$$\text{Number of moles} = \frac{20}{40}$$

$$= 0.5 \text{ moles}$$

Empirical formula

This shows the ratio between elements in a chemical compound. It is useful when discussing giant structures such a sodium chloride. The empirical formula of a compound can be calculated from the masses of each element in the compound. These masses are worked out through experimental analysis of the compound.

Step by step: Empirical formula

3 Steps

1 Divide the mass of each element present in the compound by its molar mass to get its molar ratio.

2 Divide the answer for each element by the smallest molar ratio calculated. This gives you a ratio of 1:x for each element present.

3 If the answers are not all whole numbers, multiply them all by the same number to get whole numbers. e.g. if the ratio is 1:1.5:3 then multiplying all the numbers by 2 will give you an answer with all whole numbers 2:3:6

Molecular formula

Molecular formulae are used for simple molecules. To work out the molecular formula you need to know the empirical formula and the relative molecular mass. e.g. a compound has the empirical formula CH_4. This has an empirical formula mass of 12 + (1 × 2) It has a relative molecular mass of 42. To work out its molecular formula you first divide its relative molecular mass by the empirical mass. 42/14 = 3 You write out the formula multiplying each part of the CH_2 unit by 3. This gives C_3H_6. This is the molecular formula.

Reacting quantities

When carrying out **titrations**, a chemist has to use solutions of a known concentration. These are called **standard solutions**. They have been prepared and tested to ensure they are of the specific concentration needed.

The number of moles of **solute** in a given volume of **solvent** tells you how concentrated the solution is.

When 1 mole of solute is dissolved in 1 cubic decimetre of solution, its concentration is written as:

$1\ \text{mol dm}^{-3}$.

This can be written as 1M for short. This is the molarity of the solution.

I mole of HCl has a mass of 1 + 35.5 = 36.5 g

36.5 g of HCl in 1 dm^3 of solution has a concentration of 1 mol dm^{-3} or 1M or 36.5 g dm^{-3}.

Key terms

Titration – a method of volumetric analysis used to calculate the concentration of a solution.

Solution – a liquid mixture where a solute is dissolved in a solvent

Standard solution – a solution of known concentration used in volumetric analysis.

Solute – the substance dissolved in a solvent to form a solution.

Solvent - a liquid that dissolves another substance.

Worked Example

1 How many moles of hydrochloric acid are there in 100 cm^3 of 1M hydrochloric acid solution?

Number of moles (N) = molarity (C) × volume of solution (V) (dm^3)

$$N = CV$$

The volume is given in cm^3 so this needs to be converted into dm^3 by dividing by 1000. (Remember 1 dm^3 = 1000 cm^3)

$$\text{number of moles} = \frac{100}{1000} \times 1$$

$$= 0.1\ \text{mol}$$

2 What is the concentration of a sample of sodium hydroxide solution if 10 dm^3 contains 0.5 mol?

Number of moles (N) = molarity (C) × volume of solution (V) (dm^3)

$$N = CV$$

$$0.5 = C \times 10$$

$$C = \frac{0.5}{10} = 0.05\text{M}$$

3 What volume in cm^3 of 2M sulfuric acid solution would you need to ensure you had a sample containing 0.05 mol?

Number of moles (N) = molarity (C) × volume of solution (V) (dm^3)

$$N = CV$$

$$0.05 = 2 \times V$$

$$V = \frac{0.05}{2} = 0.025\ dm^3$$

Multiply by 1000 to give answer in cm^3

$$0.025 \times 1000 = 25\ cm^3$$

4 Calculate the number of moles of HCl in 20 cm^3 of a 2 mol dm^{-3} solution of HCl(aq).

Convert 20 cm^3 to dm^3 by dividing by 1000.

$$\frac{20}{1000} = 0.02\ dm^3$$

Use the equation

Number of moles (N) = molarity (C) × volume of solution (V) (dm^3)

$$N = CV$$

$$0.02 \times 2 = 0.04 \text{ mol of HCl in solution}$$

Using a chemical equation to calculate the quantities of reactants and products

Chemical equations can be used to calculate the quantities of reactants and products. Here is an example. Calcium chloride can be produced by reacting calcium carbonate with hydrochloric acid. This is the equation for the reaction.

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Note that the equation includes state symbols. A solid substance is indicated by (s), a solution is indicated by (aq), a liquid is indicated by (l) and a gas is indicated by (g). The equation shows that one mole of calcium carbonate reacts with two moles of hydrochloric acid. One mole of calcium chloride is produced as well as one mole each of carbon dioxide and water. This is an example of **stoichiometry**.

Key term

Stiochiometry – involves using the relationships between the reactants and the products in a chemical reaction to work out how much product will be produced from given amounts of reactants.

Worked Example

1 Calculate the expected mass of calcium chloride produced when 50 g of calcium carbonate is reacted with excess hydrochloric acid.

A_r (H) = 1, A_r (C) =12, A_r (0) = 16, A_r (Cl) = 35.5, A_r (Ca) = 40

One mole of $CaCO_3$ produces one mole of $CaCl_2$.

You know this from the balanced equation

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$. This shows a one to one (1:1) ratio.

Add up the relative atomic masses for each compound.

40 +12 + (3 × 16) g = 100 g of $CaCO_3$ produces 40 + (35.5 × 2) g = 111 g of $CaCl_2$.

As one mole of $CaCO_3$ produces one mole of $CaCl_2$ then

100 g $CaCO_3$ produces 111 g $CaCl_2$.

In this case, only 50 g of $CaCO_3$ was used so

50 g $CaCO_3$ produces $\frac{111}{100} \times 50$ g $CaCl_2$.

50 g $CaCO_3$ produces 55.5 g $CaCl_2$.

You could say that only $\frac{1}{2}$ a mole of $CaCO_3$ was used so therefore only half the amount of $CaCl_2$ would be produced and this would give the same answer of 55.5 g.

This is the theoretical mass.

2 Calculate the expected mass of water if 10 g of oxygen is reacted with excess hydrogen.

A_r (H) = 1, A_r (O) = 16

Use a balanced equation to find out the ratio between oxygen and water.

$2H_2 + O_2 \rightarrow 2H_2O$

So one mole of oxygen gives 2 moles of water. This is a 1:2 ratio.

Add up the relative atomic masses for each substance. Remember there will be two lots of water.

2 × 16 g = 32 g of O_2 produces 2 × (2 × 1) +16 g = 36 g of H_2O

32 g O_2 produces 36 g H_2O

So

10 g O_2 produces $\frac{36}{32} \times 10$ g of H_2O

10 g O_2 produces 11.25 g H_2O

Assessment practice 1.5

Silver iodide is used in the manufacture of photographic paper.

Calculate the theoretical yield of silver iodide for 34 g of silver nitrate reacting with excess sodium iodide. The equation for the reaction is as follows.

$AgNO_3 (aq) + NaI (aq) \rightarrow AgI (s) + NaNO_3 (aq)$

A_r (N) = 14, A_r (0) = 16, A_r (Ag) = 108, A_r (I) = 127

Percentage yields

The **theoretical mass** is the amount of product you can produce in a reaction. In most reactions it is unlikely that the total amount of product possible is made.

Some may be lost in transferring product from one vessel to another. Some of the reactants or products may react with impurities. In **reversible reactions**, products react to become the reactants and so are not all extracted from the reaction system. Chemists need to know how efficient their reaction process is so they calculate the **percentage yield**.

The percentage yield is the actual mass compared to the theoretical mass. An efficient process would give a percentage yield as close to 100% as possible.

The formula for calculating percentage yield is:

$$\text{Percentage yield} = \frac{\text{actual number of moles}}{\text{expected number of moles}} \times 100\%$$

It can also be calculated as:

$$\text{Percentage yield} = \frac{\text{actual mass}}{\text{theoretical mass}} \times 100\%$$

The first step in working out percentage yield is to measure accurately the mass of product that you have obtained. How accurate your measurements are may depend on the equipment you have, but the mass should be measured to at least two decimal places. You should be able to use a top pan balance for this. If you are using small quantities, or if you want more accurate measurements, you may use a chemical balance which measures to 3 decimal places.

Once you have measured the mass of your product, you can work out how many moles you have produced. You can then divide this by the number of moles you were expecting to obtain and multiply by 100.

If you are using solutions, then you will need to calculate the number of moles for the volume of solution used. Calculating concentration times volume, *CV*, will give you the number of moles in the volume of solution used. This equation can be rearranged to find out what volume of a known concentration of solution is needed in a reaction.

Key terms

Theoretical mass – the expected amount of product from a reaction calculated from the balanced equation.

Percentage yield – the actual amount of yield worked out as a percentage of the theoretical yield.

Reversible reaction – a reaction where the reactants react to form products and the products simultaneously react to re form the reactants. s, for example, in NaCl.

Worked Example

1 When 50 g of calcium carbonate is reacted with excess hydrochloric acid solution to make calcium chloride, the theoretical yield is 55 g.When the reaction was carried out, only 44 g of calcium chloride was produced.

Calculate the percentage yield of calcium chloride.

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\text{Percentage yield} = \frac{44}{55} \times 100\%$$

Percentage yield is 80%

2 When I mole of oxygen reacts with excess hydrogen, 2 moles of water should be produced.

When this reaction was carried the actual yield was 1.8 moles of water.

Calculate the percentage yield.

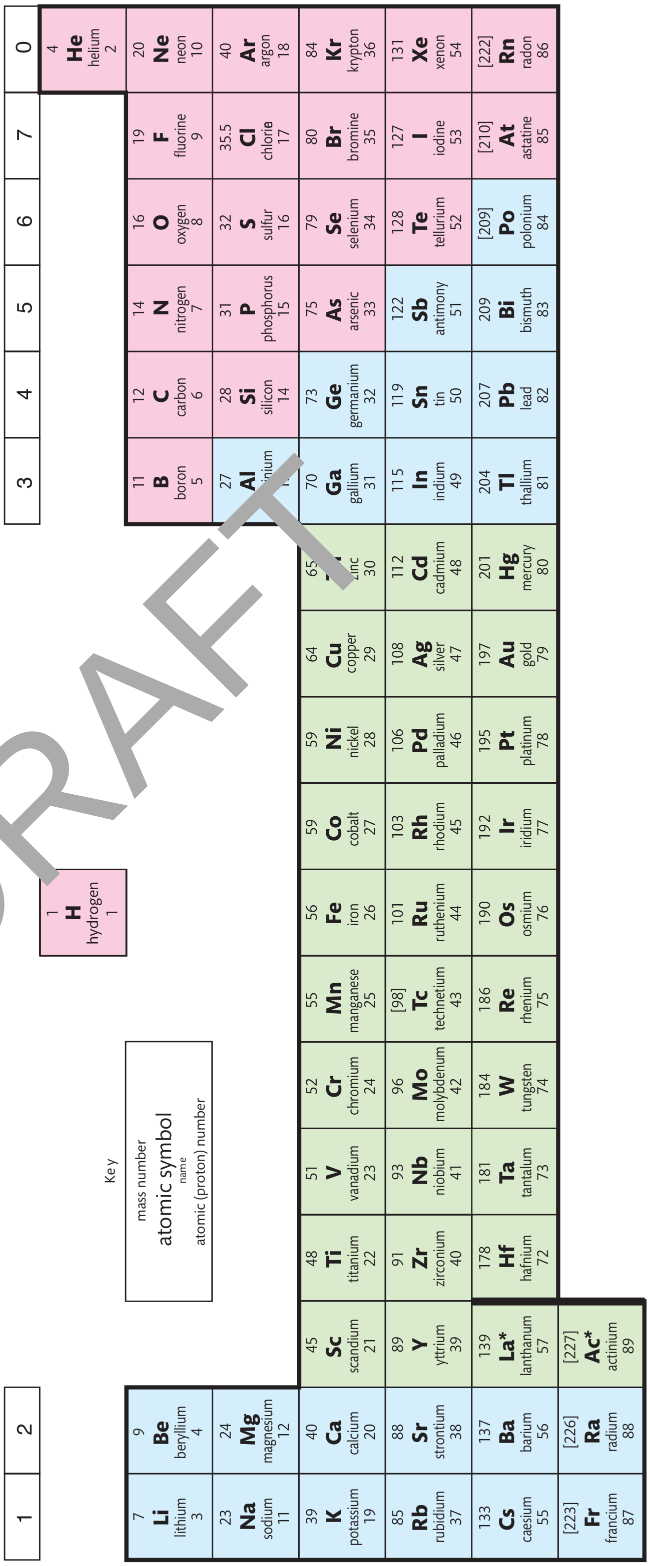
$$\text{Percentage yield} = \frac{\text{actual number of moles}}{\text{expected number of moles}} \times 100\%$$

$$\text{Percentage yield} = \frac{1.8}{2} \times 100\%$$

Percentage yield is 90%

A2 Production and uses of substances in relation to properties

The periodic table



1	2											3	4	5	6	7	0
						1 **H** hydrogen 1											4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chloria 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	64 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89															

Key

mass number
atomic symbol
name
atomic (proton) number

Figure 1.18: A section from the periodic table

Periods 1, 2, 3 and 4

The periodic table (see Figure 1.18) shows all the chemical elements arranged in order of increasing **atomic number**. Chemists can use it to predict how elements will behave, or what the physical or chemical properties of the element may be.

A laboratory technician needs to be very familiar with the periodic table. It is an information sheet on all the elements and their properties.

The elements on the periodic table are organised into groups (vertical columns) and periods (horizontal rows). Chemical properties are similar for elements in the same group. The atomic number increases as you move from left to right across a period. This is because each successive element has one more proton than the one before.

Key term

Atomic number – the number of protons in an atom. (This is the same as the number of electrons in the atom.)

Table 1.3: Characteristics of each period

Period	Characteristics
1	Contains hydrogen and helium. Both are gases. The electrons in these two elements fill the 1*s* orbital. Helium only has two electrons and, chemically, helium is unreactive. Hydrogen readily loses or gains an electron, and so can behave chemically as both a group 1 and a group 7 element. Hydrogen can form compounds with most elements and is the most abundant chemical element in the universe.
2	Contains eight elements, lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine and neon. The outer electrons in these elements fill the 2*s* and 2*p* orbitals. Nitrogen, oxygen and fluorine can all form diatomic molecules. Neon is a noble gas. Carbon is a giant molecular structure.
3	Contains eight elements, sodium, magnesium, aluminium, silicon, phosphorus, sulfur, chlorine and argon. The outer electrons in these elements fill the 3*s* and 3*p* orbitals.
4	Contains 18 elements, from potassium to krypton. The first row of the transition elements is in this period. The outer electrons on these elements fill the 4*s*, 4*p* and 3*d* orbitals.

Discussion

Look at the periodic table and write down five key features of the periodic table. Work in pairs and try to list the names of any groups in the periodic table as you can. Discuss any facts you know about the elements in the groups you have listed. These may be properties of the elements or trends within groups.

Groups – s block, p block, d block

The periodic table is also organised by element blocks. An element block is a set of elements in groups that are next to each other. Element blocks are named for the orbital the highest energy electrons are in for that set of elements. Groups 1 and 2 of the periodic table are in s block. Groups 3 to 7 and group 0 make up p block. This block contains all the non-metals except for hydrogen and helium. The transition metals are in the d block.

For example, carbon had electronic structure of $1s^2\ 2s^2\ 2p^2$. The highest energy electron in carbon is in a *p* orbital and therefore carbon is a p block element.

Assessment practice 1.6

Explain why calcium is an s block element.

PAUSE POINT Summarise what you have learned about the periodic table.

Hint Consider what you know about groups, periods and trends.

Extend Choose three elements in different areas of the table, and explain why their atomic structure and properties means they are in the position they are in.

Physical properties of elements

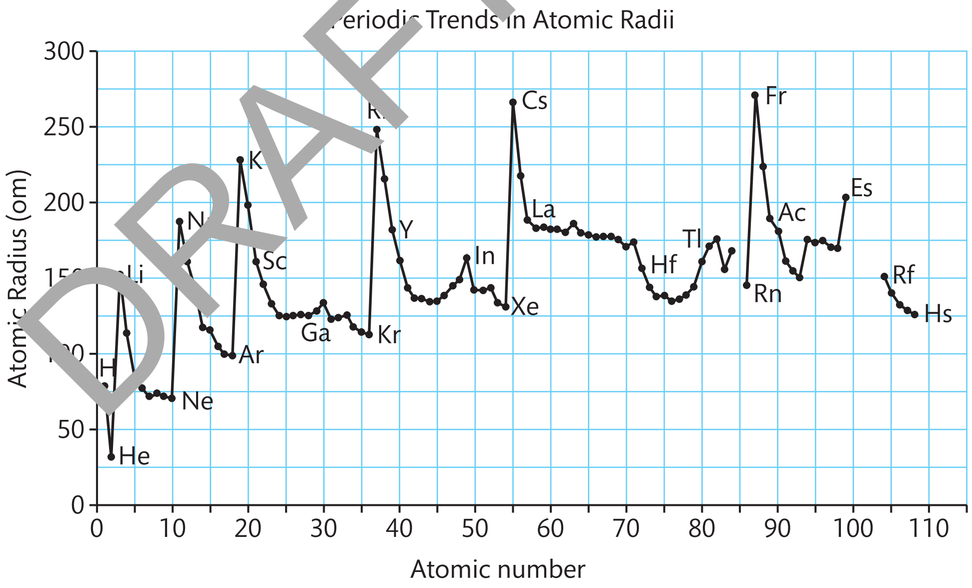
Atomic radius

The radius of an atom changes depending on what is around it. The only way to measure the radius is to measure the distance between the nuclei of two touching atoms and divide by two.

The atomic radius decreases across the period from left to right. Across the group, more protons and electrons are added. However, the extra electrons are added to the same s and p sub-shells and so the size does not increase. The extra protons increase nuclear charge. The increased nuclear charge attracts the extra electrons and pulls them closer to the nucleus. This leads to a decrease in atomic radius.

As you go down a group the atomic radii increases. This is because the extra electrons are added to additional shells and so the radius increases. Although nuclear charge increases, the number of inner shells increases and so the nuclear charge is shielded more. This means that the atomic radius increases.

The trend is slightly different for the transition metals. The atomic radii get slightly smaller as you go across the start of the transition metals but then the radii stay very similar. This is because the additional nuclear charge is balanced by the extra shielding by the 3*d* electrons of the outer 4*s* sub-shell.



▸ **Figure 1.19**

Ionic radius

The trends in ionic radius down a group follow a similar pattern to the trend for atomic radius down a group. This is because extra electrons are added to extra shells as you go down the group therefore giving a larger size.

Cations have a smaller radius than their corresponding atom. As you go across a period, the cations all have the same electronic structure. They are **isoelectronic**, therefore although number of electrons remains the same, the nuclear charge increases, for example, Na^+, Mg^{2+}, A^{3+}. However, the number of protons increases across the period. This pulls the electrons more strongly to the centre of the ion so the ionic radii of the cations decreases as you go across the period.

Key terms

Isoelectronic – having the same numbers of electrons.

Cations – ions with a positive charge.

Anions – ions with a negative charge.

Anions have a larger radius than the corresponding atom because there is more repulsion between the extra electrons. As you go across the period, the anions are all isoelectronic, for example, N_3^-, O_2^-, F^-. They have more electrons not fewer. The number of protons still increases as you go across the period whilst the number of shells and electrons stays the same so the ionic radius of the anions also decreases as you go across the period.

Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It increases as you go across a period. It decreases as you go down a group. This means that fluorine is the most electronegative element. The Group 0 gases such as argon that do not form bonds do not have electronegativity that can be reliably determined, because they do not form compounds/bonds.

Electronegativity depends on the number of protons in the nucleus, the distance from the nucleus of the bonding pair of electrons and how much shielding there is from inner electrons.

As you go across the period, the bonding pair of electrons will be shielded by the same number of electrons. However, the number of protons will increase, so the group 7 element will be more electronegative than the group 1 element.

As you go down a group, there is more shielding from inner electrons and the bonding pair of electrons are further from the nucleus. This adds up to less pull on the bonding pair from the positive charge of the nucleus and so electronegativity decreases.

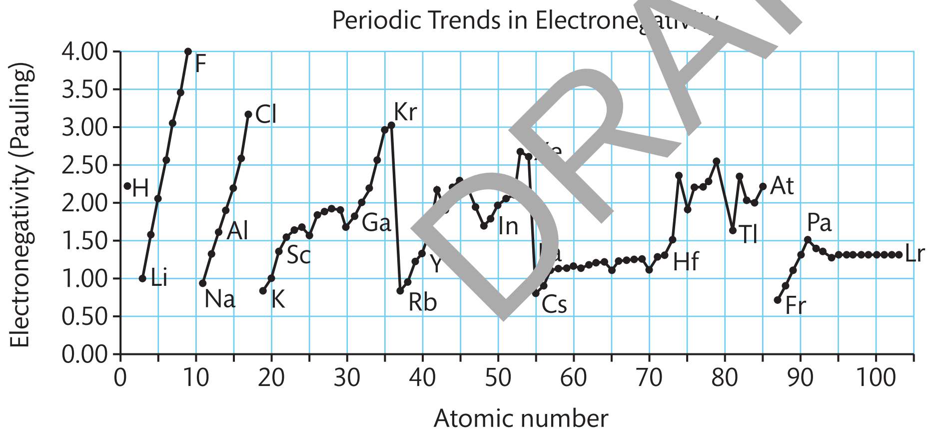


Figure 1.20

Trends in the periodic table are usually identified across periods or down groups. However, there are often similarities between elements that are diagonal to each other. For example, beryllium and aluminium have identical electronegativity. There is an increase across the period from group 2 to group 3 but then as electronegativity decreases down a group, this increase is balanced out. Other diagonal pairs also have similar electronegativity, for example, lithium and magnesium. These similarities mean they form similar bonds and may show similar chemistry.

PAUSE POINT Explain how electron affinity affects the reactivity of atoms.

Hint You will need to use the terms nuclear charge, shells, and shielding.

Extend Use this information to explain why potassium is a very reactive metal.

First ionisation energy and reasons for trends

First ionisation energy is the minimum energy needed for one mole of the outermost electrons to be removed from one mole of atoms in a gaseous state. One mole of positively charged ions is formed. First ionisation energies of the elements in a period show **periodicity**. There is an overall trend of first ionisation energy increasing across the period.

This trend is shown in the graph in Figure 1.21.

It takes more energy to remove an electron as you go across the period. This is because the number of protons increase across the period so the positive charge on the nucleus increases. This means that the force of attraction pulling on the outer electron increases. However, you can see there is not a steady increase in first ionisation energy. There is a pattern in the dips and increases for each period.

Across period 2, it dips at group 3 and group 6 elements. You can see the same pattern in period 3. Period four is a little different because it also contains the transition elements.

Key terms

Periodicity – the repeating pattern seen by the elements in the periodic table.

First ionisation energy – the energy needed for one mole of electrons to be removed from one gaseous atoms. For example, the equation shows potassium losing one electron to become a positive ion.

$K(g) \longrightarrow K^+(g) + e^-$.

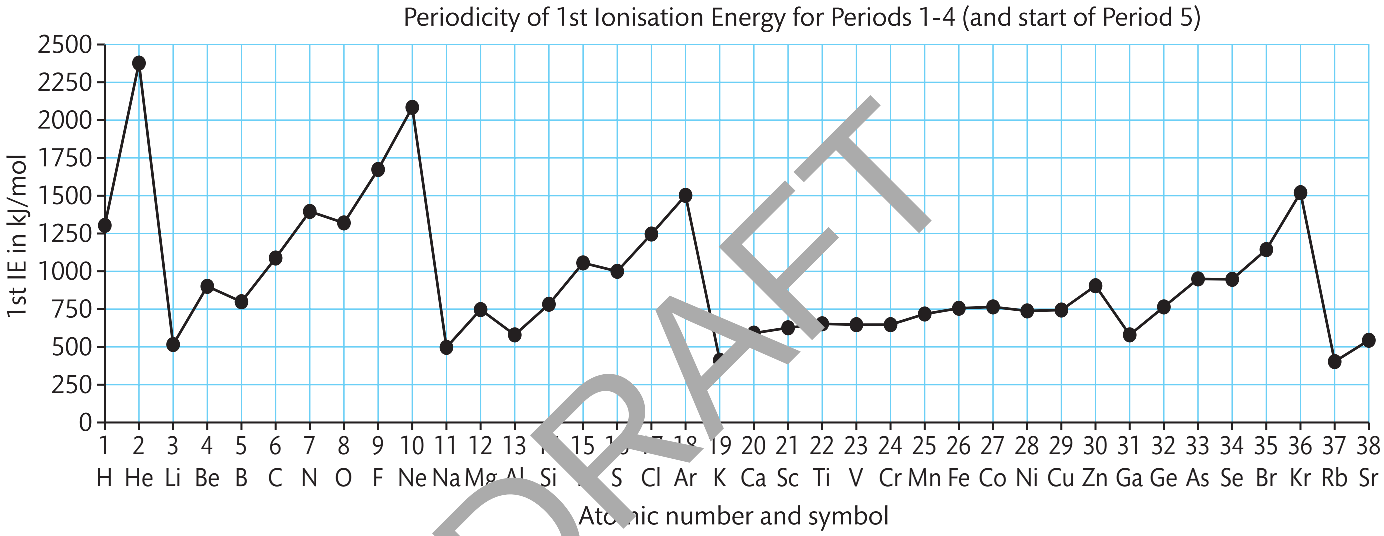


Figure 1.21: First ionisation energies of elements in periods 1–4

For periods 2 and 3, there is a pattern which suggests that electrons removed from the third energy level are arranged in different sub-levels. Across period 2, the first electrons are removed from the 2*s* sub-level. The value for beryllium is higher than for lithium because beryllium has one more proton in its nucleus. There is a decrease for boron where the electron is taken from the 2*p* sub-level. This is a higher energy level than the 2*s* sub-level and so the electron is easier to remove i.e. the 2*s* sub-shell shields the 2*p* sub-shell, making it easier to remove an electron.

Carbon and nitrogen show the expected increase because the electron removed from each element is in the same 2*p* sub-level. These electrons occupy orbital on their own, they are all unpaired. There is a second dip in the first ionisation energy at oxygen. Here the electron removed is also in the 2*p* sub-level, but it is paired with another electron in that level. The electrostatic repulsion between the two electrons in the orbital means that it is easier to remove this electron. The first ionisation energy increases then for fluorine and neon because they have increasing positive charge. A similar pattern is seen for period 3 where the electrons are removed from the 3*s* and 3*p* sub-levels.

In period 4, a similar pattern is seen for elements in groups 1 to 7 and 0. The transition elements' first ionisation energy does increase across the period, but only a little. Their outer electron always comes from a 4*s* sub-shell because this has a higher energy than a 4*d* sub-shell in the transition elements. As you go across the group, the number of protons increases, but so does the number of 3*d* electrons. These 3*d* electrons provide shielding and so cancel out the effect of the extra proton so ionisation energy only increases slightly across the d block.

As you go down a group, the first ionisation energy generally falls. You can also see this on the graph. Sodium has eight more protons than lithium so it may be expected that this increase in positive charge would increase the first ionisation energy. However, the outer electron in sodium is further away from the nucleus and has more shielding.

In lithium, the outer electron is attracted to 3 positive protons and is shielded by 2 negative electrons, so there is an overall attractive charge of +1. In sodium the outer electron is attracted to 11 positive protons and is shielded by 10 negative electrons so the overall attractive charge is also +1. The outer electron in sodium is further away from the nucleus and this lowers the effect of the +1 charge and so the first ionisation energy is lowered. This trend continues down the group and can also be observed in other groups such as groups 2 and 7.

Assessment practice 1.7

Evaluate the factors that affect first ionisation energies of elements in a period and in a group.

Electron affinity

Electron affinity can be simply defined as an atom's ability to gain an electron and become a negative ion. It is the change in energy (kJ mol^{-1}) of a neutral gaseous atom when an electron is added to the atom to form a negative ion. First electron affinity is when a –1 ion is formed. First electron affinities are negative. Table 1.4 gives the first electron affinities for group 7 elements.

Key term

Electron affinity – the charge in energy when one mole of a gaseous atom gains one mole of electrons to form a negative ion. For example, for oxygen.

$O(g) + e^- \longrightarrow O^-(g)$.

Table 1.4: First electron affinities

Element	First electron affinity kJ mol^{-1}
Fluorine	–328
Chlorine	–349
Bromine	–324
Iodine	–295

The negative sign shows that energy is released. The amount of energy released generally decreases as you go down group 7. Fluorine is an exception. Electron affinity indicates how strong the attraction is between the nucleus of an atom and the incoming electron. If this attraction is strong, more energy is released. Just as in first ionisation energy, number of protons (or nuclear charge), distance from nucleus and shielding all have an effect on electron affinity. As you go down the group the nuclear charge increases. However, there is also extra shielding from electrons as further shells are compressed. The further down the group you go, the further the distance the outer shell is from the positive pull of the nucleus and so the attraction becomes weaker. This means that less energy is released when the ion is formed.

Fluorine is a very small atom and this is why it does not follow this pattern. When fluorine gains an electron to become fluoride, this new electron is added to a region that is already full of electrons and so there is repulsion from these.

Group 6's electron affinities follow a similar pattern to that in group 7 with it decreasing as you go down the group. Oxygen does not follow this pattern. It has a lower electron affinity than sulfur for exactly the same reason as fluorine having a lower electron affinity than chlorine.

Overall group 6 elements have lower electron affinities than group 7 elements in the same period. This is because they have one less proton but the same amount of shielding. (Remember that elements in a period have the same number of electron shells as each other.)

Group 6 elements will also have a second electron affinity where the negative ion gains a second electron forming a charge of –2. It is the change in energy (kJ mol^{-1}) of a mole of gaseous –1 ions when an electron is added to the ion to form a –2 ion. The two negative charges will repel so this change in energy will be positive as energy will be needed to force an electron into the negative ion. Group 7 elements can also have a second electron affinity.

The first electron affinity for oxygen is –142 kJ mol^{-1}. The second electron affinity for oxygen is +844 kJ mol^{-1}. The high energy is needed to overcome the repulsion between negative electron and negative ion.

$$O(g) + e^- \rightarrow O^-(g)$$
$$O^-(g) + e^- \rightarrow O^{2-}(g)$$

Type of bonding in the element

The electronegativity of elements can be used to predict the type of bonding in a compound. Bonding is a spectrum from ionic to covalent bonding with most compounds sitting somewhere between the two. It is rare to have a wholly ionic or wholly covalent compound. In a hydrogen molecule, both hydrogen atoms have the same electronegativity. This means that they form a covalent bond that is not polar. When hydrogen bonds with fluorine to make hydrogen fluoride, a polar covalent molecule is formed. This is because fluorine has a high electronegativity and so attracts the bonding pair. This gives the fluorine atom a positive charge and the hydrogen atom a negative charge.

You cannot directly measure electronegativity of an element. The chemist Linus Pauling produced a scale that gives a relative value for the elements and this allows you to predict how ionic a covalent bond will be.

Electronegativities for elements in periods 1 to 3

H 2.1

Li 1.0 Be 1.5 B 2.0 C 2.5 N 3.0 O 3.5 F 4.0

Na 0.9 Mg 1.2 Al 1.5 S 1.8 P 2.1 S 2.5 Cl 3.0

If the difference between the electronegativities of the elements forming the bonds is low, then the covalent bond will be less polar than when the difference between the electronegativities is high. As the difference increases, the covalent bond will become more polar. If the difference is very large, then the bond becomes ionic.

Ionic bonds can also show polarity. The extent of the polarisation will depend on whether:

- either ion is highly charged
- the cation is relatively small
- the anion is relatively large.

A small cation that is highly charged will tend to draw electrons towards it. A large anion that is highly charged will have an electron cloud that is easily distorted. This means that some of the negative charge is shared with the cation. This gives the ionic bond some covalent characteristics.

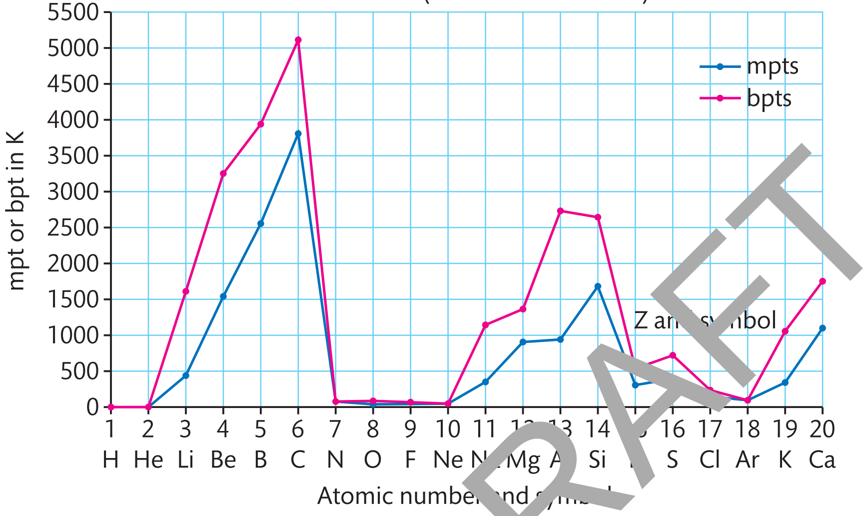
Period 2	Li	Be	B	C	N_2	O_2	F_2	Ne
Period 3	Na	Mg	Al	Si	P_4	S_8	Cl_2	Ar
Structure	giant metallic			giant covalent	simple molecular			
Forces	strong forces between positive ions and negative delocalised electrons			strong forces between atoms	weak intermolecular forces between molecules			
Bonding	metallic bonding			covalent	covalent bonding within molecules intermolecular bonding between molecules			

Trends: melting point and boiling point

The elements in the periodic table also show periodicity for melting and boiling points. Melting and boiling points depend on the strength of the forces between the atoms in an element. Going down group 1, the melting and boiling points decrease. This means that the forces of attraction get weaker. The melting and boiling points increase as you go down group 7. This means that the forces of attraction get stronger.

When an element melts, energy is used to overcome some of the attractive forces holding the atoms or molecules of the element together. When an element boils, most of the rest of the attractive forces are broken. The stronger the forces between the atoms, the higher the melting and boiling point will be. The melting and boiling points peak in the middle of period 2 and 3. The lowest values are found in group 0.

Periodicity (1) melting points and (2) boiling points for Periods 1-3 (and start of Period 4)



Period 2

As you go across groups 1 to 3, metals have increasing nuclear charge because they have increasing number of protons and increasing number of delocalised electrons and so have stronger metallic bonding. This means the melting and boiling points increase as you go across the metals in the period. Carbon, has giant covalent bonding forming a giant lattice structure with each atom bonding to 4 other carbon atoms. So its melting and boiling points are very high because it has strong covalent bonds that need a large amount of energy to break. The non-metals in groups 5 to 7 have small separate molecules and so have low melting points. There are only weak van der Waals forces that need to be overcome.

Period 3 follows the simple trend of period 2 with a few small exceptions. Sulfur, in group 6, has a higher melting and boiling point than the rest of the non-metals. This is because of the different size of the molecules of each of these elements. Phosphorus exists as P_4, sulfur as S_8, chlorine as Cl_2 molecules and argon as Ar atoms. The strength of the van der Waals forces increases as the size of the molecule increases. Therefore, because sulfur has the biggest molecule, it has the strongest van der Waals forces and so the highest melting and boiling points.

Assessment practice 1.8

Evaluate how type of bonding, intermolecular forces and molecule size affects the melting point in elements in period 3 and groups 2 and 6.

Physical properties of metals: electrical conductivity, thermal conductivity, malleability, ductility

Metallic bonding allows for electrical conductivity through a solid or liquid metal. The delocalised electrons carry the electric charge. Copper is an excellent conductor of electricity. In fact, it has the best conductivity of any metal except for silver and so is used for electrical cables and wires.

The delocalised electrons in metals also absorb heat energy which gives them kinetic energy. This energy is then transferred through the metal by these electrons. Metals are good thermal conductors. This makes many metals such as aluminium and copper useful for saucepans, heat sinks in computers, and radiators.

The structure of metals also explains why they can be **malleable** or **ductile**. The atoms in the layers are able to roll over each over. They can move to new positions without breaking the metallic bonds. Aluminium is very malleable, which along with its high thermal conductivity, make it useful for aluminium foil.

Key terms

Malleable – can be hammered into shape with breaking.

Ductile – can be hammered than or stretched into wires without breaking.

Research

Research the uses and applications of metals based on their properties. Can you find more examples of how each of the properties listed above make the metal useful?

Chemical properties of elements

The reactions between oxygen and metals are very important. How easily they react and the product they make can influence how a metal is used. For example, iron reacts very easily with oxygen and forms rust so it is often painted to protect it from oxygen in the air. Table 1.5 shows how elements react with oxygen.

Table 1.5: Products and reactivity of all period 2 and 3 elements with oxygen

Group	Element	Reactions with oxygen	Equations
1	Lithium Sodium	Rapid, burns with red flame. Metal oxide produced that forms a **basic solution** when dissolved in water. Very vigorous, burns with orange flame. Metal oxide produced that form basic solution when dissolved in water.	$4Li\,(s) + O_2(g) \rightarrow 2Li_2O(s)$ $4Na\,(s) + O_2(g) \rightarrow 2Na_2O(s)$ $2Na\,(s) + O_2(g) \rightarrow Na_2O_2(s)$
2	Beryllium and magnesium	Needs heat to react as do group 1 elements. Very vigorous reactions.	$2Be\,(s) + O_2(g) \rightarrow 2BeO(s)$
3	Aluminium	Vigorous at first. Rapidly forms a water insoluble coating of Al_20_3. This layer prevents the aluminium below from corroding and so makes aluminium an extremely useful material.	$4Al\,(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$ It is **amphoteric**.
4	Carbon Silicon	Forms slightly acidic oxides. Shows reaction with heat.	$C(s) + O_2(g) \rightarrow CO_2(g)$ $2C(g) + O_2(g) \rightarrow 2CO\,(g)$ – this is incomplete combustion $Si(s) + O_2(g) \rightarrow SiO_2(s)$ – weak acidic
5	Nitrogen Phosphorus	Forms a range of oxides with different **oxidation** states. A high temperature is need for these reactions to take place. Burns vigorously with a white flame.	It can produce NO, and NO_2 and N_2O_5. P_2O_6 if limited oxygen, P_4O_{10} if excess oxygen.
6	Oxygen Sulfur	In ozone layer O_2 and O_3 are **allotropes**. Two oxides form. Burns slowly with a blue flame.	$O + O_2 \rightarrow O_3$ $S + O_2 \rightarrow SO_2$ $2SO_2 + \frac{1}{2}O_2 \rightarrow 2SO_3$
7	Most halides react	Unstable oxides form.	Not usually formed by direct reaction.
0	Neon Argon	No reaction.	

As you move across periods 2 and 3, the general pattern for the oxides formed from left to right are 'ionic bonding to giant covalent structure to small covalent molecules'. The products change in nature from solids to gases and from alkaline to **amphoteric** to acid. This is due to the changes in bonding across the period.

Key terms

Oxidation – loss of electrons from an atom/ion.

Basic solution – a solution with a pH above 7.

Allotropes – two or more different physical forms that an element can exist in e.g. graphite and diamond are allotropes of carbon.

Amphoteric – substance that can act as both an acid and a base.

Products and reactivity of metals with oxygen, water, dilute hydrochloric acid and dilute sulfuric acid

Reactions with oxygen

Group 1 metals react rapidly with oxygen. Lithium, sodium and potassium are stored under oil to prevent contact with air due to this. The more reactive group 1 metals are usually stored in sealed glass tubes to ensure that no air or oxygen is present. The reactions of lithium and sodium with oxygen are shown in Table 1.[illegible]. Their oxides contain the simple ion O^{2-}. Sodium and potassium can also form the peroxide M_2O_2 containing the molecular ion $O_2{}^{2-}$. Potassium, rubidium and caesium ignite in air to form the super-oxides KO_2, RbO_2, and C_5O_2. These contain the molecular ion O_2^{2-}.

These more complicated ions are unstable near a small positive ion. The covalent bond between the two negative oxygen ions in $O_2{}^{2-}$ is weak. The electrons in the peroxide ion will be attracted to a positive ion but the positive ion can polarise the negative ion.

Lithium only has a +1 charge, but is it a small ion and so it has a high charge density. This causes the peroxide ion to break into an oxide and an oxygen atom. The super-oxide ions are even more unstable and are only stable in the presence of the larger, non-polarising ions at the bottom of group 1.

Group 2 metals tend to burn in oxygen or air to form metal oxides. This is the general equation.

$$2M + O_2 \rightarrow 2MO$$

Beryllium tends to form a coating of beryllium oxide. This makes it resistant to further oxidation.

Strontium and barium will also form peroxides. This is the general equation.

$$M + O_2 \rightarrow MO_2$$

Group 3 metals react with oxygen and this is the general equation for the reaction.

$$4M + 3\,O_2 \rightarrow 2M_2O_3$$

Thallium will also react to produce Tl_2O. Aluminium, like beryllium, also forms an outer coating of aluminium oxide. This means that it behaves as an unreactive metal.

The group 4 metals lead and tin can also produce oxides with the formula MO and MO_2.

When d block metals react with oxygen, the oxides are often brittle. Iron oxide is rust. Some d block metals become resistant to corrosion because they quickly form an unreactive outer oxide layer that prevents any more of the metal from reacting.

Titanium oxide is an example of this. The d block metals can form a range of oxides. Transition metals are much less reactive than group 1 and 2 metals in general.

Reactions with water

Group 1 metals are called alkali metals because when they react with water they produce a basic solution. Here is the general equation.

$$2M(s) + 2H_2O(l) \rightarrow 2M^+(aq) + 2OH^-(aq) + H_2(g)$$

They react violently with water. The reaction becomes more violent as you go down the group.

Group 2 metals also produce hydroxides in the reaction with water. Here is the general equation.

$$M(s) + 2H_2O(l) \rightarrow M(OH)_2(aq) + H_2(g)$$

Magnesium only reacts with steam, while the metals below magnesium will react increasingly easily with water. Beryllium does not react with water.

Group 3 metals are not very reactive with water. Aluminium does not appear to react at all due to its outer aluminium oxide layer. Group 4, 5 and 6 metals do not react with water. Transition metals react slowly with water and some do not react at all.

Reactions with dilute acids

Metals above copper in the reactivity series can react with dilute acids to form metal salts (an ionic compound formed from a neutralisation reaction) and hydrogen. For example, magnesium reacts with dilute hydrochloric acid to give magnesium chloride and hydrogen:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2.$$

It reacts with dilute sulfuric acid to give magnesium sulfate and hydrogen:

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2.$$

Sodium reacts with hydrochloric acid to form sodium chloride and hydrogen:

$$2Na + 2HCl \rightarrow 2NaCl + H_2.$$

This is a very violent reaction and is too dangerous to carry out in a school/college laboratory.

The reactions of calcium, strontium and barium with sulfuric acid are a little more complicated. This is because the sulfates of these metals are insoluble. They form a protective layer that prevents more of the metal reacting.

PAUSE POINT List the different types of metal reaction.

Hint Think about the chemical substances they react with and what products they form.

Extend Describe how the reactions change as you go down group I metals.

Positions of metals in the reactivity series in relation to position in the periodic table

The reactivity series (see Figure 1.22) is a list of metals in order of how reactive they are with oxygen, acids and water. The higher a metal is in the series, the more reactive it is. This is because it has a higher tendency to lose an electron and form a complete outer shell. The more reactive a metal is, the more difficult it is to extract from its ore and the more likely it is to be found in a compound.

The most reactive metals are in group 1, as reactivity decreases across the period. It also increases down the group, which means that francium is the most reactive metal and is at the top of the reactivity series. Most reactivity series do not list francium, as it is so radioactive and unstable that it is rarely seen uncombined. Most reactivity series have potassium at the top and gold and platinum at the bottom as they are so unreactive. It can be useful to place carbon and hydrogen. Carbon is present at it can extract/displace some metals from their compounds and ores as it explains why some metals cannot be extracted from their ores by reaction with carbon. Hydrogen is present as metals below hydrogen will not react with dilute acids or water to displace hydrogen. The order of reactivity is group 1, group 2, group 4, transition metals. The key reason is that, going across the period, the nuclear charge increases, therefore it is harder to lose an electron and react.

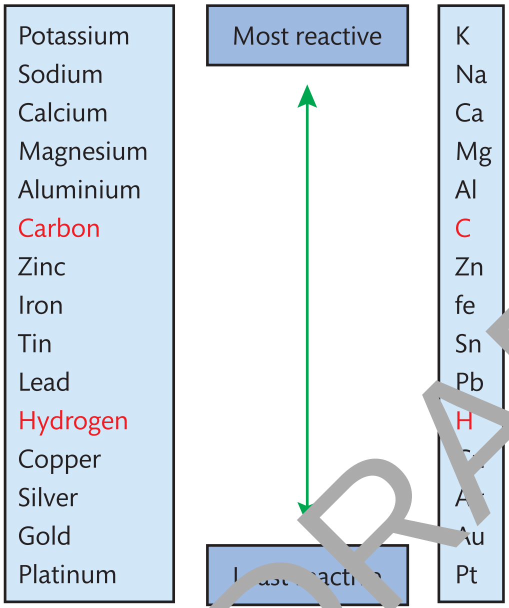


Figure 1.22: Reactivity series of metals

Oxidation and reduction

An atom becomes an ion when it loses or gains an electron or electrons. The term **redox** refers to the transfer of electrons that occurs during chemical reactions. When atoms of an element lose electrons, it is called oxidation. For example, $Mg \rightarrow Mg^{2+} + 2e^-$ (this is a half equation).

When electrons are gained, it is called **reduction**. For example, $\frac{1}{2}O_2 + 2e^- \rightarrow O^{2-}$

The two **half equations** together show the reaction between magnesium and oxygen:

$$Mg + \frac{1}{2}O_2 \rightarrow MgO$$

This process of electron transfer allows the reaction, oxidation and reduction occur simultaneously.

In reactions that make ionic compounds, it is easy to see where electrons are lost or gained. You cannot write half equations for reactions where covalent compounds are formed, for example, in the formation of water. This is where **oxidation states** come in. Oxidation states of atoms in a molecule of an element are always zero. So in O_2 both oxygen atoms have an oxidation state of 0. In water the oxygen is more electronegative than hydrogen so it has more power to attract electron density. So oxygen can be given the oxidation state of –2 as if it has taken two electrons in the covalent bound.

Key terms

Redox – the transfer of electrons during chemical reactions.

Reduction – when an atom/ion gains electrons. The phrase OIL RIG will help you remember the difference between oxidation and reduction: **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons).

Half equation – an equation that shows the loss or gain of electrons during a reaction

Oxidation state – the number assigned to an element in a chemical compound. It is a positive or negative number depending on how many electrons the element has lost or gained. (Also called oxidation number.)

Each hydrogen can be said to have 'lost' an electron so will have an oxidation state of +1. Compounds and molecules always have an overall oxidation state of 0, as the oxidation numbers of all the elements in the compound will add up to 0.

So oxidation occurs when the oxidation state of an atom increases and reduction occurs when the oxidation state of an atom decreases. In the formation of water, hydrogen has a change in oxidation state from 0 to +1 and oxygen changes from 0 to −2.

Key term

Redox reactions – reactions in which atoms have their oxidation state changed.

Not all reactions are **redox reactions**. If the oxidation states do not change, then the reaction is not a redox reaction. For example, the reaction between hydrochloric acid and sodium hydroxide,

$$HCl + NaOH \rightarrow NaCl + H_2O$$

oxidation state +1 −1 +1 −2 +1 +1 −1 2(H) −2

In this case, the oxidation states for each atom is the same in the reactants as in the products so this is not a redox reaction.

Step-by-step: Assigning oxidation states

5 Steps

1 The oxidation state of an atom in an element is always zero. For example, in sodium, Na, it is 0 and in O_2, oxygen, it is 0.

2 The oxidation state in an element or its ion is always its charge, including for polyatomic ions

3 The oxidation state of fluorine in a compound is always −1 as it is the most electronegative element.

4 The oxidation state of oxygen is nearly always −2 (except in peroxides and FO, where it is −1, +1).

5 The oxidation state of chlorine in a compound is usually -1 unless bonded with F or O.

6 The oxidation state of hydrogen is +1 unless bonded to a metal when it is −1. Group 1 metals are +1, group 2 metals are +2, aluminium is +3.

7 The sum of oxidation states in a compound is always 0. In polyatomic ions, the sum of the oxidation state of each element in the formula is the overall charge.

Variable oxidation states of transition metal ions

Transition metals have variable oxidation states due to their highest valence electrons being in the d sub-shell. This is a defining property of transition elements. When a period 1 transition metal loses electrons to form a positive ion, the 4*s* electrons are lost first, followed by the 3*d* electrons. The maximum oxidation state increases as you go along the period until manganese, which has a maximum oxidation state of +7 where 2 electrons are lost from the 4*s*, and 5 from the 3*d* orbitals. There is no simple rule to predict possible oxidation states, so you may want to learn some of the common states for the commonly used elements.

Scandium and zinc only have one oxidation state when in a compound. The others in the first period have two or more. For example, iron has possible oxidation states of +2 or +3 and these are written as Fe (II) and Fe (III).

Transition metals and their compounds have a large range of uses. Most of these uses are because of their variable oxidation state. Many are used as **catalysts**. For example, iron is used in the Haber process and platinum is used in catalytic converters in cars. Some transition metal compounds are also used as catalysts. For example, vanadium (V) oxide is used in the process for making sulfur dioxide (contact process). In the decomposition of hydrogen peroxide, manganese (IV) oxide is used as a catalyst. It is oxidised by the hydrogen peroxide to form manganese (VII) oxide. This then decomposes back to manganese (IV) oxide and oxygen. This means that the catalyst is ready to be reused.

Transition metals also have the metal properties discussed previously. They are good electrical and thermal conductors and they are malleable and ductile. This makes them useful when a conducting material is needed, as well as in structural materials. They have greater strength and are less likely to corrode than group 1 metals and so are often added to alloys to improve the properties of the material.

Key term

Catalysts – substances that increase the rate of a chemical reaction but are unchanged at the end of the reaction.

Theory into practice

Transition metals are extremely important in the chemical industry. They are used as a catalyst in a range of manufacturing processes.

1 Research one use of a transition metal as a catalyst.
2 Describe how the transition metal is used in the process.
3 Explain how the transition metal acts as a catalyst to increase productivity.

PAUSE POINT Work out the oxidation state of chlorine in the following compounds

HCl, HClO, NaClO$_2$, KClO$_3$, [illegible], Cl$_2$O$_7$.

Hint Think about the number of electrons lost or gained. Follow the rules above.

Extend In the following equations, which elements have been reduced, oxidised, neither oxidised nor reduced?

$2KBr + Cl_2 \rightarrow 2KCl + Br_2$

$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

Displacement reactions of metals/halogens

A metal will displace a less reactive metal in a metal salt solution. For example, when iron is added to blue copper sulfate solution, the solution will lose its colour as iron sulfate is formed. You will also see a pink-brown metal forming as copper is displaced out of the solution. Here is the equation for the reaction.

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

oxidation state 0 +2 +6 4 (–2) +2 +6 4 (–2) 0

The iron has been oxidised and the copper has been reduced. Iron is more reactive than copper as can be seen in the reactivity series. You can predict which metals will displace which from their salts by using the reactivity series.

Table 1.6

	Magnesium	**Zinc**	**Iron**	**Copper**
Magnesium sulfate	No reaction	No reaction	No reaction	No reaction
Zinc sulfate	displacement	No reaction	No reaction	No reaction
Iron sulfate	displacement	displacement	displacement	displacement
Copper sulfate	displacement	displacement	displacement	displacement

Key term

Oxidising agents – substances that withdraw electrons from other atoms or ions.

Halogens are **oxidising agents** which means they withdraw electrons from another atom or ion. The oxidising power of a halogen decreases as you go down group 7. If chlorine reacts with potassium bromide, then the bromine will be displaced and potassium chloride will form.

$$CL_2 + 2KBr \longrightarrow 2KCl + B_2$$

This is because chlorine is a stronger oxidising agent than bromine and so withdraws an electron from the bromide ion.

Table 1.7

	Chlorine	**Bromine**	**Iodine**
Potassium chloride	No reaction	No reaction	No reaction
Potassium bromide	displacement	No reaction	No reaction
Potassium iodide	displacement	displacement	No reaction

Uses and applications of substances produced within this unit

Knowing the chemical and physical properties of elements and the compounds is important to chemists when they are researching for a substance for a specific industrial application. The substances in this unit have a range of applications. Some are given below, but it would be useful for you to research more applications whenever you discuss or investigate a chemical substance.

- Metal and non-metal oxides have a range of applications. For example, magnesium oxide is used as a starter material for industrial processes such as producing magnesium alloys or fibreglass.
- Metal salts are used to make the colours in fireworks.
- Sodium chloride is used for many different manufacturing processes such as making glass, paper and rubber as well as being used in water softening systems.
- Sulfates are used in detergents.
- Copper sulfate is used in water treatment to kill algae.

Discussion

The uses listed here are not exhaustive. Research further uses of the products in this chapter. Can you link the uses to their physical and chemical properties? Discuss this with your group.

PAUSE POINT How do the physical properties of elements change across the periods?

Hint Think about reasons for trends in ionisation energy across periods 2–4 and down groups 1, 2 and 7.

Extend Research how physical properties and ionisation energies may affect reactions in industry. You may want to pick one specific reaction.

Further reading and resources

www.rsc.org The website of the Royal Society of Chemists.

www.sciencebuddies.org A website giving hands on science projects.

www.virtlab.com A series of hands on experiments and demonstrations in Chemistry.

Getting started

Biology is the study of living organisms. Cells are found in all living organisms. They are the fundamental unit of structure and function in all living organisms. From single prokaryotic cells, to the millions of cells that make up animals and plants, cells are vital to life. It is essential that you understand the structure and function of cells in order for you to understand the fundamental concept of biology. See if you can list parts of a plant and animal cell. When you have completed this unit, you should be able to add more to your list.

B Structure and function of cells and tissues

B1 Cell structure and function

In this section you will learn about cell theory, microscopy, and the ultrastructure, and function of animal, plant and **prokaryotic cell**. You will also use micrographs to identify cell organelles and carry out **magnification** calculations

Key terms

Magnification – the number of times larger the image appears compared to the actual size of the object being viewed.

Prokaryotic cell– A cell with no true nucleus or nuclear membrane.

Cell theory

Cell theory is the concept that cells are the fundamental unit of structure, function and organisation in all living organisms. Cell theory states that both plant and animal tissue is composed of cells and that cells are the basic unit of life. It also states that cells can only develop from existing cells. In 1655, the English scientist Robert Hooke used an early light microscope to observe the structure of finely sliced cork. He made observations and described what he saw as 'cells'. This was the start of cell theory. Developments in microscopy meant cells could be observed in detail for the first time. Figure 1.23 shows the timeline for cell theory development.

1665:
Robert Hooke first described cells

1674–1683:
The first living cell was observed
Anton van Leeuwenhoek was the first person to observe bacteria and protoctista from pond water samples, after developing powerful glass lenses.

1831:
Nucleus observed
Robert Brown an English botanist was the first to observe and describe the nucleus in a plant cell.

1839:
Universal cell theory
Matthias Schleiden suggested that all plant material is composed of cells. Jan Purkyne observed that animal tissue is composed of cells and the structure is similar to plant tissue. The scientist credited for the Universal Cell Theory is Theodor Schwann, a German physiologist. He proposed that 'all living things are composed of cells.

1852:
Evidence for the origin of new cells
Robert Remak observed cell division in animal cells. His findings were not accepted at the time, but in 1855 Rudolf Virchow published the findings as his own to show new cells form from existing ones.

1860:
Spontaneous generation disproved
Louis Pasteur demonstrated that bacteria will only grow in sterile nutrient broth after it has been exposed to air. This disproved the theory of spontaneous generation of cells.

Figure 1.23: Timeline of cell theory development

Key terms

Organelle – specialised structures found within a living cell.

Nucleus – an organelle found inside a cell which contains genetic information.

Mitochondria – an organelle where aeobic respiration takes place.

Chloroplast – a plant organelle where the stages of photosynthesis take place, Bind in plant cells, photosynthetic bacteria and algae.

Resolution – the ability to distinguish between objects that are close together.

Microscopy

Before microscopes were invented, people knew nothing about cells, sperm and bacteria or any other micro-organisms. Microscopes have given us the power to see these sorts of things in microscopic detail. With high power microscopes, it is possible to observe cell **organelles**.

A microscope is an instrument that is used to magnify objects that are too small to see with the naked eye. Using microscopes to see distinct cells that make up multi-cellular organisms allows us to observe how their structure relates to their function. When hospitals receive tissue samples, the cytology department need to determine if these samples are healthy or diseased. Histopathologists will analyse the samples using microscopy. It is important that they are able to recognise what they see and record observations accurately.

Light Microscopy

Light microscopes were first developed in the 16th century and continue to be improved and developed. Light microscopes use visible light and magnifying lenses to oberve small objects. There are limitations to using light microscopes because they have a lower magnification and **resolution** than other, more advanced, microscopes. The maximum magnification of a light microscpe is × 1500 and the maximum resolution is 200 nm. However, light microscopes do allow us to observe sub-cellular structures, known as organelles. For example, a light microscope can magnify a cell **nucleus, mitochondria**, and **chloroplasts** in plant cells. The image below shows a human cheek cell observed down a light microscope. You can clearly see the nucleus in the photograph.

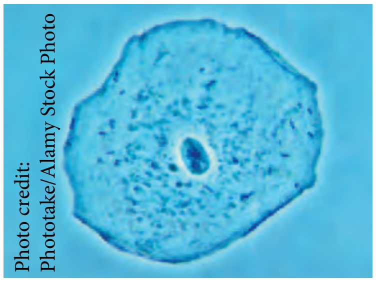


Photo credit: Phototake/Alamy Stock Photo

Light microscope used to observe human cheek cells – magnification

Electron microscopy

Electron microscopes were first developed in the 20th century. They use a beam of electrons in a vacuum with a wavelength of less than 1 nm to visualise the specimen. They allow much more detail of cell ultrastructure to be observed and produce images called electron micrographs, with a magnification of up to × 500 000 and higher resolution, as great as 0.1 nm. Samples are stained using methyleneblue for light microscopes and radioactive salts for electron microscopy.

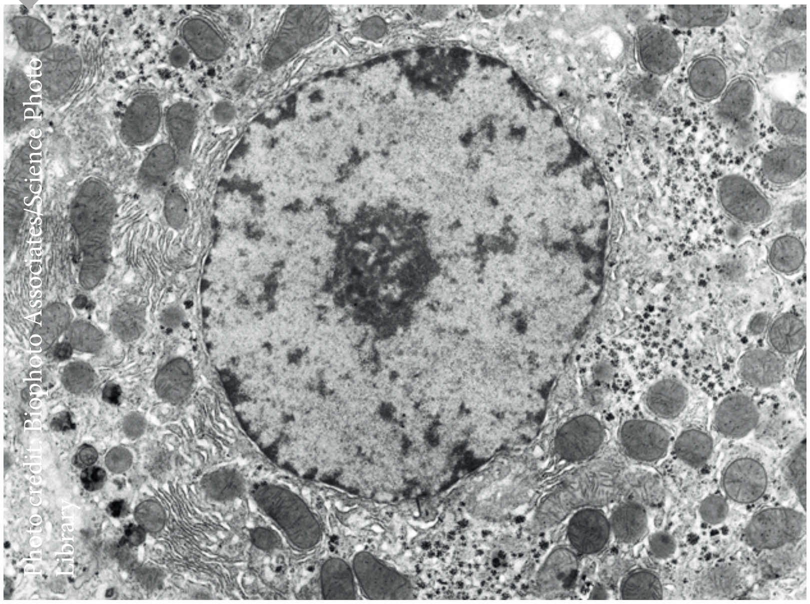


Photo credit: Biophoto Associates/Science Photo Library

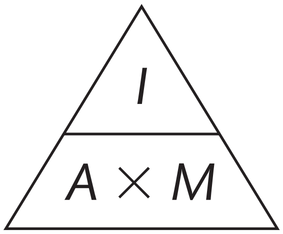
Electron micrograph of animal cell – magnification

Calculating magnification

We can use the equation below to work out magnification.

$$\text{Magnification (M)} = \frac{\text{size of image (I)}}{\text{actual size (A)}}$$

The size of the image refers to the length of the image when you measure it with a ruler. Ensure that you always measure in millimetres and convert the actual size to the same units that you have measured in. You will usually be given the magnification or the actual size in the exam question. You will therefore have one unknown and you can rearrange the equation to work out the unknown answer. Always include units in your answer and place your answer on the given line in the exam question. Finally, make sure you show your working out, including the equation above.



Figures 1.25 Use this triangle to help you to rearrange the magnification equation.

Worked Example

Calculate the magnification of the image. Use the equation above to work out the magnification.
Remember to convert all units to make them the same.
1000 nanometres (nm) = 1 micrometre (µm)
1000 micrometres (µm) = 1 mm

1000 mm = 1 m

1. Use your ruler to measure the size of the image in mm. The line measures 50 mm.
2. The image states that the actual size is 50 µm. You need to convert this to mm so they are both in the same units.
   $\frac{50}{1000} = 0.05$ mm
3. Magnification = 50/0.05
4. 50/0.05 = 1000
5. Magnification = × 1000

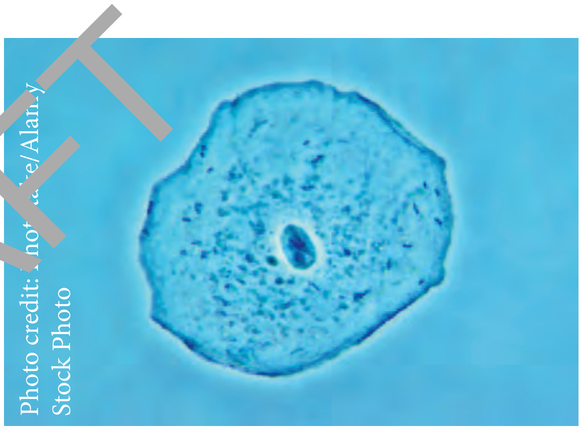


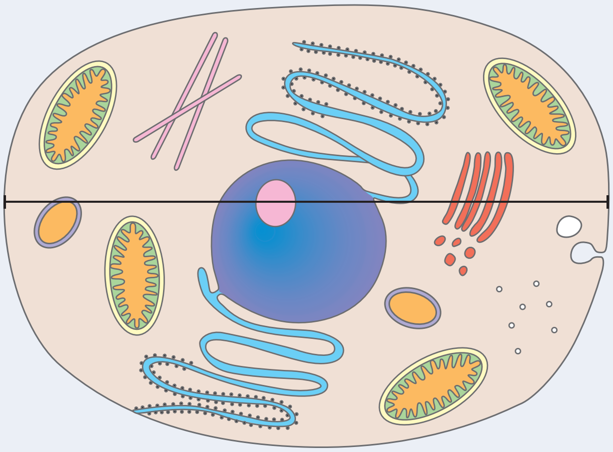
Photo credit: Phototake/Alamy Stock Photo

Calculating magnification

Assessment practice 1.9

Work out the magnification for the diagram.
The actual size of the image is 200 µm.

1. Use your ruler to measure the size of the image in mm.
2. The actual size is 200 µm. You need to convert this to mm so they are both in the same units.
3. Put both figures into the magnification equation and work out the magnification.



Calculating the magnification of a cell

Assessment practice 1.10

Work out the magnification of the nucleolus in the image.

1. Use your ruler to measure the size of the nucleolus in mm (size of image).
2. The actual size of the nucleolus is stated on the picture.
   You need to convert this to mm so they are both in the same units.
3. Put both figures into the magnification equation and work out the magnification.

Calculating the magnification of a nucleolus

PAUSE POINT Can you explain the concept of cell theory?

Hint Close the book and see if you can produce a time line with important dates in relation to the development of cell theory?

Extend Think about the differences between the light microscope and the electron microscope

Ultrastructure and function of organelles in cells

A cell is the basic unit of life. You will need to be able to recognise different types of cells when using a microscope, by observing the differences in their ultrastructure.

There are two types of cell.

- **Prokaryotic** cells are single-celled organisms. They are simple structures and do not have a nucleus or any membrane-bound organelles.
- **Eukaryotic** cells make up multi-cellular organisms such as plants and animals. They are complex cells with a nucleus and membrane-bound organelles.

Key term

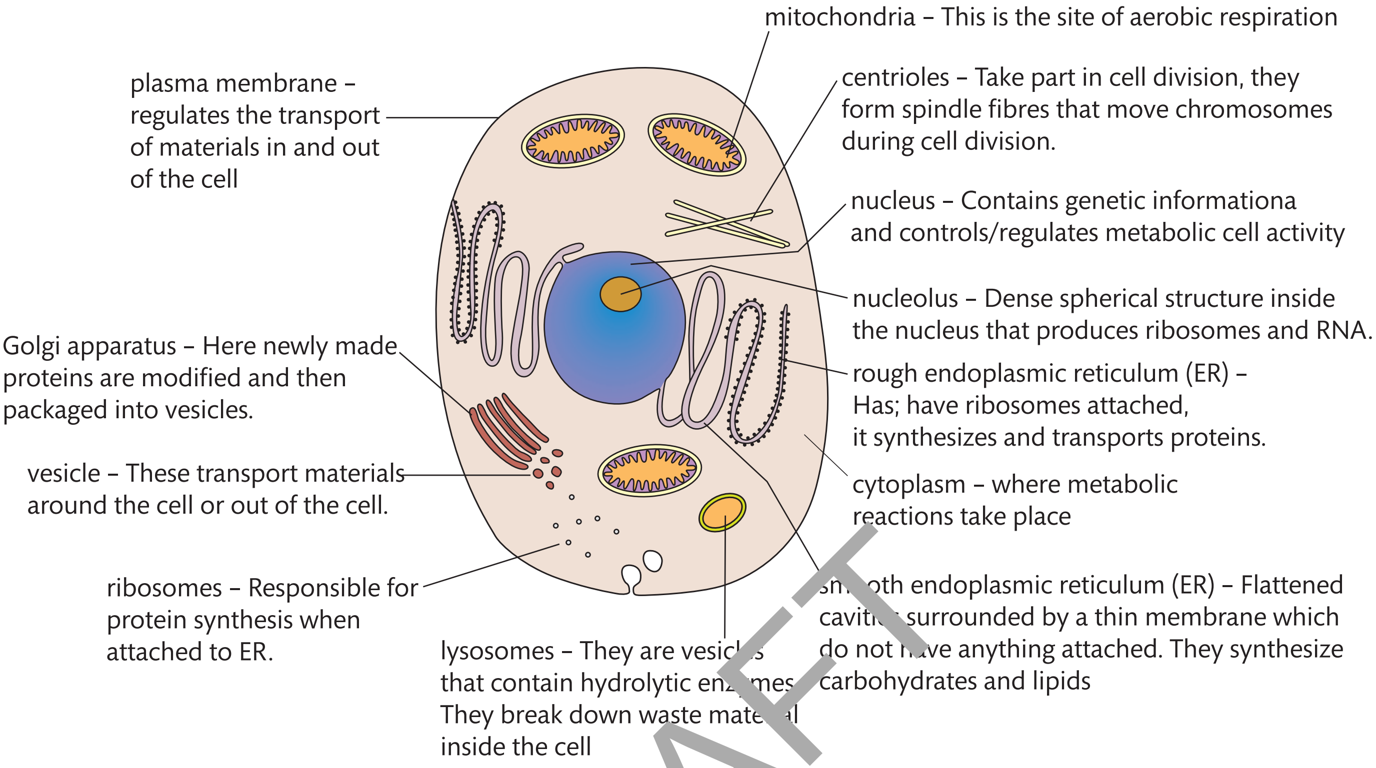
Eukaryotic – an organism that contains the genetic information as linear chromosomes within the nucleus of the cells and numerous specialised organelles.

Eukaryotic cells

Eukaryotic cells are approximately 10–100 μm and the ultrastructure can be seen using an electron microscope. Chemical reactions occur in the cytoplasm of a cell. The cell surface membrane or plasma membrane separates the cell cytoplasm from the external environment. Inside the cell cytoplasm there are a number of different structures called organelles. There are a number of organelles that are common in both plant and animal cells. You will study the structure of plant cells later in this unit.

Animal cell ultrastructure

Figure 1.24 shows the ultrastructure of an animal cell.



mitochondria – This is the site of aerobic respiration

plasma membrane – regulates the transport of materials in and out of the cell

centrioles – Take part in cell division, they form spindle fibres that move chromosomes during cell division.

nucleus – Contains genetic informationa and controls/regulates metabolic cell activity

nucleolus – Dense spherical structure inside the nucleus that produces ribosomes and RNA.

Golgi apparatus – Here newly made proteins are modified and then packaged into vesicles.

rough endoplasmic reticulum (ER) – Has; have ribosomes attached, it synthesizes and transports proteins.

vesicle – These transport materials around the cell or out of the cell.

cytoplasm – where metabolic reactions take place

ribosomes – Responsible for protein synthesis when attached to ER.

smooth endoplasmic reticulum (ER) – Flattened cavities surrounded by a thin membrane which do not have anything attached. They synthesize carbohydrates and lipids

lysosomes – They are vesicles that contain hydrolytic enzymes They break down waste material inside the cell

▶ **Figure 1.24:** Ultrastructure of an animal cell with organelles labelled

Table 1.8 describes the structure and function of organelles in an animal cell.

▶ **Table 1.8:** Structure and function of animal cell components

Organelle	Description of structure	Function
Plasma membrane	Composed of a phospholipid bilayer, with proteins embedded in the layer.	The membrane is selectively permeable and regulates the transport of materials into and out of the cell. Separates cell contents from the outside environment.
Cytoplasm	Cytoplasm is a thick, gelatinous, semi-transparent fluid.	The cytoplasm maintains cell shape and stores chemicals needed by the cell for metasolic reactions.
Nucleus	The nucleus is the largest organelle and is surrounded by a nuclear envelope. The envelope has nuclear pores which allow the movement of molecules through it. The nucleus contains chromatin.	The nucleus controls/regulates cellular activity and houses genetic material called chromatin, DNA and proteins, which comes the instruction for making proteins.
Nucleolus	Dense spherical structure in the middle of the nucleus.	The nucleolus makes RNA and ribosomes.
Rough endoplasmic reticulum (ER)	Network of membrane bound flattened sacs called cisternae studded with ribosomes.	Protein synthesis takes place on the ribosomes and the newly synthesised proteins are transported to the Golgi apparatus.
Smooth endoplasmic reticulum (ER)	Network of membrane bound flattened sacs called cisternae. No ribosomes..	Responsible for synthesis and transport of lipids and carbohydrates.
Golgi apparatus	A stack of membrane bound flattened sacs.	Newly made proteins are received here from the rough ER. The Golgi apparatus modifies them and then packages the proteins into vesicles to be transported to where they are needed.

Table 1.8: – *continued*

Organelle	Structure	Function
Vesicles	Small spherical membrane bound sacs with fluid inside.	Transport vesicles are used to transport materials inside the cell and secretory vesicles transport proteins that are to be released from the cell, to the cell surface membrane.
Lysosomes	Small spherical membrane bound sacs containing hydrolytic enzymes.	They break down waste material including old organelles.
Ribosomes	Tiny organelles attached to rough ER or free floating in the cell. They consist of two sub-units and they are not surrounded by a membrane.	Protein synthesis occurs at the ribosomes.
Mitochondria	They have two membranes. The inner membrane is highly folded to form cristae. The central part is called the matrix. They can be seen as long in shape or spherical depending on which angle the cell is cut at.	They are the site of the final stages of cellular respiration.
Centrioles	They are small tubes of protein fibres.	They form spindle fibres during cell division.

Function of animal cells

One of the key functions of a cell is to synthesise proteins for use inside the cell, to lead to cell multiplication and for secretion out of the cell for example, insulin. Proteins are synthesised on ribosomes attached to rough endoplasmic reticulum. The newly synthesised proteins are transported through the cisternae of the rough ER and packaged into vesicles. They are transported to the Golgi apparatus, where vesicles fuse with the surface of the Golgi apparatus and the proteins enter. It is here that the newly synthesised proteins are modified and then packaged into vesicles. Secretory vesicles will transport proteins that are to be released from the cell to the cell surface membrane. They will fuse with the membrane and release the protein by **exocytosis**.

Key term

Exocytosis – process of vesicles fusing with plasma membrane and secreting contents.

Plant cell ultrastructure

Plant cells have all the cellular components that are listed in the animal cell except centrioles (see Table 1.8). However, plant cells have additional structures and organelles because their main function is to produce carbohydrates during photosynthesis. (See Table 1.9.)

Table 1.9: Structure and function of plant cell components

Plant cell structure	Structure	Function
Cell wall	Made of cellulose forming a sieve like network.	Protects and supports each cell and the whole plant.
Chloroplast	Has a double membrane and is filled with a fluid called stroma. The inner membrane is a continuous network of flattened sacs called thylakoids. A stack of thylakoids is called a granum (grana is plural). Grana contain chlorophyll pigments.	Site of photosynthesis. Light energy is trapped by the chlorophyll and used to produce carbohydrate molecules from water and carbon dioxide.
Vacuole	Membrane bound sac in cytoplasm that contains cell sap.	Maintain turgor to ensure a rigid framework in the cell.
Tonoplast	The partially permeable membrane of the vacuole.	Selectively permeable to allow small molecules to pass through.
Amyloplast	A double membrane bound sac containing starch granules.	Responsible for the synthesis and storage of starch granules.
Plasmodesmata	Microscopic channels which cross the cell walls of plant cells.	Enable transport and communication between individual plant cells.
Pits	Pores in the cell walls of the xylem.	Allow water to enter and leave xylem vessels.

Figures 1.25 and 1.26 show the ultrastructure of a plant cell.

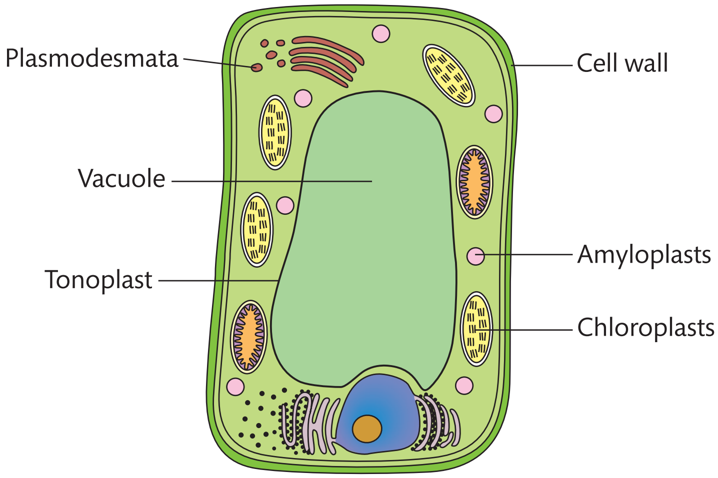


Figure 1.27: Ultrastructure of a plant cell

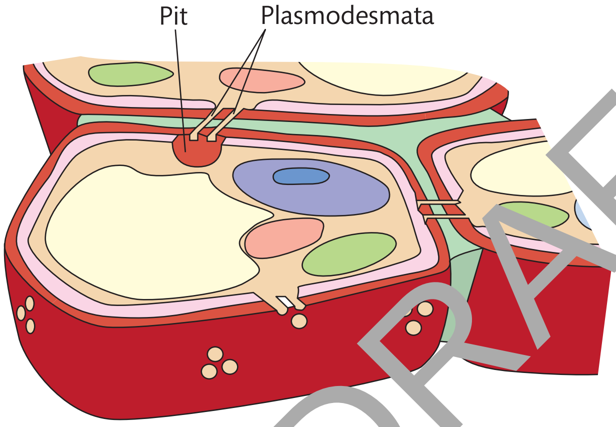


Image credit: adapted from www.phschool.com/science/biology_place/biocoach/plants/walls.html, Pearson Education, Inc., Used by permission of Pearson Education, Inc. All Rights Reserved

Figure 1.28: Ultrastructure of a plant cell with plasmodesmata and pits

PAUSE POINT Can you list all the organelles present in a eukaryotic cell?

Hint Think about the differences and similarities between a plant and animal cell.

Extend Find different images on the internet of plant and animal cells and identify the organelles.

Prokaryotic cell (bacteria) ultrastructure

Prokaryotes are single celled micro-organisms that are much smaller that eukaryotic cells. They are generally 1–5 μm in diameter. They are simple in structure, with no **membrane-bound organelles** and fewer organelles. Their **DNA** is not contained in a nucleus. Table 1.10 lists the parts of a prokaryotic cell, and their structures and functions.

Key terms

Membrane-bound organelles – organelles surrounded by a phospholipid membrane. For example, lysosomes and Golgi apparatus.

DNA – deoxyribonucleic acid, the hereditary material in cells.

Table 1.10: Structure and function of prokaryotic cell components

Organelle	Structure	Function
Cell wall	Prokaryotic cells are surrounded by a cell wall made of peptidoglycan.	Protects and supports each cell.
Capsule	Slippery layer outside the cell wall of some species of bacteria.	Protects the cell and prevents dessication.
Ribosomes	Smaller than ribosomes found in eukaryotic cells. They consist of two sub-units and they are not surrounded by a membrane.	Protein synthesis occurs at the ribosomes.
Nucleoid	The nucleoid (meaning nucleus-like) is the irregularly-shaped region that holds nuclear material without a nuclear membrane and where the genetic material is localized. The DNA forms one circular chromosome.	The nucleoid is the region where generic information can be found. and controls cellular activity.
Plasmid	Small circular loops of DNA.	Plasmids carry genes that may benefit the survival of the organism.

Figure 1.27 shows the ultrastructure of a prokaryotic cell with flagello.

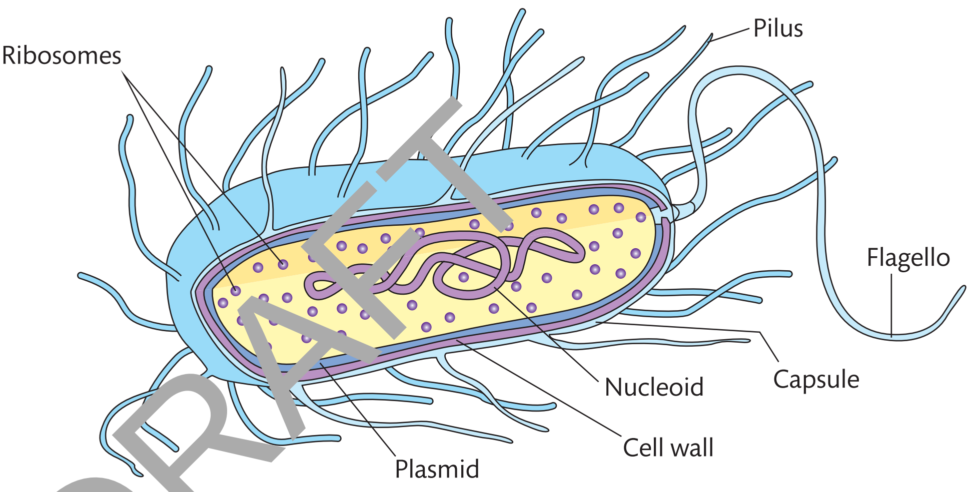


Image credit: after http://www.jlhahnconsulting.com/prepsteps1/bioB2.html, McGraw-Hill

Figure 1.27: Ultrastructure of a prokaryotic cell

Ribosome size is determined by their ability to form sediment in a solution. Eukaryotic ribosomes are determined as 80S whereas prokaryotic cell ribosomes are smaller and are 70S**.**

Function of bacterial cells

Bacterial cells, like many other animals and plants, produce and secrete toxins that have an effect on other organisms. DNA is free in the cytoplasm of a prokaryotic cell in the area called the nucleoid. A section of DNA containing a genetic code for a metabolve unwinds and hydrogen bonds break. RNA nucleotides line up (**complementary base pairing**). Messenger **RNA** is formed. This process is known an transcription. The next process is the production of the bacterial protein. This is called translation and it occurs at the ribosomes. Transcription and translation can occur simultaneously because the genetic material is free in the nucleoid surrounded by ribosomes. The newly made protein toxin is moved to the surface membrane ready to be secreted to cause infection. Note that many bacteria are beneficial to humans and to eukaryotes.

Key terms

Complimentary base pairing – the way in which nitrogenous bases in DNA pair with each other. Adenine (A) always bonds with Thymine (T) (or uracil (U) in mRNA) and Guanine always bond with Cytosine.

RNA – ribonucleic acid, a molecule with long chains of nucleotides.

Link

Go to *Unit 11: Genetics and Genetic Engineering Learning aim A* to find more information about DNA base pairing rule and about transcription and translation.

Classifying bacteria as Gram positive or Gram negative

It is important that microbiologists can correctly identify bacteria that cause infections to enable them to decide the most effective treatment.

Gram stain

Hans Christian Gram, a Danish microbiologist, developed a staining technique to distinguish between two groups of bacteria:

- Gram positive
- Gram negative.

Both types of bacteria have different cell wall structures and respond differently to antibiotics. Penicillin stops the synthesis of the cell wall on growing Gram-positive bacteria, but it does not have the same effect on Gram-negative bacteria. Gram-negative bacteria have a thinner cell wall and two lipid membranes.

During the staining technique, two stains are added to the bacterial smear: crystal violet and safranin. If you see a purple stain when observing the smear under a microscope it shows that Gram-positive bacteria are present. If the smear has retained the pink safranin stain, this shows that Gram-negative bacteria are present. This is because their thinner cell walls and lipid membranes allow ethanol (applied during the method) to wash off all the crystal violet purple stain and to then retain the pink safranin stain.

PAUSE POINT Do you know the functions of both animal and bacterial cells?

Hint Think about the products made by each cell.

Extend Can you think what might go wrong if these processes were interrupted?

Assessment practice 1.11

Produce revision cards on all the organelles you would find in an animal cell, plant cell and bacterial cell.

- On one side, write the name of the organelle.
- On the other side, write its function.

Assessment practice 1.12

Copy and complete the following table to show **three** ways in which prokaryotic and eukaryotic organisms **differ** in the **structure** of their cells.

Prokaryotic	Eukaryotic
1	....
2	....
3..	....

Assessment practice 1.13

Copy and complete the table below by putting a tick or a cross in the box to indicate whether the feature is present or absent.

	Cell type		
Feature	**Plant**	**Animal**	**Bacteria**
Mitochondria			
Chloroplast			
Cellulose cell wall			
Nucleus			
Ribosome			

B2 Cell Specialisation

Many organisms are multi-cellular, meaning that they are made from billions of cells. It is important that cells within these organisms become specialised for different roles with particular functions. Multi-cellular organisms in higher animals and higher plants are organised as follows:

- specialised cells
- tissues
- organs
- organ systems
- organism.

Cell specialisation: structure and function

Palisade mesophyll cell

Key term

Turgor – rigidity of plant cells due to pressure of cell contents on the cell wall.

Palisade mesophyll cells found in leaves are rectangular box shaped cells that contain chloroplasts (see Figure 1.28). The chloroplasts are able to absorb a large amount of light for photosynthesis. They also move around in the cytoplasm in order to maximise the amount of light absorbed. These cells are closely packed together and form a continuous layer in the leaf. Palisade cells are surrounded by a plasma membrane and a cell wall made of cellulose. This helps to protect the cell and keep it rigid. They also have a large vacuole to maintain **turgor** pressure (the plasma membrane pushes against the cell wall of plant to maintain its rigid structure).

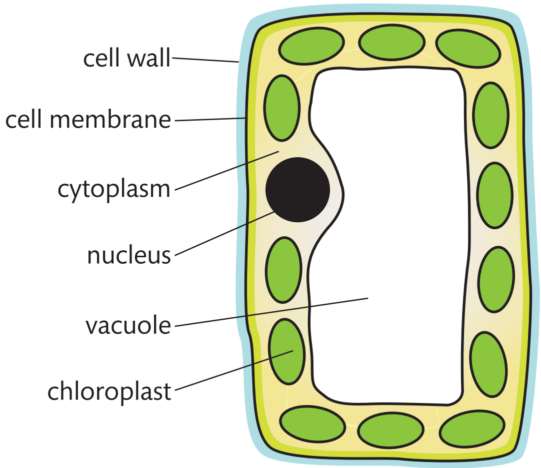


Figure 1.28: Basic structure of a palisade mesophyll cell found in leaves

Root hair cell

These cells are found at a plant's roots, near the growing tip (see Figure 1.29). They have long hair-like extensions called root hairs. The root hairs increase the surface area of the cell to maximise the movement of water and minerals from the soil into the plant root. The cells have thin cellulose walls and a vacuole containing cell sap with a low **water potential**. This encourages the movement of water into the cell.

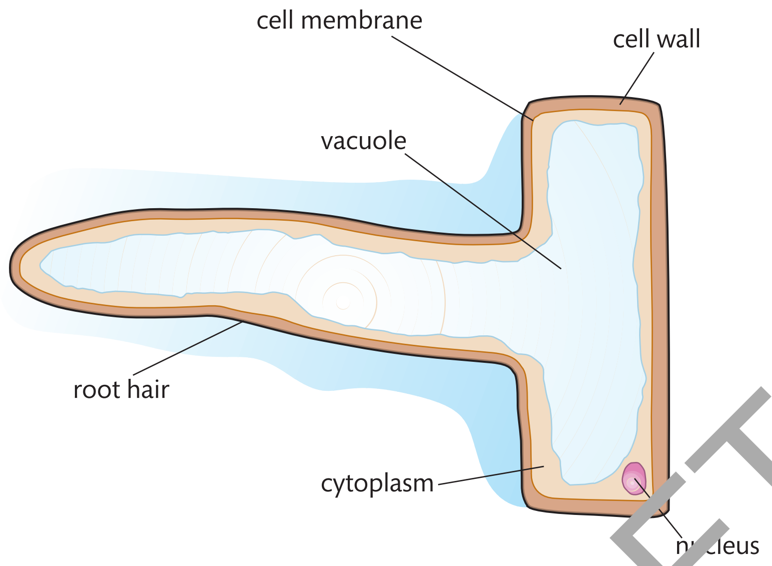


Figure 1.29: Basic structure of a root hair cell in plants

Sperm cells

Sperm cells are male **gametes** in animals (see Figure 1.30). They have a tail-like structure called a undulipodium so they can move. They also contain many mitochondria to supply the energy needed for this movement. In human sperm, the mid-piece of the tail is 7 µm long and the end is approximately 40 µm in length. The sperm head is 3 µm wide and 4 µm long. It is made up of an acrosome, which contains digestive enzymes. These enzymes are released when the sperm meets the egg, to digest the protective layer and allow the sperm to penetrate. The sperm's function is to deliver genetic information to the egg cell or ovum (female gamete). This is fertilisation.

Key terms

Water potential – the tendency of a cell to draw in water from outside by osmosis.

Gamete – one set of chromosomes compared to two sets in the parent cells.

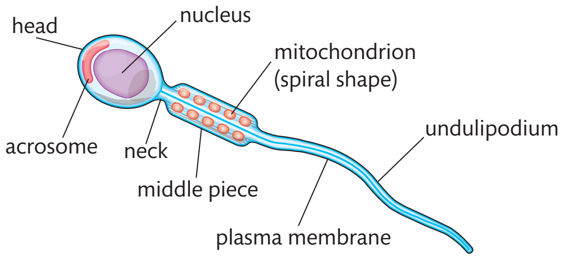


Figure 1.30: Sperm cell

Egg cells

Egg cells, or ova, are the female gametes in aminals (see Figure 1.31). An egg cell is one of the largest cells in the human body, and is approximately 0.12 mm in diameter. It contains a nucleus which houses the genetic material. The zona pellucida is the outer protective layer/membrane of the egg. Attached to this is the corona radiata, which consists of two or three layers. Its function is to supply proteins needed by the fertilised egg cell.

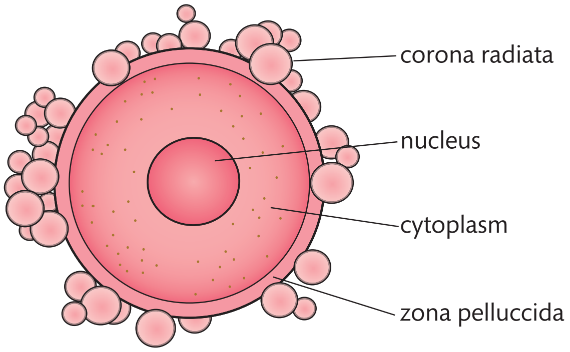


Figure 1.31: Basic structure of an ovum

Red blood cells

Red blood cells or erythrocytes are a bi-concave shape (where both sides concave inwards, see Figure 1.32). This increases the surface area to volume ratio of an erythrocyte. They are flexible so that they can squeeze through narrow blood capillaries. Their function is to transport oxygen around the body. In mammals, erythrocytes do not have a nucleus or other organelles. This increases space for the **haemoglobin** molecules inside the cell that carry oxygen.

Key term

Haemoglobin – protein molecule in red blood cells. It carries oxygen from the lungs to other parts of the body and carbon dioxide back to the lungs.

Figure 1.32: Erythrocyte

White blood cells

Neutrophils are a type of white blood cell and they play an important role in the immune system (see Figure 1.33). They have multi-lobed nuclei, which enables them to squeeze though small gaps when travelling to the site of infection. The cytoplasm holds lysosomes that contain enzymes that are used to digest **pathogens** that are ingested by the neurophil.

Key term

Pathogen – a microorganism that can cause disease.

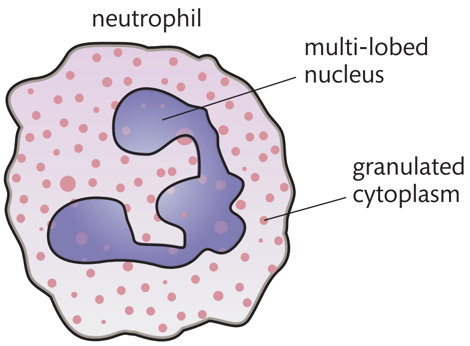


Figure 1.33: Basic structure of a neutrophil

PAUSE POINT Name four specialised cells and their functions.

Hint Think about their shape and how this enables them to function.

Extend Research two disorders in humans that may cause these specialised cells to change in structure and therefore not be able to carry out their function efficiently.

B3 Tissue structure and function

A collection of differentiated cells that perform a specific function is called a tissue. There are four main tissue types in animals:

- epithelium
- muscle
- connective (supports, connects or separates different types of tissues and organs in the body)
- nervous.

Epithelial tissue

Epithelial tissues are found lining organs and surfaces. Epithelial tissues can be divided into different types:

- squamous epithelial tissue
- columnar epithelial tissue.
- endothelium tissue.

Squamous epithelial tissue

Simple squamous epithelial tissue is a lining tissue and is one cell thick (see Figure 1.34(a)). It is made from flattened specialised squamous epithelial cells. These cells form a thin, smooth, flat layer. This makes them ideal when rapid diffusion is necessary. They line various structures. An example is the alveoli in the lungs where they provide a short diffusion pathway to allow rapid diffusion of oxygen into the blood and carbon dioxide into the lungs.

Epithelium cells can be damaged by smoking. Smoking irritates and causes inflammation and scarring in the epithelium tissue of the lungs. The alveoli walls become thicker due to scarring and produce more mucus. The damage to the air sacs causes emphysema and the lungs lose their natural elasticity. This causes:

- breathlessness
- persistent coughing
- phlegm.

These symptoms are all associated with Chronic Obstructive Pulmonary Disorder (COPD).

Ciliated columnar epithelial tissue

Ciliated columnar epithelium tissue is made up of column-shaped **citiated cells** with hair-like structures called cilia covering the exposed cell surface (see Figure 1.34(b)). Ciliated epithelium line the trachea in the respiratory system in order to protect the lungs from infection. They do this by sweeping any pathogens away from the lungs. Goblet cells are column shaped and are also present in the respiratory tract. They secrete mucus to help trap any unwanted particles that are present in the air that you breathe in. This protects your lungs because it prevents bacteria reaching the alveoli.

Key term

Ciliated cells – cells with tiny hair-like structures.

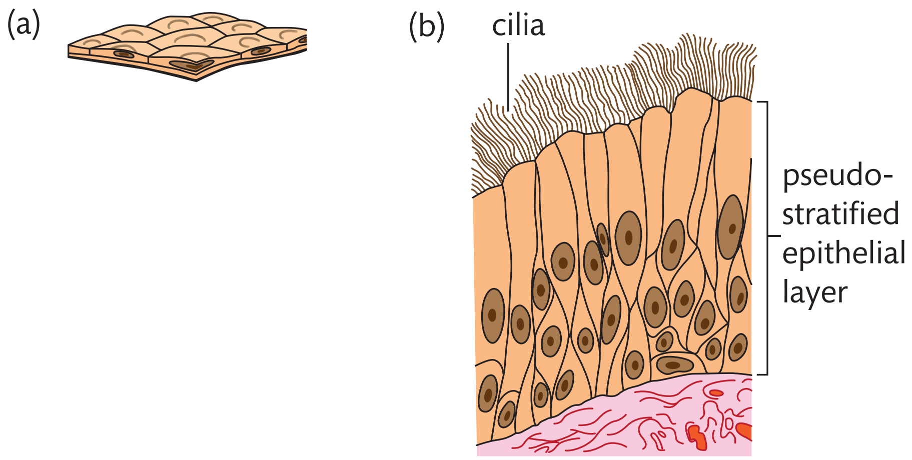


Figure 1.34: Epithelial tissue

Endothelial tissue

Endothelial tissue consists of a layer of flattened cells, one later thick. It is found lining the heart, blood vessels and lymphatic vessels (vessels that make up the lymphatic system). The cells provide a short diffusion pathway for the movement of various substances, such as:

- products of digestion into blood capillaries
- blood plasma and tissue fluid in and out of blood capillaries.

There are a number of risk factors that can cause damage to the endothelium. Carbon monoxide and high blood pressure can damage the inner lining of the arteries. White blood cells repair the damage and encourage the growth of smooth muscle and the deposition of fatty substances such as cholesterol under the endothelium lining of arteries, not on the surface. This process of deposition is called atherosclerosis. These deposits, called atheromas, may build up enough to break through the inner endothelial lining of the artery, eventually forming plaque in the **lumen** of the **artery**. This reduces the size of the lumen and restricts blood flow.

Key terms

Lumen – the space inside a structure.

Artery – blood vessel that carries blood away from the heart.

Assessment practice 1.14

Copy and complete the table below by comparing two types of epithelium: squamous and ciliated. For each type state one function and one specific location in the human body where it can be found.

Type of epithelium	Function of tissue	Specific location in the human body
Squamous		
Ciliated		

Muscle tissue

Muscles are composed of cells that are elongated and form fibres. Muscle cells contain protein filaments called actin and myosin that enable muscles to contract and cause movement.

There are three types of muscle tissue:

- **Skeletal** muscle is found attached to bones. You can control its contraction and relaxation, and it sometimes contracts in response to reflexes.

- **Cardiac** muscle is found only in the heart. It contracts at a steady rate to make the heartbeat. It is not under voluntary control.
- **Smooth** muscle is found in the walls of hollow organs, such as the stomach and bladder. It is also not under voluntary control.

Skeletal muscle fibre

Muscle tissue needs to be able contract (shorten in length) in order to move bones. In a muscle, cells join up to make muscle fibres. These are long strands of cells sharing nuclei and cytoplasm, which is known as the sarcoplasm. Inside the muscle cell cytoplasm are many mitochondria, specialised endoplasmic reticulum known as sarcoplasmic reticulum and a number of microfibrils. Each muscle fibre is surrounded by a cell surface membrane called the **sarcolemma**.

Skeletal muscle shows a stripy/banding appearance under a microscope. Skeletal muscle is made up of thousands of muscle fibres. Each muscle fibre is made up of myofibrils.

Key terms

Sarcolemma – cell membrane of a striated muscle cell.

Myofibril – basic rod-shaped unit of muscle cell.

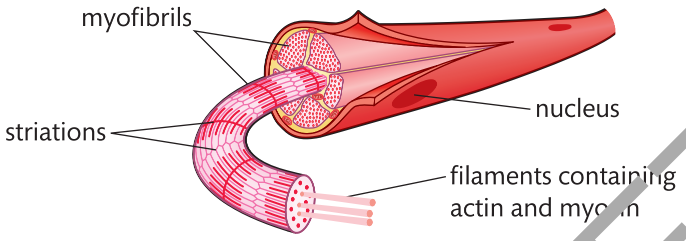


Figure 1.35: Structure of a muscle fibre

Myofibril fibres are made from proteins called myofilaments, which enables contraction to take place because of the contractile nature of the proteins in the filament. They appear as different coloured bands: A-band and I-band (you can remember which is which as A-bands are d**A**rk and I-bands are l**I**ght).

Sarcomere

The span from one z-line to the next in Figure 1.36 is known as the sarcomere. When the muscle is relaxed, this is approximately 2.5 μm in length. This length reduces when the muscle contracts because the I-band and H-zone lengths are reduced. The A-band does not change in length during contraction.

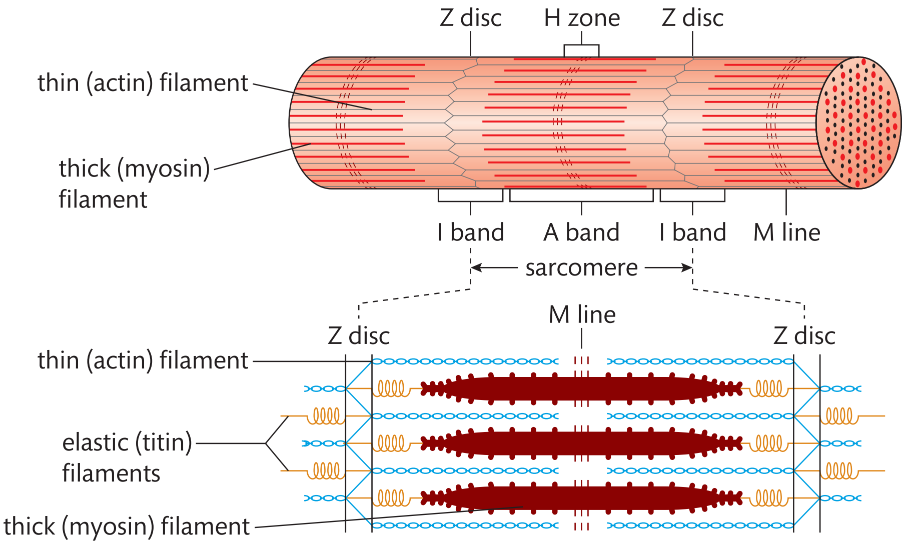


Figure 1.36: Muscle fibre microscopic structure

Image credit: after Essentials of Human Anatomy & Physiology, Benjamin Cummings (Marieb, E. 2003) Pearson Education, Inc.

There are two protein filaments found in muscle cells. This filament made of actin and thick filaments made of myosin. During muscle contraction, the thin actin filaments move and overlap the thick myosin filaments. The sarcomere shortens, decreasing the size of the overall muscle.

There are two types of muscle fibres: slow twitch and fast twitch. These fibres influence how muscles respond during physical activity. Human muscles contain a mixture of both.

Slow twitch muscle fibres

Slow twitch muscles are more effective at using oxygen to generate energy in the form of ATP, for continuous and extended muscle contractions over a long time. These fibres help marathon runners and endurance cyclists to continue for hours. Slow twitch fibres have:

- less sarcoplasmic reticulum
- more mitochondria for sustained contraction
- more myoglobin
- a dense capillary network.

These fibres release ATP slowly by **aerobic respiration**.

Fast twitch muscle fibres

Fast twitch muscle fibres can be divided into two different kinds.

- Fast twitch oxidative muscle fibres are similar in structure to slow twitch muscle fibres. They contain many mitochondria, myoglobin and blood capillaries, but they are able to **hydrolyse** ATP much more quickly and therefore contract quickly. They are relatively resistant to fatigue.
- **Fast twitch glycolytic muscle fibres** have relatively less myoglobin, few mitochondria and few capillaries. They contain a large concentration of **glycogen** that provides fuel for **anaerobic respiration**. They contract rapidly but also fatigue quickly.

Key terms

ATP – adenoisine triphosphate, an enzyme that transports chemical energy within cells for metabolism.

Hydrolyse – a chemical reaction involving breaking down a compound with water.

Anaerobic respiration – respiration without oxygen.

Glycogen – many glucose molecules connected and stored in the liver and muscles.

Dendrons – extension of a nerve cell.

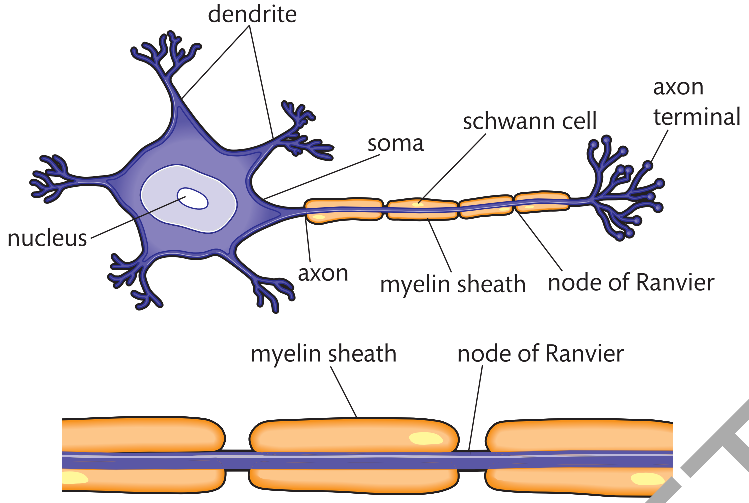
Aerobic respiration – respiration with oxygen.

Nervous tissue

The central nervous system (CNS) consists of the brain and spinal cord. It is made up of billions of non-myelinated nerve cells and longer, myelinated axons (axons with myeline sheath) and **dendrons** that carry nerve impulses. Nervous tissue is made of nerve cells called neurons.

Neurons

Neurons are cells that receive and facilitate nerve impulses, or action potentials, across their membrane and pass them onto the next neuron. They consist of a large cell body called a soma with small projections called dendrites and an axon. The end of the axon is called the axon terminal. It is separated from the dendrite of the following neuron by a small gap called a synapse.



▶ **Figure 1.37:** Myelinated neuron

Information travels along neurons in the form of electrical signals called nerve impulses. A nerve impulse is known as an action potential. Action potentials arise from a change in the ion balance in the nerve cell which spreads rapidly from one end of the neuron to the other. Neurons are bundled together to form nerves and nerves form a network all around the body. When the action potential travels to the axon terminal, neurotransmitters (chemicals) are released across the synapse and bind to the post-synaptic receptors, continuing the nerve impulse in the next neuron.

Sensory neurons receive information from receptors, for example, ears, and take this information to the CNS. The brain processes the information, then motor neurons take the information from the brain to the effector, for example, muscle.

Link

Go to *Unit 9: Human Regulation and Reproduction Learning aim A* to find more information about the nervous system.

Resting potential and action potential

Resting potential is the term given to a neuron that is not transmitting an action potential and is at rest. However, the neuron is actively transporting sodium and potassium ions across its membrane to maintain a negative potential in the interior of the cell compared to the outside. The membrane is more permeable to potassium ions, so for every three sodium ions actively transported across the membrane, only two potassium ions are actively transported. This process requires energy in the form of ATP. The cell membrane is described as polarised. At rest, the gated sodium ion channels in the membrane are closed and it is a sodium/potassium pump that is used to transport the sodium and potassium ions across the membrane creating a potential difference of −60 **mV**.

Key term

mV – millivolts, a small voltage/potential across a cell membrane.

When a nerve impulse is stimulated by a receptor cell or another neuron, an action potential is generated. The neuron is always ready to conduct an impulse. The axon membrane is polarised, which means that the fluid on the inside is negatively charged with respect to the outside. An action potential is the electrical potential which results from the process of ions moving across the neuron cell membrane when the correct channels are open in response to a stimulus, causing the inside of the neuron to be more positive than the outside.

Depolarisation

At rest, there are more positive ions outside the neuron. When an action potential is generated, there is a quick change in the permeability of the axon membrane that spreads down the whole neuron as a wave of depolarisation. The voltage across the membrane changes. A small number of voltage gated sodium channels detect this change, and open to allow a few sodium ions to diffuse into the axon. The membrane depolarises and becomes less negative than the outside, with a potential difference of -50 mV (threshold). If the stimulus is large enough and this threshold is reached, then the rest of the sodium gated channels open to allow rapid diffusion of sodium ions into the axon, making the inside positively charged in comparison to the outside, with a potential difference of +40 mV across the membrane. Sodium ion channels close and potassium channels open. Potassium ions therefore diffuse out of the cell. This makes the inside of the axon negatively charged again. This is called repolarisation and it restores resting potential.

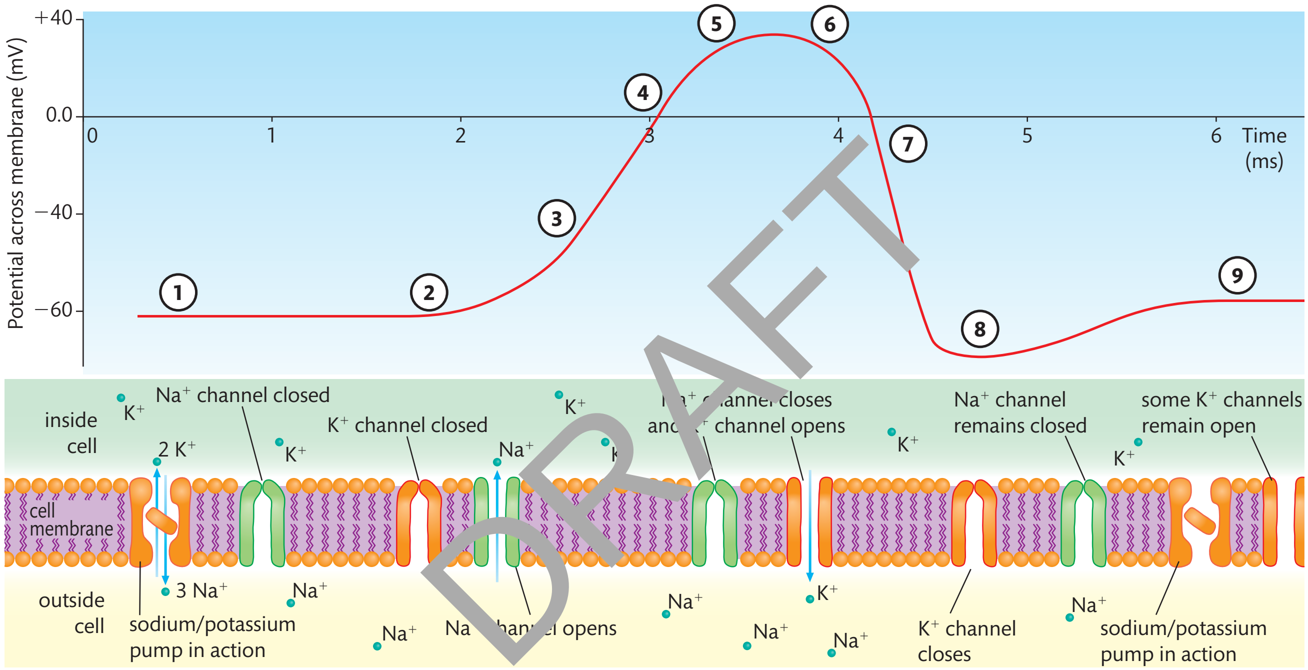


Figure 1.38: Depolarisation

Myelinated neurons and saltatory conduction

Some neurons have an axon covered with a fatty sheath called myelin (see Figure 1.37). Myelin is made from specialised cells called Schwann cells that wrap themselves around the axon when they develop in an embryo. The Schwann cells are thick and form a lipid insulating layer around the neuron called the myelin sheath. This insulates the axon and makes the action potential travel faster.

Key term

Nodes of Ranvier – the gap in the myelin sheath of a nerve cell, between Schwann cells.

Saltatory conduction happens only in myelinated nerves and it greatly increases the speed of the action potential. The myelin sheath insulating the axon means that ion exchange can only occur at the **nodes of Ranvier** (see Figure 1.37) that are in between the Schwann cells, where the axon membrane is exposed and not covered with Schwann cells.

Saltatory conduction is the process of the signal jumping (saltatory comes from the Latin *saltare*, meaning 'to dance'). When the action potential reaches a node of Ranvier, sodium ions diffuse into the axon membrane. They displace the potassium ions down the axon because they are both positively charged, and like charges repel. The movement of the potassium to the node further down the axon makes the next node more positive and depolarises it until the threshold is reached. The impulse quickly jumps from node to node, making the action potential quicker. Only a small part of the axon is being used, so less ATP is needed and fewer ions are being exchanged.

The speed of an action potential in humans

The speed at which a nerve impulse travels in humans is 1 to 3 m/s in unmyelinated fibres and 3 to 120 m/s in myelinated fibres. The speed of travel (conduction) depends on:

- axon diameter – the larger the axon, the faster the conduction
- myelination of neuron – the nerve impulse travels faster if the neuron is myelinated
- number of synapses involved – the fewer synapses there are to cross, the faster the communication.

Synapses

When the nerve impulse reaches the end of the neuron, it must cross a gap called a synapse (see Figure 1.39) to get to the next neuron or the effector cell. A nerve impulse crosses the synapse in the form of a chemical transmitter called a neurotransmitter. Neurotransmitters diffuse across the synapse and initiate an action potential in the neuron at the other side. The presynaptic neuron ends in a swelling called the synaptic bulb and it contains many mitochondria as ATP is needed. The neurotransmitters are stored in temporary vesicles in the synaptic bulb that can fuse with the surface to release the neurotransmitters into the synapse. They also contain voltage-gated calcium ion channels.

There are hundreds of neurotransmitters. The most common neurotransmitter is acetylcholine and synapses that have this as their transmitter are called cholinergic synapses. Acetylcholine molecules are released by exocytosis and they diffuse across the deft. The acetylcholine molecules bind to the receptor sites on the sodium ion channels in the postsynaptic neuron to generate a new action potential.

Step by step: chemical transmission across the synapse

8 Steps

1 The action potential arrives at the synaptic bulb

2 Calcium channels open in the presynaptic membrane. Calcium ions diffuse into the neuron membrane down a concentration gradient.

3 As the calcium concentration increases, the synaptic vesicles containing neurotransmitters move towards the presynaptic membrane.

4 The vesicles fuse with the membrane and release the chemicals into the synaptic cleft.

5 Neurotransmitters diffuse across the synaptic deft. This is known as synaptic delay because it is slower than an electrical signal travels.

6 The neurotransmitter binds to the postsynaptic cell membrane receptor sites on the sodium channels.

7 Some neurotransmitters open sodium channels in the membrane, causing sodium ions to pass in. This creates a excitatory postsynaptic potential (EPSP) and makes the membrane receptive to the signals coming in. If this reaches the threshold, the action potential is generated.

8 The neurotransmitter will excite the cell and, once it has acted on the membrane, enzymes act on the neurotransmitter to break them down.

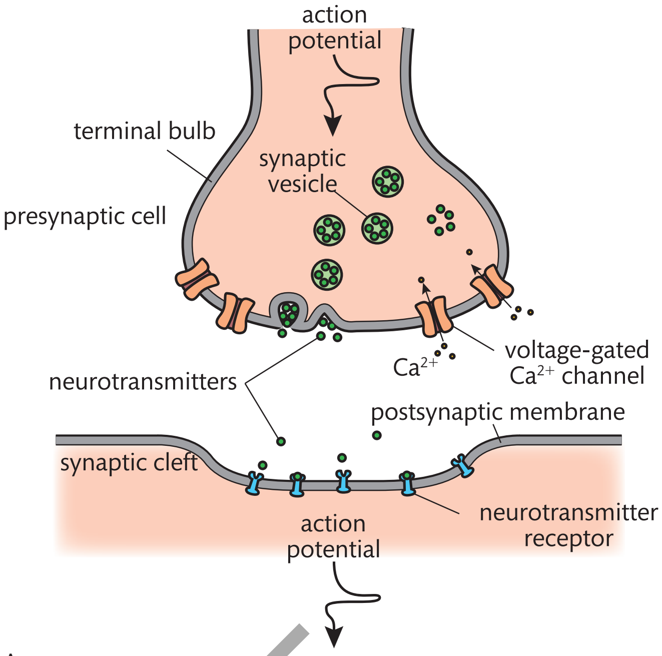


Figure 1.39: A synapse

An electroencephalograph (EEG) is a test that looks at the activity of the brain cells. When the brain is working, nerve impulses travel from one cell to another. These produce electrical signals that can be picked up by detectors attached to a person's skull. The detectors send the signals that they detect to a recorder. The recorder produces a graphical trace which can be interpreted to see whether there is any abnormal activity, such as that which may suggest that an epileptic seizure has occurred.

Problems that can occur

Key term

Genetic – related to heredity and variation.

Parkinson's disease is a **genetic** disease that affects the nervous system. Parkinson's sufferers are not able to produce the naturally occurring chemical dopamine, a neurotransmitter that helps smooth and normal movements. Without this, people show symptoms of:

- slow movement
- tremors when moving
- speech problems
- poor balance.

The drug, L-dopa, replaces the dopamine that is lost in people with Parkinson's disease.

Serotonin is another of the body's naturally occurring neurotransmitters. It is normally active in the brain and can cause problems if it is not produced. Some forms of depression are caused by a reduced amount of serotonin in the brain.

PAUSE POINT Synapses are an integral part of the nervous system. Outline the role of the synapse in the nervous system.

Hint Draw a synapse and add commentary to show the function of the synapse on the nervous system.

Extend Research beta blockers and Prozac™ and discuss their effect on the synapse.

Websites

Boyle, M. & Seenior, K. (2008) Biology (Third edition). Collins

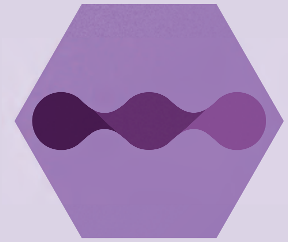
Kennedy, P. and Sochacki, F. (2008). *OCR Biology AS*. Oxford: Heinemann Educational (ISBN 9780435691806).

Getting started

Modern communications involve technology: phone, email, radio and TV, social media. In scientific work, modern technology is used to make and record observations as well as to analyse and share them: spectroscopy, endoscopy, data-logging, satellite imaging and so on.

All these communications and measurement technologies depend on ways in which waves behave. To get the best out of them, you need to understand waves.

What do you already know about waves? How many kinds of wave can you think of and picture? Write a list of the main terms you know for describing waves. When you have completed this unit, see if you can add to that list.



C Waves in communication

C1 Working with waves

Waves generally start with a disturbance. For example, wind blowing across the surface of the sea, or a stone being thrown into a pond. The energy imparted by that disturbance causes a regular repeating motion, backwards and forwards or up and down, which is called an **oscillation**.

Oscillations, period and amplitude

Examples of oscillations are the pendulum of a clock, a child on a swing, a weight bouncing on a spring. Often, as in the suspension of a car, our aim is to try to damp down the oscillations as soon as possible. These oscillations are not themselves waves because they don't travel anywhere, but like waves they do have a **frequency** and **periodic time**, which describe the rate at which the oscillation repeats itself.

In an oscillation, something is displaced from its rest position, but it also has a tendency to bounce back. In a physical oscillation, like the examples above, the **displacement** is a distance moved by something from rest. But, for example, in an electrical oscillation, the displacement would be a change of voltage or of current going regularly up and down in value. In either case, we measure the size of an oscillation by its **amplitude**

How can an oscillating system sometimes produce a wave motion?

Wave motion

Waves transfer energy from one place to another, but without causing any net movement of material.

The energy transfer depends on the way an initial oscillating system is connected to its surroundings. If that connection can carry some energy from that first oscillation and transfer it to a similar system next to it, then that system will also start oscillating. However, that second oscillation will not be quite in time with the first one.

Key terms

Oscillation – a regularly repeating motion about a central value.

Period (or 'periodic time') – the time taken for one whole cycle of an oscillation, i.e. before the motion starts to repeat itself. (Symbol: *T*; SI unit: s.)

Frequency – $f = \frac{1}{T}$ – i.e. the number of whole cycles occurring in one second. (Symbol: *f*; SI unit: Hertz, Hz.)

Displacement – how far the quantity that is in oscillation has moved from its mean (rest) value at any given time. (Symbol and unit: various according what is the quantity that is oscillating.)

Amplitude – the maximum value of displacement in the oscillation cycle – always measured from the mean (rest) position.

Key term

Wavelength – the distance along the wave in its direction of travel (propagation) between consecutive points where the oscillations are in phase.

Have you experienced a heavy person sitting down next to you on a springy sofa or bed? As they sit down at one point, you, some distance away, may find yourself going up. If it is a very springy seat, then you might both find yourselves bouncing a couple of times before the oscillation dies away. If you watch children jumping on a bouncy castle, or two gymnasts sharing a trampoline, you might see something similar happening. This is the start of a wave.

When a wave transfers the energy of an oscillation, it takes time. So a short distance away, though a similar oscillation happens, it is delayed in time. If you travel with the wave for one whole **wavelength**, you will find another place where the oscillation does once again occur exactly in time with the first one. In fact, it is now delayed by one whole cycle – that is, the time delay is equal to the oscillation period. Two such points along the wave are said to be 'in phase' with one another. (We'll explore the idea of phase a bit more in the section below about graphs of waves.)

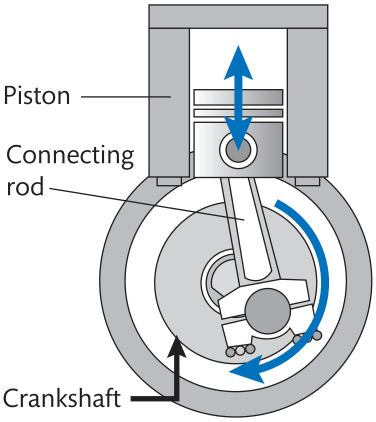
Wave speed

A wave travels one wavelength during its periodic time. So that means you can calculate its speed, v, as wavelength, λ, divided by periodic time, T. However, instead of the periodic time, frequency is more commonly used, f, where $f = \frac{1}{T}$. Frequency is measured in cycles per second or Hertz (Hz). It is often easier to use frequency, because periodic times are usually tiny fractions of a second. The faster the oscillations, the larger is the value of the frequency. Using frequency, you can rewrite the equation for the speed of a wave as:

$$v = f\lambda$$

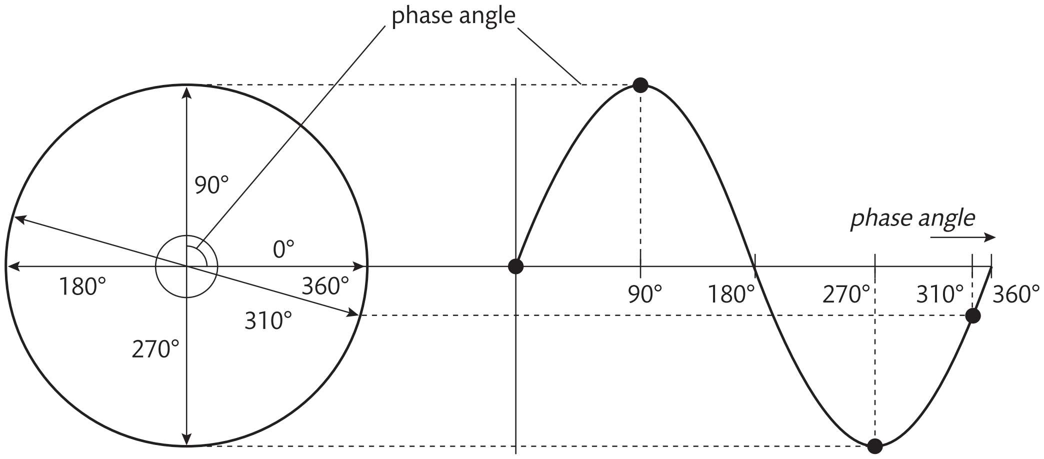
Graphical representation of wave features

One example of an oscillating system is a piston in a motor car engine. The piston is connected to a rotating crankshaft, and that is what (via the transmission system) drives the vehicle's wheels in circular motion. One complete oscillation of the piston corresponds to one whole turn of the crankshaft. They both have the same periodic time (and frequency).



▶ **Figure 1.40** Pistom in a motor car engine.

So the mathematics of oscillation and of circular motion are closely connected. Figure 1.41 shows you how. This is why the graphs you draw of oscillations and waves are typically sine waves. The sine is a mathematical function of the angle through which you can imagine a crankshaft turning to drive the motion.



▶ **Figure 1.41:** How a rotating vector generates a sine wave

You can use this idea of the angle generating a cycle of oscillation when you compare two wave motions that are not in phase with one another. The **phase difference** is usually given as an angle, where 360° (or 2π radians) equates to a whole cycle – a shift equivalent to one wavelength in distance or one period in time.

Table 1.11: Phase angles in degrees and radians, and also compared to wavelengths.

Wave cycle position	**Start**	**$\frac{1}{4}$ of a cycle**	**$\frac{1}{2}$ of a cycle**	**$\frac{3}{4}$ of a cycle**	**1 whole cycle**	**1.5 cycles**	**2 whole cycles**
Phase/°	0	90	180	270	360	540	720
Phase/rad	0	$\frac{\pi}{2}$	π	$\frac{3\pi}{2}$	2π	3π	4π
Number of wavelengths	0	$\frac{\lambda}{4}$	$\frac{\lambda}{2}$	$\frac{3\lambda}{4}$	λ	$\frac{3\lambda}{2}$	2λ

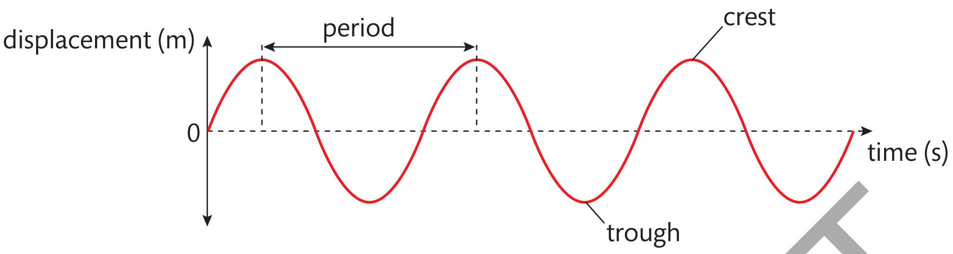


Figure 1.42: Graph showing how the displacement varies over time at one fixed position in space as a wave travels past

Key term

Phase difference – the difference in phase angle between two waves of the same frequency and wavelength, where 360° (2π radians) represents a single whole cycle of the waveform.

Look at the two graphs in Figures 1.42 and 1.43. They represent the same wave motion, but there is too much going on in it for everything to be captured in one still picture. So one graph concentrates on the changes happening at a fixed point in space, while the other is for a fixed point in time and shows how the wave extends through space.

The quantity represented on the vertical axis of both these wave graphs is the displacement of whatever is oscillating from its rest value. So for different kinds of wave, this will be a different physical property.

- For water ripples it is the water level.
- For sound waves it is microscopic movements of molecules linked to pressure variations. For light, radio and other electromagnetic waves it is an oscillating electric field.

Waves create a pattern of displacements in space as well as in time. So, taking a snapshot in time, you can picture a wave using a graph of displacement against position along the direction in which the wave travels (propagates). See Figure 1.43.

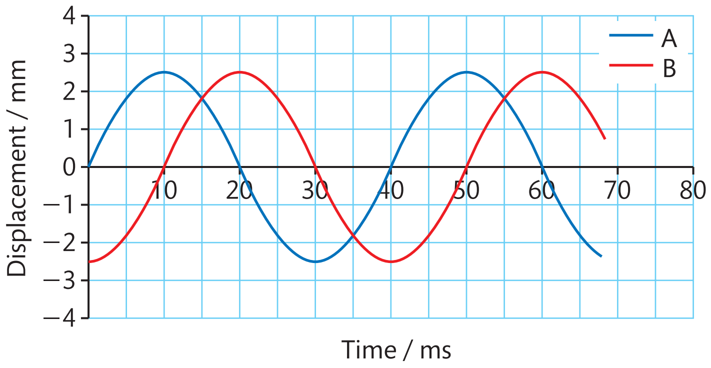


Figure 1.43: The same wave, but showing how displacement varies with distance along the wave in the direction of its travel: a snapshot at one fixed point in time

On this graph, one whole cycle along the horizontal distance axis marks out a wavelength. Note that moving forward along the wave in distance is equivalent to moving backwards in time so far as the phase of the oscillations at that point is concerned.

PAUSE POINT

Close the book and try using graphs to explain some of the key wave terms:- displacement, period, phase, frequency, wavelength, amplitude. First draw a rotating vector and wave diagram using a pair of compasses and some graph paper.?

Hint

Extend

On the time axis, mark off the wave period, allowing space on the axis for two whole cycles. Mark phase angles along the same axis. Use measurements of **displacement** taken from the generating circle to help you plot out the points.

What happens to the shape of the wave if you double the **frequency**?

What other type of graph can you draw in order to show the wavelength?

Types of wave motion: transverse and longitudinal

When the displacement occurs in the same direction that the wave travels, for example in a sound wave, it is a longitudinal wave. By contrast, in a transverse wave the displacement is at right angles to the direction of propagation of the wave, for example, water ripples and electromagnetic waves.

Transverse waves are easy to picture because they look like the sine wave graphs you normally draw, with displacement on the vertical axis and the distance travelled by the wave plotted horizontally. (See Figures 1.42 and 1.43.)

However, in a longitudinal wave, the different displacements of particles along the direction in which the wave is propagating, lead to a series of compressions (where particles are packed closer together) and rarefactions (where they are further apart). You can create a longitudinal wave in a spring by making it oscillate along its length (see Figure 1.44). A soft 'slinky' spring is best for seeing clearly the compressions and rarefactions that travel down its length.

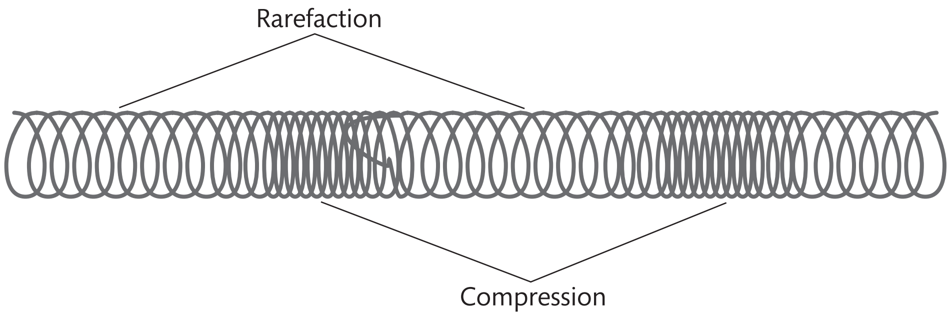


Figure 1.44: Compressions and rarefactions travelling down a spring

Because of this, sound and other longitudinal waves are sometimes described as pressure waves: oscillations in pressure travelling through a solid or fluid medium.

Earthquakes and other seismic events below the earth's surface generate two types of shock wave: a longitudinal 'pressure' wave and a transverse 'shaking' wave. They travel at different speeds and so will each arrive at different times, making earthquakes quite complex events to study

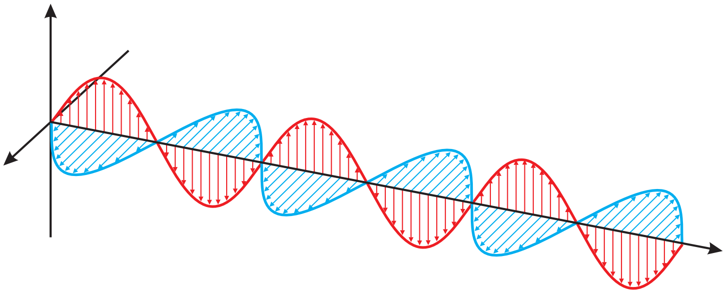


Figure 1.45: Electromagnetic waves are transverse in three dimensions: so some can be 'polarised' in one plane and some in another plane at 90° to the first one

Theory into practice

Because electromagnetic waves are transverse, they can be polarised. The direction of displacement plus the direction of propagation define a plane of polarisation. Passing electromagnetic radiation through a polarising filter will remove all the components that have oscillations in the plane at 90° to that of the filter. Polarising sunglasses work because when sunlight is passed through a polarising filter, half of sunlight's intensity is removed. As light that has been reflected tends to be mostly polarised in one plane, polarising sunglasses are great for cutting out glare from reflective surfaces.

Explain why there cannot be an equivalent to this for sound waves. What is different about them?

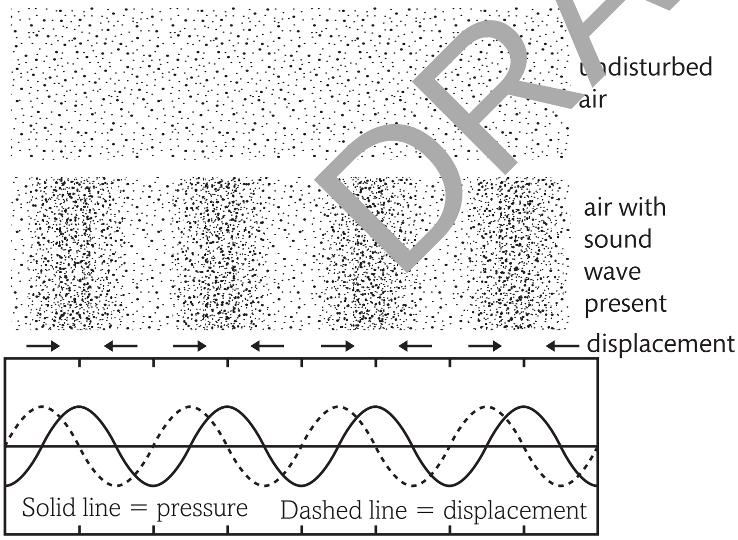


Figure 1.46: Transmission of sound as a longitudinal wave in air

Diffraction gratings

Diffraction

Diffraction is a key characteristic of all waves. It means the tendency of a wave to spread out in all directions, transferring energy to its surroundings as it does so.

When a wave is moving straight forward, for example, in a beam of light, you can picture moving wave-fronts. Each is at right-angles to the direction of propagation (i.e. the direction of travel and energy transfer).

These wave-fronts will also be straight lines. In three dimensions they form a flat plane. (See Figure 1.47.)

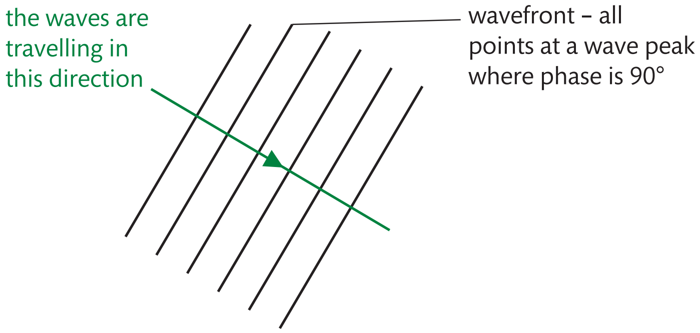


Figure 1.47: Wave-fronts are perpendicular to the direction of a ray, and show the points where the wave oscillations have the same phase

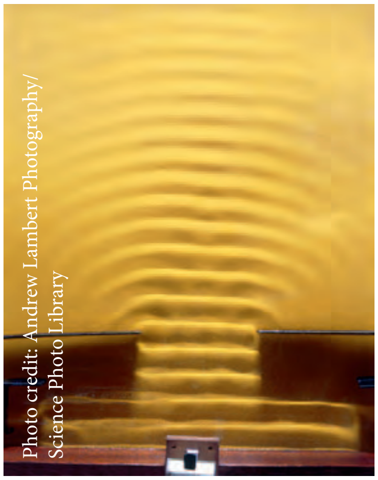


Photo credit: Andrew Lambert Photography/ Science Photo Library

Ripple tank photo showing diffraction through a gap

If the advancing wave-fronts encounter a flat obstacle in front of them, like a wall, most of the wave's energy is either absorbed or reflected by the wall. However, if the obstacle has edges or gaps, wave energy can travel round the edges or through the gaps. It is then that you may notice diffraction occurring. Although after going through a gap much of the wave energy does keep moving forwards, some of it spreads out in other directions.

A good way to see this effect is to use water ripples generated in a specially-designed glass-bottomed tray (a 'ripple tank'). By shining a light downwards through the water in the tray onto a horizontal white screen or sheet of paper, the moving ripples can be clearly seen as bright lines. If obstacles are introduced with edges, or with gaps about whose size is a few times the wavelength of the ripples, you can observe diffracted ripples with curved wave-fronts, even though the original ripples had straight line wave-fronts.

Every point along a wave-front has oscillations and energy. So each point on the wavefront can act as a secondary source of circular ripples spreading out in all directions. In this way, all the little secondary ripples add together to make a straight (or plane) wave-front keep moving forward as a straight line, until it meets an obstacle.

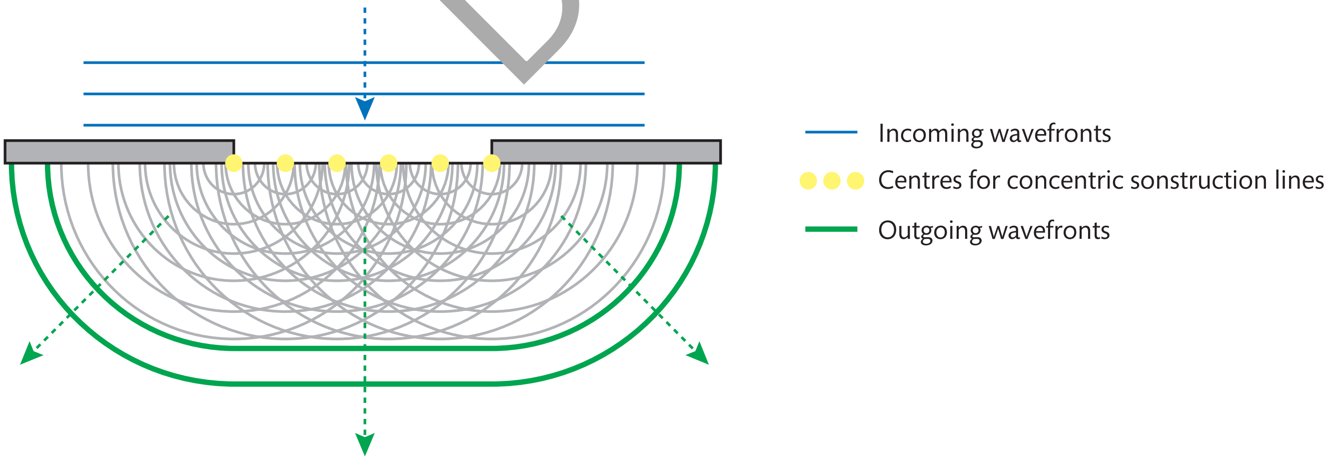


Figure 1.48: Huygens' construction correctly predicts diffraction through a gap

The dutch mathematician and scientist, Christiaan Huygens, developed a geometrical construction to predict the shape of waves in water (see Figure 1.48). In 1678, he was the first to apply this wave-front propagation principle to light and he showed that wave theory could explain all the behaviour of lenses and mirrors. However, it was Thomas Young's experiments on diffraction and interference of light in 1801 that finally convinced the whole scientific community to use a wave motion theory for light.

Gratings

A **diffraction grating** is a flat plane object. It has a series of regular lines formed on it that block parts of an advancing wave-front. For microwaves, these lines could be a series of regularly spaced metal bars or wires. For light you would usually use a piece of glass with a series of very fine and regularly spaced scratches on its surface. You can get a similar effect for X-rays by using the regularly spaced rows of atoms or molecules in a crystal.

When a wave-front meets a diffraction grating, some of the wave's energy continues propagating forward through the gaps between the grating lines. This is **transmission**. Some more of the wave's energy may be absorbed in the grating itself, but the remainder of the energy is scattered backwards as a **reflection**.

Think just about the forwards (transmission) direction.

The grating creates a set of regularly spaced secondary sources, where each gap in the grating allows energy through. Each gap generates a set of circular wave-fronts spreading out from that location. The spacing, *d*, between the grating lines is very important. If it is close in size to the wavelength, λ, of the incident radiation (the incoming waves) then the grating will produce an **interference pattern** of regularly spaced bright and dark lines (fringes). (See below). These correspond to strong intensity created at certain angles after diffraction and no wave energy at all at certain other intermediate angles.

Key terms

Transmission – wave energy passing through an object, e.g. a diffraction grating, and mostly continuing forward in the original direction, though some energy will be diffracted through angles of less than 90°.

Reflection – wave energy that bounces off a surface and has its direction of travel altered by more than 180°.

The interference pattern is due to **superposition** of the waves from the separate, regularly spaced sources, coming through the gaps in the diffraction grating. Wherever the **path difference** between waves from adjacent sources works out to be a whole number, *n*, of wavelengths, then the displacements due to waves from all the separate sources will be in phase. They add together constructively to give a bright spot of high intensity of radiation at that point. The same effect can be clearly demonstrated with water ripples (see below).

Photo credit: Edward Kinsman/Science Photo Library

▸ Superposition of waves from two separate sources demonstrated in a ripple tank

Key terms

Superposition – is the adding together of wave displacements that occurs when waves from two or more separate sources overlap at any given location in space. The displacements simply add mathematically.

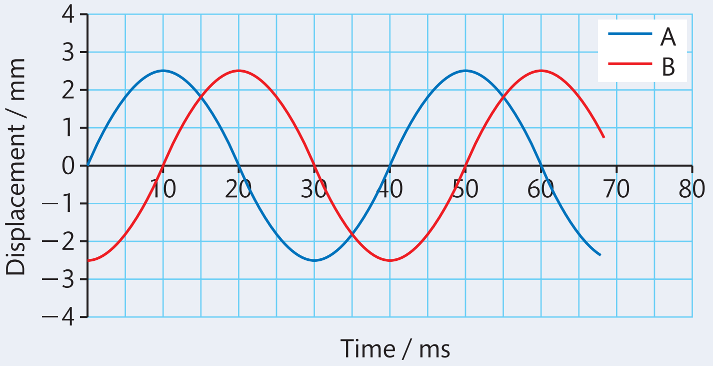
Path difference – is the difference in length between two (straight line) rays, e.g. one from a particular grating gap to a given point in space and the ray from the next-door grating gap to the same point.

Interference pattern – a stationary pattern that can result from the superposition of waves travelling in different directions, provided they are **coherent**.

Coherent – literally means 'sticking together' and is used to describe waves whose superposition gives a visible **interference pattern**. To be coherent, waves must share the same frequency and same wavelength and have a constant phase difference.

Assessment practice 1.15

The graph shows two waves of the same frequency passing through a single point in space.



a) What is the phase difference between wave A and wave B?

b) Sketch a graph of the resultant wave motion at that point in space and determine:

   i) its amplitude

and

   ii) its phase angle relative to wave A.

Where the path difference between waves from adjacent gaps in the grating works out to be half a wavelength – or an odd number of half wavelengths – the interference between the waves will be destructive. That means that the wave displacement due to the wave energy coming from one grating gap is completely out of phase with that coming from the next door gap. They have a phase difference of 180° (π radians) so that whenever one has a positive displacement, its neighbour has an equal but negative displacement. This means they cancel one another out and the resulting wave intensity is zero – a dark spot.

Bright positions of high intensity radiation do not only occur in the straight ahead 'transmission beam' direction. They also occur at any angle, ϑ, to the transmission beam direction that satisfies the equation:

$$n\lambda = d \sin\vartheta$$

where n = order of diffraction, λ is wavelength, d is grating spacing and ϑ is diffraction angle.

The bright interference line at the angle where $n = 1$, i.e. the where path difference is one wavelength, is called first order diffraction. Where $n = 2$, it is second order diffraction … and so on.

If the radiation consists of a mixture of waves of different wavelengths, for example, white light, then the transmission beam ($n = 0$, so $\vartheta = 0$) has all the wavelengths in one place and so is still 'white'. However, for the first and higher orders of diffraction the angle, ϑ, will vary with the wavelength, λ, and as a result you will see a separated (or 'dispersed') spectrum of different coloured lines (see below). The separation of the lines is greater for the second order diffraction compared with the first, but the intensity of the lines lessens as the angle ϑ increases.

It is easy to imagine and to demonstrate diffraction by a grating in transmission mode, as described above, for water ripples and for radio waves or microwaves. Diffraction of light can also be done with a grating in transmission mode. However, for light, and even more for X-rays and γ-rays, it is more common to use reflection mode.

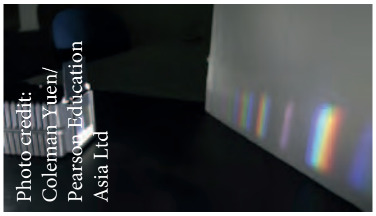


Photo credit: Coleman Yuen/ Pearson Education Asia Ltd

▸ Coloured spectra produced by diffraction of white light by a grating

Gratings in reflection mode

In reflection mode, instead of looking at what comes through a grating, you look at the part of the wave energy that is bounced back off the grating surface. Once again, because the grating lines are regularly spaced, an interference pattern is produced. The geometry and equations are just the same as for transmission, but on the opposite side of the grating. That is because you are still looking at a regularly spaced set of secondary sources of waves, but now it is the reflected wave-front from the grating surface between each line that counts.

The advantage of reflection mode is that you do not have to worry about the transparency of the grating. So long as the wave energy hitting the grating 'lines' gets lost, it does not matter whether it has been absorbed or transmitted. It is the reflectivity of the grating's surface that counts for giving a strong, measurable signal.

Coherent light sources

Wave theory is used to describe light (and all the other kinds of electromagnetic radiation) because it gives a good explanation of basic light properties you observe, notably diffraction and refraction, but also reflection.

However, when light is emitted from or absorbed by matter, you can only explain what happens by thinking of light as being composed of tiny particles or 'packets of energy', which are called **photons**.

So, when thinking about the **coherence** of light, you have to combine ideas from wave theory with the idea of individual photon particles – what is called 'wave-particle duality'. To try to visualise this you can use the concept of 'wave packets' or 'wave trains' (see Figure 1.49). These are snippets of wave motion of a given frequency that start up from zero with a growing amplitude and then die away again.

Key terms

Photon – a **quantum** of electromagnetic radiation. Photons have zero mass and zero charge, but a definite energy value linked to their frequency.

Quantum – the smallest unit that can exist independently which has clearly defined values of energy, mass, charge and other physical quantities.

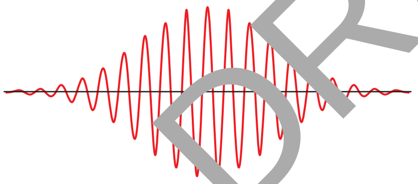


Figure 1.49: The displacement-time and displacement-distance graphs for a photon wave packet might look like this – a very short burst of waves or 'wave train'

More evidence about the nature of light and about coherence comes from interference, for example, the patterns produced by diffraction gratings. You can only demonstrate interference between beams of light that were produced from the same source but which have subsequently been split up, for example, by:

- mirrors
- passing through a double slit (Young's experiment)
- the lines of a diffraction grating.

However, if the path difference between the beams gets larger than a certain size – called the coherence length – then interference fringes fade and disappear. That can happen if the spacing of a double slit is gradually increased. Similarly, if the phase difference is increased beyond a value equivalent to a time known as the coherence time, by adjusting the position of a beam-splitting mirror, again the interference pattern disappears.

So, for light and other electromagnetic waves, the conditions for coherence go beyond the simple frequency and phase requirements that apply for other waves like water ripples or sound.

No detector is fast enough to directly record the frequencies and phase relationship of two light waves. So you cannot directly measure whether or not they are coherent. Instead, you have to base your understanding of light wave/particle packets (i.e. photons) on the results of interference experiments.

Photons produced from separate sources are emitted randomly at different instants, and it seems they do not generally overlap in time enough to give interference. So they must be described by bursts of waves of very short duration. Measuring the energies of individual photons, emitted at different moments from the same source, gives a small spread of energies (bandwidth) and hence also of frequencies – another obstacle to coherence.

It is because they have a long coherence length that all the photons produced in the same pulse from a laser are able to pack so closely and to travel together in such a tightly directed beam.

But if you direct light beams from two different LED sources, or even from two separate lasers, at the same screen, you never see interference patterns occurring. They are not **coherent**.

PAUSE POINT

What *is* light really? Is it a wave motion, a stream of particles - or can it be both at the same time? Try out the idea of wave trains to see if you can explain the conditions under which interference will or will not occur.

Hint

Sketch two wave trains. Show them overlapping by different amounts along the horizontal axis. Think about how superposition will work and what the resultant wave form will look like. Then explain what coherence length means.

Extend

a) Why do displacement-time and displacement-distance graphs have a similar shape?

b) A source like an LED emits photons with very slightly differing frequencies. What effect will that have on the coherence time??

Emission spectra

The **quantum theory** of light and other electromagnetic radiations is based on the experimental observation that there is a simple relationship between the frequency, *f*, of the radiation and the energy, *E*, carried by each photon:

$$E = hf$$

where h is the Planck constant, $-6.626\ 070 \times 10^{-34}$ Js. That constant of proportionality between energy and frequency has been very precisely measured and experiments indicate it is universal.

Key terms

Quantum theory – combines ideas from wave motion and particle mechanics theories to create a new 'wave mechanics'. At the sub-atomic level all the particles – protons, neutrons, electrons, photons, etc. – also behave like waves (e.g. they can be diffracted). When they are bound into an atom or molecule, these particles behave like stationary waves with a fixed wavelength and energy.

If a chemical element or compound is vaporised by heating in a flame, or if you pass an electric current at high voltage through a gas, you typically see light emitted of a characteristic colour, according to the chemical nature of the material you are testing.

When you look at the spectrum of that light, by splitting it up using a prism or a diffraction grating, what you see is a number of bright, coloured lines at definite frequencies. This is an emission spectrum. Each line in the spectrum matches to photons all emitted with very nearly the same frequency – and therefore they also each have virtually the same energy.

You can explain these lines by thinking about the possible **energy levels** for electrons in the atom or molecule concerned. When a gas or vapour is cool, most of the atoms/molecules will be in the **ground state** where their orbital electrons all in the lowest energy state possible. Just as water naturally flows downhill, things in general tend to gravitate towards the lowest energy state possible.

However, in a hot or electrically excited gas or vapour, many of the outer electrons get knocked into higher energy levels. Then, as they begin to drop back down to lower levels and eventually back to the ground state, they have to lose energy.

De-excitation occurs one electron at a time, at randomly unpredictable instants in time. Electrons in highly excited states (i.e. those in higher energy levels) may make the journey back down to the stable ground state in two or three jumps. Rather like the balls in an arcade game of bagatelle, they may spend some time resting in one of the intermediate energy levels. Each transition from a higher to a lower energy level means that a specific amount of energy (the difference between the two energy levels) has to be lost by the electron.

That extra energy is emitted as a photon of light. Because the energy levels are fixed, the energy differences between them are always the same and are typical of the particular atom or molecule in which the electron is bound. The gas samples that are investigated, for example, by passing an electric current through them in a discharge tube, contain vast numbers of atoms. So there are always many atomic electrons making these de-excitation transitions, and you see a spectrum of lines (light frequencies) corresponding to all the transition paths they can take. Even for hydrogen, the simplest atom of all with just one electron, there are several energy levels and even more possible electronic transitions between them – so there are lots of lines in the hydrogen spectrum.

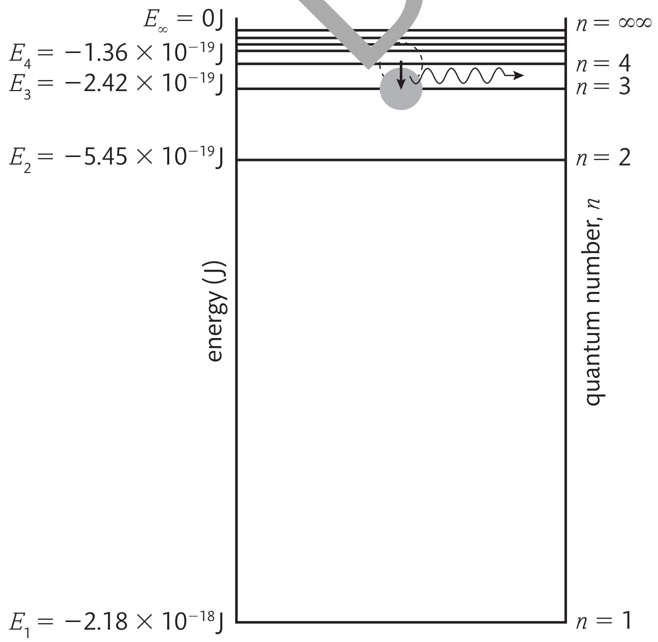


Figure 1.50: Electron de-excitation in a hydrogen atom – a photon of a specific frequency is emitted.

Key terms

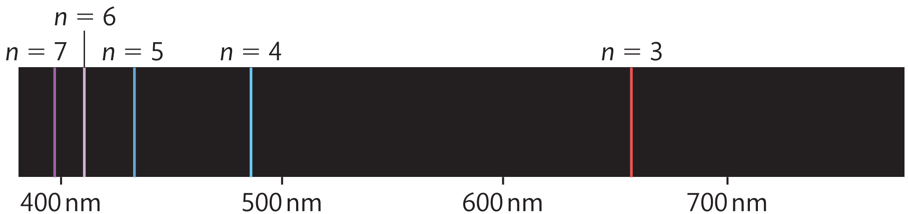
Energy level – one of the fixed, allowed values of energy for an electron that is bound in an atom, or for a proton or neutron that is bound in a nucleus.

Ground state – the lowest energy state possible for a given bound particle.

Theory into practice

You can use the pattern of lines in an emission spectrum as a 'fingerprint' to identify which atoms and molecules are present in a particular chemical sample. First you record the spectra from pure samples of each gas. Then you use those results as a reference to identify the frequencies of spectral lines obtained from a sample of unknown composition.

Discuss in pairs: As well as identifying the presence of a particular gaseous material, what could the relative intensities of different lines you measure tell you about the composition of the gas sample?



▸ Hydrogen visible emission line spectrum (Balmer series)

Worked Example

What is the frequency and wavelength of the photon emitted in Figure 1.50?

The electron drops from level $n = 4$ down to level $n = 3$.

The energy difference between the levels (the energy lost by the electron) must equal the energy of the emitted photon.

From the diagram, the energy level difference, ΔE, is from -1.36×10^{-19} J down to -2.42×10^{-19} J.

$$\Delta E = E_3 - E_4 = -2.42 \times 10^{-19} - -1.36 \times 10^{-19} = -1.06 \times 10^{-19}\ \text{J}$$

The fact that this energy value is negative indicates that energy is given out here.

But ΔE = the photon energy = hf

So $f = \dfrac{\Delta E}{h}$

$$f = \frac{1.06 \times 10^{-19}\ \text{J}}{6.63 \times 10^{-34}\ \text{Js}} = 1.60 \times 10^{14}\ \text{Hz}$$

$c = f\lambda$ where c is speed of light, $3 \times 10^{8}\ \text{m s}^{-1}$

$$\lambda = \frac{c}{f}$$

$\lambda = \dfrac{3.00 \times 10^{8}\ \text{m s}^{-1}}{1.60 \times 10^{14}\ \text{Hz}} = 1.875 \times 10^{-6}$ m = 1875 nm so this spectral line will appear in the infra-red.

PAUSE POINT

Test your understanding of spectra by investigating further the hydrogen spectrum. In what parts of the electromagnetic spectrum would you find lines corresponding to the highest energy electron transitions?

Hint

Reading off data from Figure 1.49, use the method shown in the worked example to also calculate frequencies for the transitions $n = 3$ to $n = 2$ and from $n = 2$ to $n = 1$.

Extend

The hydrogen spectrum contains several series of lines (Lyman, Balmer, Paschen, Brackett, ...) in different parts of the spectrum. Is it the upper or the lower energy level of an electron transition that determines to which series it belongs? Why?

Stationary waves resonance

In a **stationary wave** (or standing wave) energy is stored rather than transferred to other locations. Oscillations of different amplitudes occur along the length of the wave in a pattern that does not change over time. Points of minimum (ideally zero) amplitude are called **nodes** and occur at every half-wavelength along the wave's extent. Intermediate between the nodes are **antinodes** – points of maximum amplitude. You can think of a stationary wave as being made from two coherent travelling waves that pass through each other in opposite directions (see Figure 1.51).

Key terms

Stationary waves – (or standing waves) are wave motions that store energy rather than transferring energy to other locations.

Nodes – are points along a stationary wave where the displacement amplitude is at a minimum (ideally zero).

Antinodes – are points of maximum amplitude and occur halfway between each pair of nodes.

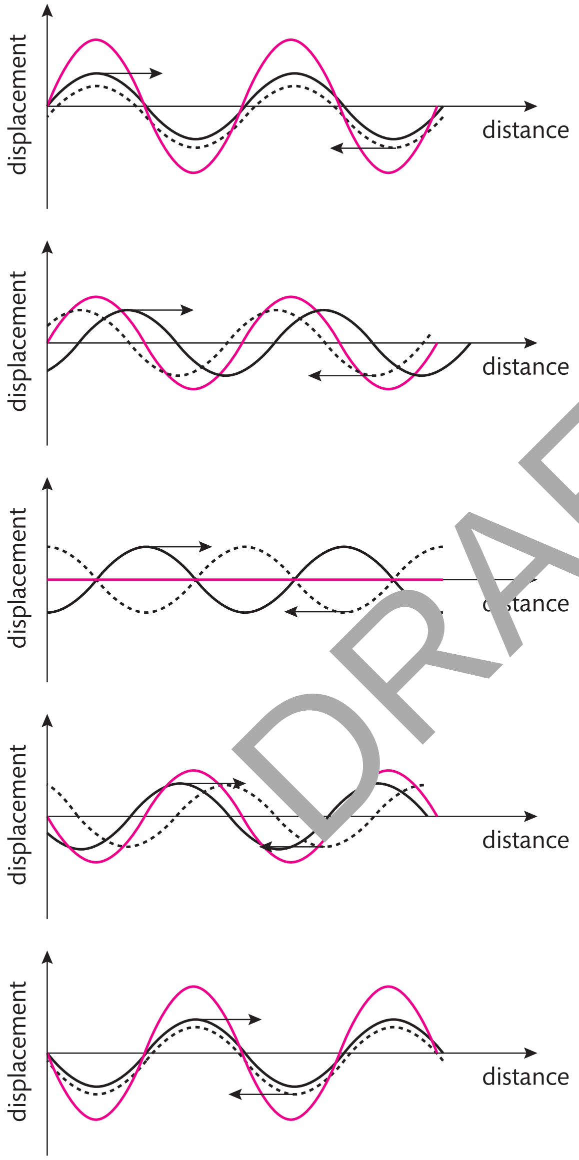


Figure 1.51: When two coherent waves pass through each other, the resultant wave (shown by a red line) is a stationary wave pattern with nodes and antinodes

Stationary wave patterns most often occur in resonators, where the wave motion is confined in a fixed space. The resonator has boundaries that prevent the wave progressing further and reflect its energy back. So you can picture a travelling wave starting down in one direction, hitting the boundary and being reflected back along its own path. The result of superposing these two waves of more or less equal amplitude, but travelling in opposite directions, is a stationary wave pattern. You can demonstrate this in a laboratory by causing vibrations in a stretched string (see Figure 1.52).

Key terms

Resonance – the storing of energy in an oscillation or a stationary wave, the energy coming from an external source of appropriate frequency.

Natural frequency – a resonator has a series of natural frequencies (or 'modes' or 'harmonics') each of which corresponds to an exact number of half wavelengths fitting within its boundaries.

Forcing frequency (or driving frequency)– the frequency of wave energy from an external source that is coupled to a resonator. Efficient energy transfer into the resonator only occurs when this is close to one of the natural frequencies.

The resonator will also have a mechanism for interacting with and absorbing travelling wave energy from outside itself. Small amounts of energy collected over a period of time can be stored up in the stationary wave and build up a much larger amplitude oscillation. This effect is **resonance**. It happens when the wave energy coming in from outside has a **forcing frequency** equal or very close to a **natural frequency** of the resonator.

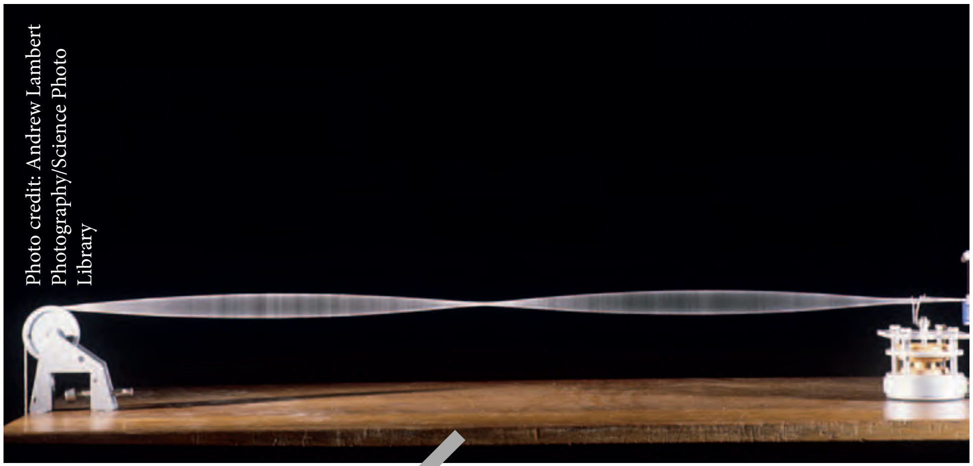


Photo credit: Andrew Lambert Photography/Science Photo Library

Figure 1.52: Stationary waves in a stretched string can be demonstrated with a vibration generator, a pulley and weights to adjust the tension

A resonator can also be set into stationary wave motion by a sudden impact that transfers a large amount of energy in an instant, which is then stored in the wave. This is the case for a gong or bell. Alternatively, resonators can be excited continuously by an external vibration, which is the case for most musical instruments other than percussion instruments.

Musical instruments

Both stringed and wind instruments depend on resonance to produce their musical notes.

In a stretched string, the oscillations are transverse, and the speed, v, at which waves travel down its length, L, depend on the string tension, T, and on the string's mass, m, per unit length, μ ($= \frac{m}{L}$). The wave speed can be calculated using the formula:

$$v = \sqrt{\frac{T}{\mu}}$$

The fixed ends of the string are nodes because they are points of no vibration. So the harmonic with the lowest frequency (the fundamental harmonic) has just one antinode between those nodes and a wavelength equal to twice the length of the string (i.e. $\frac{\lambda_1}{2} = L$). Using the wavelength and the wave speed, you can determine the frequency, f_1, of this fundamental harmonic.

But there are lots of other harmonics with higher frequencies that will fit neatly onto the same stretched string, with first one and then more extra nodes appearing between those at either end. Any number of half wavelengths can be fitted into the string's length, giving harmonics that make the string resonate at frequencies: f_1, $2f_1$, $3f_1$, $4f_1$, ...etc. The smooth, rich tone of a violin and other stringed instruments is due to the fact that bowing the string excites not just the fundamental harmonic frequency, but also smaller amounts of vibration at all these higher harmonics.

Tubes can also act as resonators, and are the basis of all wind instruments.

An open-ended cylinder (pipe) will naturally have maximum oscillations (i.e. oscillation antinodes) at both its open ends, and one or more nodes in between. This gives a set of harmonic frequencies similar to those for a stretched string. However, you cannot control the speed of the waves, which is just fixed at the speed of sound in air, so to tune a pipe you need to alter its length.

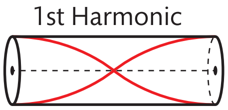


Figure 1.53: Fundamental harmonic of an open-ended pipe

Another useful way to look at sound vibrations in pipes is to think about the wave pressure rather than the oscillation of the molecules of air. Pressure is always 90° out of phase with molecule displacements, so wherever there is an oscillation antinode it corresponds to a pressure node – and vice versa. So open-ended pipes have a pressure node at each end.

However, if you open a hole in the side of a pipe, you can reduce the pressure there to near atmospheric and so force a pressure node (oscillation antinode) to occur there. The woodwind family of instruments operates in just that way – opening and closing holes in the side of the pipe to alter the stationary wave pattern produced inside. The recorder, the flute, the piccolo and many organ pipes all operate as open-ended pipes, with the sound vibrations being excited by blowing air over a sharp edge.

A closed-ended cylinder instrument actually has one end closed but the other open to the air to let the sound out. That makes for a displacement node at the closed end with an antinode at the open end and gives lower musical notes: the fundamental harmonic will have just a quarter wavelength fitting in the pipe.

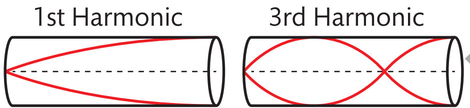


Figure 1.54: Harmonics of a pipe with its left-hand end closed

PAUSE POINT

Brass instruments are closed-ended cylinders, in which different harmonic notes can be blown by the players' altering the pressure and vibration of their lips on the mouthpiece. Which harmonics are possible in a closed-ended instrument?

Hint Sketch the first few harmonics that will fit in a pipe giving a node at one end and antinode at the other end.

Extend Why do brass instruments sound harsher than stringed or woodwind instruments? Think about how the wavelengths and frequencies of the brass instruments' harmonics compare with that of the fundamental harmonic. Which ratios of harmonic are missing?


Photo credit: Lebrecht Music and Arts Photo Library/Alamy Stock Photo

Making music using stationary waves resonance

Other applications of stationary waves

Radio and TV antennas have a reflector element that bounces the incoming waves back and creates a stationary wave pattern. The detector is placed at an antinode position for the particular wavelength of radiation the aerial has been designed to pick up.

In **microwave ovens**, stationary wave patterns caused by reflections for the metal sides of the oven cause hot and cold spots corresponding to antinodes and nodes. Hence the need for a turntable.

Bound electrons in atoms and molecules behave like stationary waves bouncing around in the space they are restricted to by the attraction of the nuclear positive charge. The discrete energy levels that electrons can occupy each correspond to a stationary wave pattern. Wave patterns with higher numbers of nodes correspond to higher energy levels.

Assessment practice 1.16

The free section of a stretched guitar string is 70 cm long and it produces a fundamental harmonic with a frequency of 450 Hz.

a) What is the wavelength of the fundamental vibration in the string?

b) What is the speed of the waves travelling up and down this string?

c) If the mass per unit length of this string is 0.001 kg/m, what must be the tension in the string?

d) Explain why the waves travelling up and down the string produce a stationary wave pattern at that frequency but not at lower or slightly higher frequencies.

e) What is the next higher frequency at which a stationary wave could be formed in this string?

C2 Waves in communication

Fibre optics have become a vital backbone for modern communication systems. In this section you will learn how and why they work, and about their importance in scientific investigations.

The principles of fibre optics

To understand fibre optics, you need first to know about refractive index and total internal reflection.

The laws of refraction

Light (or electromagnetic radiation of other frequencies) travels best through a vacuum. Its rapidly oscillating electric field generates an oscillating magnetic field, and the changing magnetic field in turn generates another nearby oscillating electric field. And so the wave progresses rapidly through space.

When the waves have to travel through matter, their progress is impeded by the electronic charges in the atoms and molecules. Metals, which are full of freely moving electrons, just stop the wave oscillation completely. So the light wave energy is reflected back, just as a sound wave is reflected back at the fixed end of a vibrating string. Metals therefore look shiny and make good mirrors. Many other materials absorb some or all of the light and so look coloured or even black.

In transparent materials, like water, glass and many plastics, the waves are not stopped or absorbed, but they are slowed down. The ratio of the speed of light in vacuum, *c*, to its speed in the material medium, *v*, is called the **refractive index**, *n*, of the medium. That is:

$$n = \frac{c}{v}$$

Discussion

'Refraction' of a marching column – Try acting out this 'thought experiment' with some colleagues.

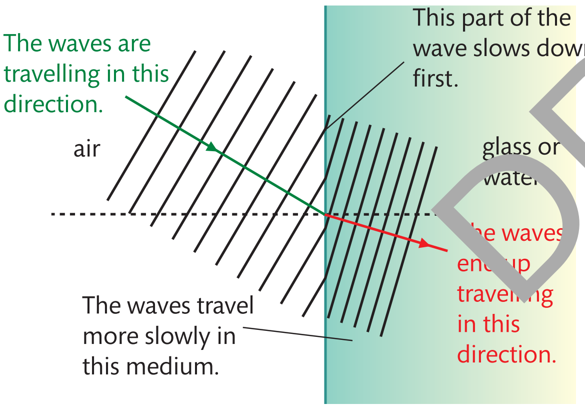
You are marching or walking holding hands with your friends in a line. On a good hard surface you can all keep the same speed, so the line stays straight and moves directly forwards.

Then suppose the person on your left finds herself reaching the edge of the paved area and stepping into long grass. As she slows down she pulls back and twists the line. Then you enter the grass too, and you also slow down.

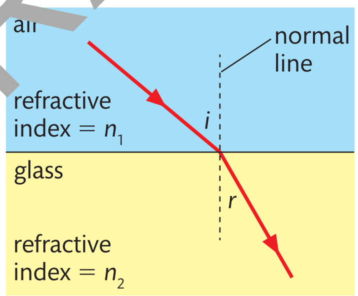
In which direction is the marching line turned?

If there were several rows of marchers, what would happen to the distance between the marching lines as they crossed onto the grass?

Figure 1.55 shows some wave-fronts of light waves entering a smooth-sided glass block at an angle:



▶ **Figure 1.55:** Light wave-fronts refracted as they enter a glass block



▶ **Figure 1.56:** Defining refraction

[redacted]

In a similar way to the thought experiment with a marching column of people, the light wave-fronts are turned as they slow down in the glass. The wave travels more directly into the glass block – i.e. closer to an imaginary line drawn at right-angles to the surface of the block, which is called the **normal** line.

If, instead of wave-fronts, you just draw light rays, the diagram looks like Figure 1.56:

You can label and measure two important angles: the angle of **incidence**, *i*, and the angle of **refraction**, *r*.

When you do the mathematics of this geometry, it turns out that:

$$n = \frac{c}{v} = \frac{\sin i}{\sin r}$$

where *n* is the refractive index of the glass, *v* is the speed of light in the glass, and *c* represents the speed of the light just before it entered the glass, which is actually its speed in air but that is almost the same as its speed in vacuum, hence the use of the letter *c*. So this equation describes how the speed change on entering the glass is what determines the change in direction of the ray – i.e. how much it is refracted (bent) from its incident angle.

Research

Carry out your own practical investigation of refraction, for example, using a light ray box and a glass block with parallel sides. Mark the path of the rays as they enter and leave the block and construct the ray inside the block by joining up the points of entry and exit. Measure the angles of the rays. *(Remember: always draw a 'normal' line perpendicular to the side of the glass block, and measure the angles the rays make with that.)*

What happens to the equation for refraction when you apply it to a ray leaving the glass block? (*Hint:* The angles *i* and *r* seem to change places.) Can you explain why that is?

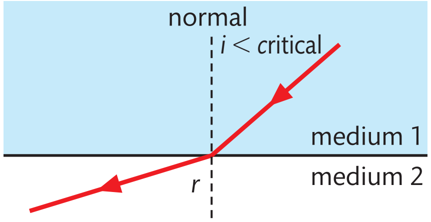
Experiment with using larger angles of incidence for the incoming light ray. What happens to the intensity of the outgoing ray? Can you spot where the light has gone?

Total internal reflection - calculation of critical angles

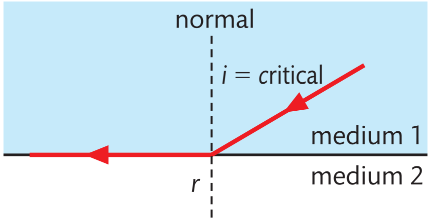
When the light ray comes to the other side of the glass block and tries to leave, back out into the air, or ideally into a vacuum, the light wave will speed up. That makes the wave-front turn back towards the direction in which it was travelling before it entered the glass.

Partial transmission and partial reflection of light at a media interface

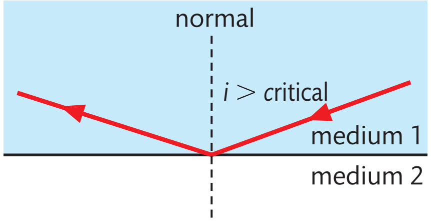
However, speeding up is yet another change, and it is not that easy, particularly if the direction change is substantial. So some of the wave energy gets reflected back inside the glass. This is **internal reflection**. The larger the angles involved, the more light is reflected and the less energy gets through in the refracted beam to escape from the glass into the air.



i is less than critical angle



i equals critical angle



angle of incidence greater
than critical angle

Figure 1.57: Moving from refraction to total internal reflection

When the angle in the glass between the ray and the normal line is increased to a **critical angle**, *C*, the refracted ray is bent so far it would need to run at 90°, that is, along the surface of the glass-air interface. That is not possible, so from that incident angle upwards all the wave energy is reflected internally and there is no refracted beam at all. This effect is called **total internal reflection**. Mathematically, since the sine of 90° is 1, the equation for refraction becomes:

$$\frac{1}{n} = \frac{v}{c} = \sin C \qquad \text{i.e. } \sin C = \frac{1}{n}$$

Assessment practice 1.17

The sparkle of diamonds is due to light being very effectively trapped in them by total internal reflection. The critical angle for a diamond-air interface is only 24.4°. Calculate:

a) the refractive index of diamond

b) the speed of light in diamond.

Optical fibres

Optical fibres are very long thin cylinders of glass or, sometimes, plastic. Light is fed into the cut end of the fibre, so when it hits the sides of the fibre, it almost always does so at angles greater than the critical angle. That means all the rays of light get totally internally reflected and keep bouncing down the length of the fibre. No wave energy gets lost through the walls of the fibre, although as glass is not perfectly transparent, some is gradually absorbed. When the light waves arrive at the far end of the fibre, up to a few

[redacted]

kilometres away, their intensity is still large enough to measure as a signal. If the joints between them are carefully made, optical fibres joined end to end can pass their light signals on to the next stage in a fibre network, again with only a small loss in intensity.

This makes light in optical fibres a much more efficient way of transmitting signals than sending electrical pulses down copper cables. Copper cables suffer from quite large losses due to electrical resistance, meaning that after a few hundred metres most of the signal has been attenuated away and amplifiers are needed to boost it up again.

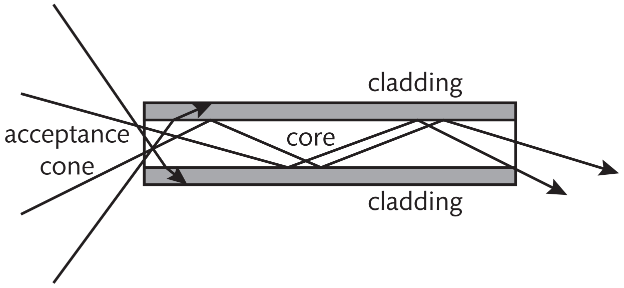


Figure 1.58: Light paths through a multimode optical fibre: total internal reflections off the lower density glass cladding material

Theory into practice

The dense glass core of a communication fibre is not surrounded by air, but by a cladding material. The cladding material is normally another type of glass with a lower refractive index than the core. Light in the fibre is totally internally reflected at the core/cladding interface.

Explain why the critical angle is larger than it would be for a bare glass fibre. (Hint: How much speed change is there for light between cladding and core glass. Compare that with air and glass.)

You might think glass fibres would be far too brittle to use. However, the brittle nature of glass is due to microscopic scratches that are easily introduced onto the outer surface. So, when first drawn, glass fibres (i.e. core plus cladding) are scratch free, strong, flexible and durable. Each glass fibre is given a plastic protective sheathing which keeps it scratch free. Sketch and label a diagram of a communications optical fibre.

Using fibres to carry signals

Installing and jointing optical fibres is a job for specially trained technicians using the right specialist tools and equipment. The jointing can be temporary, using a proprietary connector, or permanent. Permanent joints can be achieved by compression or by a welding process using heat.

Reliability of signal transmission depends on the quality of joints between fibres, both along a cable run and also those to the optical transmitter (an LED or laser) at one end and the optical receiver at the other. These are more or less the only places where substantial losses of signal intensity occur, that is, unless the cable is damaged.

Otherwise the attenuation of signal along a fibre is gradual, and a readable output of 1% or more of the input intensity can still be obtained after a kilometre or more. The optical signal can then be boosted in line by an optical amplifier. These devices are an offshoot of solid state laser technologies, and they increase the intensity of the light beam without needing to convert it back into an electrical signal.

Applications of fibre optics in medicine

Fibre optics are widely used in endoscopes. Endoscopes are optical instruments with long tubes that can be inserted into a body organ through an opening such as the throat, nose, ear canals or anus. These allow a trained medical practitioner to see inside a body organ, for example, the upper oesophagus and stomach or the colon and intestines, without undertaking surgery. Endoscopes are also used during keyhole surgery to guide the use of surgical instruments with remote handling, which are often incorporated into the same tube system.

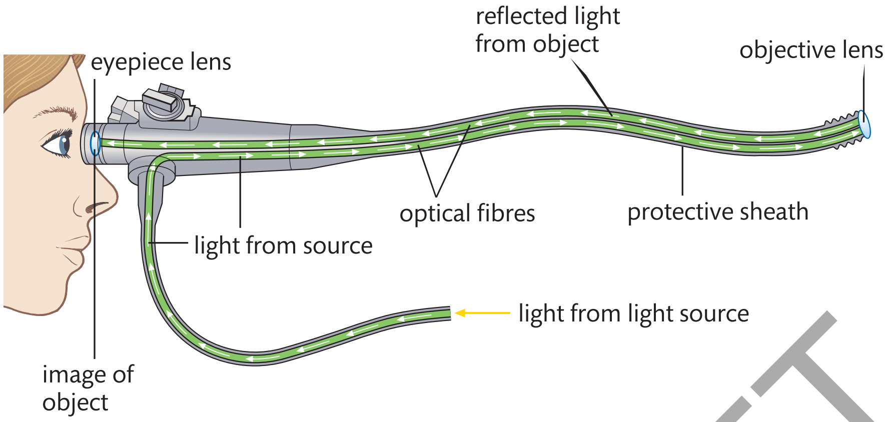


Figure 1.59: Medical endoscope used to reduce the need for invasive surgery

Light is piped in from a source outside the body using a small bundle of optical fibres. The image from inside the body is then focused by an objective lens like that in a camera. To pipe the image back out to be displayed for the medical practitioner, some shorter rigid tubular endoscopes use a series of relay lenses. In longer flexible 'fibrescopes', the image is also conveyed by a second fibre bundle back to an eyepiece or a camera lens.

Each fibre in the bundle is as thin as a human hair and consists of:

- a core
- cladding
- a protective plastic buffer coating.

The image transmitted is pixelated (i.e. formed of coloured dots), since each fibre only transmits one pixel of coloured light. So the resolution of the image depends on the number of fibres in the bundle.

Both the fibre bundles are protected inside the endoscope's long, flexible insertion tube. This tube also carries control lines that can move the bending section at the distal end. That is where the fibres terminate and interface with the lenses, one to distribute the incoming light and the other to collect and focus the image. Sometimes there may be additional control lines to independently move and operate a surgical tool.

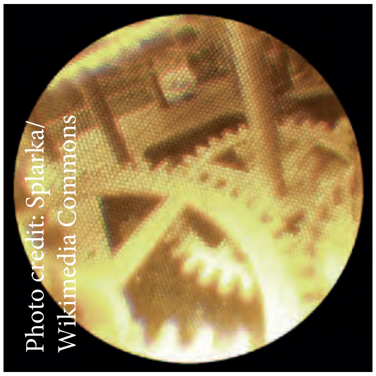


Photo credit: Splarka/ Wikimedia Commons

The image from a low-resolution medical endoscope. Each fibre produces just one dot of colour.

Applications of fibre optics in communication

Analogue and digital signals

The information transferred by each fibre of an endoscope is an **analogue signal.** The colour and intensity of the light from a single point on the image formed by the lens inside the patient's body is directly relayed to reproduce an equivalent image point at the eyepiece or camera outside the body. Just one dot in the picture (i.e. one pixel of information) is delivered by each fibre.

Other examples of analogue signals are:

- the electrical signals made by a microphone, which mimic the shape and intensity of the sound waves they are detecting
- the position of the pointer on a pressure dial gauge
- the waveform displayed on a cathode ray oscilloscope, which copies and shows the variation of an AC voltage with time.

Key terms

Analogue signal – a signal whose strength is proportional to the quantity it is representing.

Digital signal – conveys in binary code a number that represents the size of the measured quantity.

An optical fibre could convey much more information than just one pixel of coloured light. The light travelling down the fibre is a very high frequency wave. However, provided that you operate at significantly below that frequency, you can chop the light signal into on/off bursts and use it to transmit **digital signals**. Digital signals are numbers in binary code, i.e. a series of ones and zeros.

Because of the very high frequencies being used, a huge amount of digital data can be transmitted in a short time. So, instead of just one pixel, a whole television picture can be digitised as a string of numbers that represent the brightness and colour of every pixel in it, and that can arrive in your home travelling along a single digital fibre cable. And not just one TV picture, but a whole host of other channels all available at the same time. Current data networks operate at up to Gigabit per second rates, but higher frequencies are possible with future technological development.

PAUSE POINT Up to about what frequency could you potentially operate an optical switch and send readable signals down a fibre using near infra-red radiation?

Hint Use the formula $v = f\lambda$ to calculate the frequency of near infra-red light, wavelength 830 nm. (The speed of light is approximately 3×10^8 ms^{-1}.) Assume that you should allow at least 100 cycles on followed by 100 cycles off to represent a single bit of data.

Extend What would currently limit the bit rates at which data can be transmitted?

Digitising information not only makes it possible to send more data faster than using analogue transmission. It also makes the transmission much more reliable and interference free. After a kilometre of travel down an optical fibre, the colour (i.e. frequency) of the light will be maintained, but its intensity may have dropped to only about 1% of the original. Nevertheless, if the signal is chopped into digital bursts, you can still reliably read whether it is on or off. That signal can be amplified and no information will have been lost.

Converting a signal from analogue to digital is carried out electronically using an analogue to digital (A to D) converter. At regular intervals this device samples (i.e. measures the size of) the analogue signal. Each sample value is measured and converted into a whole number of units, according to what sensitivity has been. So if the sensitivity is 0.1 volt, then values of 0, 0.1, 0.2, 0.3, 0.5 volt ...etc can be recorded, but no values between these. The sample value is then output as a digital signal which is a binary code number.

Binary code uses base 2 rather than base 10, and so represents any number as a series of ones and zeros. Electrical or optical signals represent this by switching on for a 'one' and off for a 'zero'. The switching is done at a predetermined speed called the clock speed. The clock speed determines how much data can be transmitted in a given time. The sampling rate needs to be set at least an order of magnitude (i.e. 10 times) higher than the frequency of any waveform you want to represent. For example, for sound signals, the sampling rate needs to be set at 10 times the frequency of the highest pitch harmonic within the audible range you want to reproduce. (The reason that voices sound odd on the telephone is because only a limited range of frequencies is reproduced by the digital sampling of telephone systems. This saves data space but loses quality.)

Step by step: Analogue to digital conversion

5 Steps

1 Select a transducer, a device that produces an analogue electrical voltage signal proportional to the quantity you want to measure, for example, a pressure sensor, a thermocouple or thermistor for temperature, a microphone for sound .

2 Connect the output of the transducer to the input of an A to D converter, using a screened cable. To avoid picking up electrical interference the screening must be well earthed.

3 Set up the A to D converter to sample the analogue signal. This is equivalent to taking measurements to plot results out on a voltage-time graph.

4 Select an appropriate sampling rate, which is your sensitivity on the time axis.

5 Select an appropriate sensitivity for the conversion of the voltage signal into a number. The smallest difference you will be able to convey with the digitally converted data is one unit

5 Connect the A to D converter output to a switch/transmitter to send the digital information: either electronically down copper cables; wirelessly using Bluetooth, WiFi or similar protocols; or by optical signals along a fibre network. Optical fibres are free from electromagnetic interference and virtually impossible to hack into.

Case study

Fibre optic broadband networks

Broadband is used as a relative term to indicate the speed and carrying capacity of a data channel. In connection with the internet it has been used to market the improvement from earlier telephone dial up connections, which were very limited and slow. Fibre optic broadband has been progressively replacing copper cable connections with consequent gains in data speed. Speeds of up to 100 Mbit per second to the desktop are now not uncommon. Speeds in the network backbone have, naturally, to be higher.

The name ‘broadband’ derives from the idea of ‘bandwidth’. That is the practice of dividing the signal space up into a number of separate frequency bands. Each band can then be used to carry a separate channel of data, all at the same time. In the case of optical fibres, this is achieved by ‘wave division multiplexing’ (WDM). In WDM, light of several different colours (frequencies) is sent down the same fibre at the same time. Each different frequency of light carries another channel of information.

Multimode fibre is the standard fibre cable used for sending optical signals over short to medium distances – for example, connections to instruments, jumpers in cabinets, small local area networks. (It is also used elsewhere where high light power is needed, for example, in endoscopes.) The optical cores are around 100 μm (microns) in diameter, which is wide enough to allow a variety of separate beams to bounce along the inside of the fibre, striking the outer surface of the core with various different angles. If this is used for distances longer than a few hundred metres, it results in some of the beams travelling a significantly longer path and thus arriving late, which then degrades the quality of the signal. The ‘ones’ and ‘zeros’ of digital signals will thus no longer have sharp beginnings and ends.

Single mode fibre has an even narrower core (8 μm to 10 μm), which is less than ten wavelengths of the infra-red light that is used in them. This means there is just no space for different beams travelling at different angles down the core. Instead, the light wave moves as a single wave-front straight down the centre of the fibre, and all the signal energy reaches the far end of the fibre at the same instant. Millions of kilometres of this high quality cable is laid every year to build the fibre optic networks for telephone, cable TV and broadband internet communications.

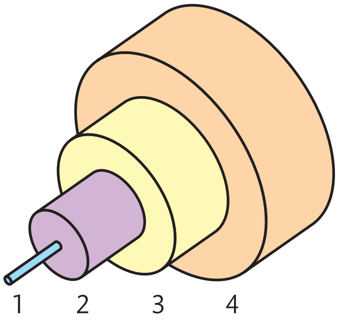


Figure 1.60: Structure of a typical single mode optical fibre cable: 1. Core 8 μm diameter, 2. Cladding 125 μm dia., 3. Buffer 250 μm dia., 4. Jacket 400 μm dia.

Check your knowledge

1 What distances are involved in the local area network (LAN) that serves your laboratory?

2 What broadband data speeds do you need?

3 Would standard multimode fibre be adequate for the LAN cables? For which data connections are you still using copper cables?

Assessment practice 1.18

Describe how a medical endoscope transmits an image from inside the body.

How do the signals in the fibres of an endoscope differ from those transmitted through a fibre optic broadband network?

C3 Use of electromagnetic waves in communication

Speed of electromagnetic waves in a vacuum

Light, and all forms of electromagnetic radiation, travel at the same speed through vacuum: $2.997\,925 \times 10^8\ \mathrm{m\,s^{-1}}$. This is a physical constant value that is usually denoted by the letter, c. The fact that light always travels at this huge speed, and that nothing has ever been observed travelling faster than that, is not a ‘law’. It is an experimentally observed fact, and it is the basis for Einstein’s theories of relativity. Scientists are always on the lookout for any exceptions, but so far none have ever been seen.

Inverse square law for intensity of a wave

Waves transfer energy, and energy is a quantity that is always conserved. Wave-fronts propagating out from a point or a spherical source will themselves be spherical. As each wave-front increases in radius it also increases in area. The formula for the surface area of a sphere of radius r is $4\pi r^2$. The energy in the moving wave-front is distributed over that expanding area, and so its intensity decreases accordingly:

$$I = \frac{k}{r^2}$$

where *I* is intensity of wave, *k* is a constnat and *r* is distance from source.

Regions of the electromagnetic spectrum

Although all types of electromagnetic radiation have the same nature and can be described by the same equations – Maxwell's equations – you experience them quite differently.

- Your eyes can only detect a very small range of frequencies. These are visible light.
- You can sense frequencies just a little lower than that of red light as radiant heat warming you. These are infra-red radiation (IR).
- There are frequencies just above your visible range that can be seen by bees and some other animals, which help plants grow and which cause sunburn. These are ultra-violet light (UV), because the frequencies are above those of violet light.

The remaining types of radiation are named according to how they are produced. At the highest frequencies the frequency ranges for X-rays and for γ-rays (gamma rays) overlap somewhat. X-rays are produced by high energy atomic electron transitions and are just a higher energy version of light and UV radiation. On the other hand, γ-rays come from nuclear disintegrations and from collisions between high energy sub-atomic particles.

For every frequency of radiation, *f*, you can calculate a corresponding wavelength in vacuum, λ, using the speed of light, *c* in the wave equation, $c = f\lambda$.

Table 1.10 shows the main types of electromagnetic radiation, how each is produced and some of the things they are used for.

Table 1.12 SI unit prefixes

Prefix	Symbol	× 10 to the power	Factor
exa	E	18	1 000 000 000 000 000 000
peta	P	15	1 000 000 000 000 000
tera	T	12	1 000 000 000 000
giga	G	9	1 000 000 000
mega	M	6	1 000 000
kilo	k	3	1000
(none)	(none)	0	1
milli	m	-3	0.001
micro	µ	-6	0.000 001
nano	n	-9	0.000 000 001
pico	p	-12	0.000 000 000 001
femto	f	-15	0.000 000 000 000 001
atto	a	-18	0.000 000 000 000 000 001

Table 1.13: Frequencies, sources and applications

The e/m spectrum		Frequency range	Wavelengths	Produced by		Used for		
Radio:	Long wave Medium wave Short wave (HF) VHF UHF	30 kHz to 300 kHz 300 kHz to 3 MHz 3 MHz to 30 MHz 30 MHz to 300 MHz 300 MHz to 3 GHz	10 km to 1 km 1 km to 100 m 100 m to 10 m 10 m to 1 m 1 m to 100 mm	Electronic oscillators coupled to broadcast antennas	Astronomical radio sources for example, neutron stars	Radio and TV broadcasting Mobile phones (UHF) Plasma heating for fusion reactors Industrial ovens		
Microwaves	SHF EHF	3 to 30 GHz 30 to 300 GHz	100 mm to 10 mm 10 mm to 1 mm	Klystron or magnetron tubes, or solid state diodes	Thermal emission Cold ⟶ Very hot	Domestic ovens RADAR Satellite and terrestrial communications links		
Infra-red (IR):	Far Mid-range Near	300 GHz to 30 THz 30 to 120 THz 120 to 400 THz	1 mm to 10 μm 10 to 2.5 μm 2500 to 740 nm	Light emitting diode (LED) or laser		Night vision cameras Optical fibre comms. Movement detectors Remote controls		Spectroscopy
Visible light		400 to 800 THz	740 to 370 nm	Emission by outer electron transitions		Illumination Imaging Signalling	Photosynthesis in plants	
Ultra-violet (UV)		800 THz to 30 PHz (10^{15} Hz)	370 to 10 nm			Insect vision Tanning		
X-rays		30 PHz to 30 EHz (10^{18} Hz)	0.01 to 10 nm	Inner electron excitation and decay		Medical imaging		
γ-rays		generally > 30 EHz	generally < 0.01 nm	Nuclear reactions and particle decays		Radiation sterilisation Medical tracing		

Table 1.14: Frequencies, sources and applications

Application	Power and mode of transmission	Frequency band	How it is used and regulated
Satellite communications	High power signals over very long distances; concentrated by dish antennae.	1 to 40 GHz (Microwaves)	Satellite transponders receive incoming upload signals, amplify them and retransmit them as a download signal on a different frequency band. *For more info search 'satellite frequency bands' on the European Space Agency website www.esa.int*
Mobile phones	High power networked system, range several km.	800 MHz to 2.6 GHz (UHF radio to microwave borderline)	5 or 10 MHz bands allocated to different operators. 2G, 3G and 4G cellular networks offering increasing speeds for data. Higher frequencies have greater data capacity but travel less distance through air and penetrate into buildings less well.
Bluetooth®	Low power device to device links, range up to about 10 m	2.4 to 2.4835 GHz – the Industrial, Scientific, Medical (ISM) unlicensed band – borderline between UHF radio and microwave frequencies	Early Bluetooth devices interfered with Wi-Fi devices because both would use the same channel for an extended period of time. Modern Bluetooth uses frequency-hopping – i.e. broadcasting in short bursts on a number of different frequency channels across the band. This reduces the amount of data lost, and in most cases both Bluetooth and Wi-Fi can maintain service. *For more info search 'Bluetooth and Wi-Fi' at IntelligentHospitalToday.com*
Wi-Fi	Medium power networked system, range ~10 to 100 m		
Infrared	Low power device to device links, range only a few metres	IR wavelength 870 nm or 930 to 950 nm (frequency about 320 THz)	Used for remote controls and for data transfer between computers, phones etc. The longer wavelength band is better because it does not suffer from 'sunlight blinding'. Atmospheric moisture blocks that range in sunlight.

PAUSE POINT What types of material are transparent or partially transparent to radio waves or to microwaves? How do you know that? What benefits or problems might be caused by these waves being diffracted or refracted?

Hint What happens to phone and radio reception inside buildings? What kinds of structure completely block reception inside? In other buildings, what could be the explanation for dead spots and good reception points?

Extend Look at some rooftop UHF TV antennas. Can you identify the reflector? Sketch how the reflector forms a stationary wave pattern along the antenna. Mark where the nodes occur and where the detector dipole is placed.

Worked Example

Calculate how many times smaller the intensity of light falling on a surface 6.0 metres away from a light bulb will be compared with the intensity at just 0.5 metre away.

Answer

$$I = k/r^2$$

So

$$I_1 / I_2 = (r_2/r_1)^2 = (0.5/6.0)^2 = (1/12)^2 = 1/144$$

The intensity will be 144 times smaller at 6.0 m compare with just 0.5 m away from the light bulb.

Assessment practice 1.19

Use the inverse square law to calculate approximately how many times smaller a mobile phone signal will be 2 km from the transmitter compared with reception 500 m from the transmitter mast.

What other factors can affect signal strength?

Assessment practice 1.20

Infra-red and Bluetooth® are used for device to device signal transmission. What are the advantages and disadvantages of each? Compare and contrast these two with mobile phone communications.

Further reading and resources

www.esa.int: European Space Agency. This site gives lots of useful information about satellites and the way they are used for communications. The page on ‘satellite frequency bands’ shows shows how the SHF band is subdivided and what applications use each sub-band.

www.intelligenthospitaltoday.com: Intelligent Hospital Today. Searching this site for ‘Wi-Fi and Bluetooth’ you should find a very useful article that explains the history of Bluetooth® and how it has been adapted to co-exist with Wi-Fi, which uses the same frequency band and so could potentially cause interference and loss of signals.

THINK FUTURE

Photo credit: laviejasirena/Fotolia.com

Sara Logan,

Laboratory Technician in a Consumer Product Company

I started work as a laboratory technician straight from sixth form college. The company I work for makes toiletries such as shampoo and shower gel. They also make cleaning products like bathroom and kitchen cleaners. When I started, I mainly carried out tests on the new products to check they were safe. This included carrying out titrations to ensure the products were the correct pH and that they contained the necessary chemical substances in the right proportions.

This is still a large part of my job but now I have worked here for five years, I am also involved in the development of new products. This means my knowledge of the periodic table and bonding has been invaluable. I know what substances will react and what sort of properties the products will have before I start experimenting. I also understand which chemical substances are safe to react and which ones are too dangerous to use.

We are given a brief that tells us what the product should do, whether it is a tear-free shampoo for babies or a strong cleaner for filthy ovens. We experiment and test using our knowledge of other products. This usually takes many months.

It is important for us to be able to communicate well and work in a team. The whole team work together to research, produce and test a new product. I have to listen carefully to my brief/instructions and ensure I carry them out in order to meet deadlines. I must write a report at the end showing my findings. Sometimes I have to present this report to the client and my boss. I often have to use Excel and PowerPoint to produce my report and presentations.

It is always an exciting day when we realise we have a product that exactly matched the brief.

Focusing your skills

Think about the role of a laboratory technician. Consider the following:

- What types of people will you work with and how will you support them?
- What role will you play in helping them achieve their goals?
- What types of company will you work for? Will you work in research and development or quality assurance for example?
- What skills do you currently have? What skills do you think may need further development?

Getting ready for assessment

This section has been written to help you to do your best when you take the assessment test. Read through it carefully and ask your tutor if there is anything you are still not sure about.

About the test

The assessment test will last 90 minutes and there are a maximum of 90 marks available. The test is in 3 sections and will ask a range of short answer questions as well as some longer answer questions worth up to 6 marks.
Each section, Biology, Chemistry and Physics will include

- short answer questions worth 1-4 marks
- a longer answer question worth 6 marks.

Remember that all the questions are compulsory and you should attempt to answer each one. Consider the question fully and remember to use the key words to describe, explain and analyse. For longer questions you will be required to include a number of explanations to your response: plan your answer and write in detail.

Preparing for the test

To improve your chances on the test you will need to make sure you have revised all the key assessment outcomes that are likely to appear. The assessment outcomes were introduced to you at the start of this unit.

To help plan your revision, it is very useful to know what type of learner you are. Which of the following sounds like it would be most helpful to you?

Type of learner	**Visual**	**Auditory**	**Kinaesthetic**
What it means	Need to see something or picture it, to learn it	Need to hear something to learn it	Learn better when physical activity is involved – learn by doing
How it can help prepare for the test	• Colour code information on your notes • Make short flash cards (so you can picture the notes) • Use diagrams, mind-maps and flowcharts • Use post-it notes to leave visible reminders for yourself	• Read information aloud, then repeat it in your own words • Use word games or mnemonics to help • Use different ways of saying things – different stresses or voices for different things • Record short revision notes to listen to on your phone or computer	• Revise your notes while walking – use different locations for different subjects • Try and connect actions with particular parts of a sequence you need to learn • Record your notes and listen to them while doing chores, exercising, etc. and associate the tasks with the learning

Do not start revision too late! Cramming information is very stressful and does not work.

Useful tips

- **Plan a revision timetable** – schedule each topic you need to revise and try and spend a small time more often on each of them. Coming back to each topic several times will help you to reinforce the key facts in your memory.
- **Take regular breaks** – short bursts of 30–40 minutes revision are more effective than long hours. Remember that most people's concentration lapses after an hour and they need a break.
- **Allow yourself rest** – do not fill all your time with revision. You could schedule one evening off a week, or book in a 'revision holiday' of a few days.
- **Take care of yourself** – stay healthy, rested and eating properly. This will help you to perform at your best. The less stressed you are, the easier you will find it to learn.

Sitting the test

Listen to, and read carefully, any instructions you are given. Lots of marks are often lost because people do not read questions properly and then do not complete their answers correctly.

Most questions contain command words. Understanding what these words mean will help you understand what the question is asking you to do.

Remember the number of marks can relate to the number of answers you may be expected to give. If a question asks for two examples, do not only give one! Similarly, do not offer more information than the question needs. If there are two marks for two examples, do not give four examples.

Planning your time is an important part of succeeding on a test. Work out what you need to answer and then organise your time. You should spend more time on longer questions. Set yourself a timetable for working through the test and then stick to it. Do not spend ages on a short 1 or 2 mark question and then find you only have a few minutes for a longer 4 or 6 mark questions. It is useful when reading through a question to write down notes on a blank page. This way you can write down all the key words and information required and use these to structure an answer.

If you are writing an answer to a longer question, try and plan your answers before you start writing. Have a clear idea of the point your answer is making, and then make sure this point comes across in everything you write, so that it is all focused on answering the question you have been set.

If you finish early, use the time to re-read your answers and make any corrections. This could really help make your answers even better and could make a big difference in your final mark.

Hints and tips for tests

- Revise all the key areas likely to be covered. Draw up a checklist to make sure you do not forget anything!
- Know the time of the test and arrive early and prepared.
- Ensure that you have eaten before the test and that you feel relaxed and fresh.
- Read each question carefully before you answer it to make sure you understand what you have to do.
- Make notes as you read through the question and use these to structure your answer.

- Try answering all the simpler questions first then come back to the harder questions. This should give you more time for the harder questions.
- Remember you cannot lose marks for a wrong answer, but you cannot gain any marks for a blank space!

Q. Ethane and fluoromethane have similar-sized molecules.

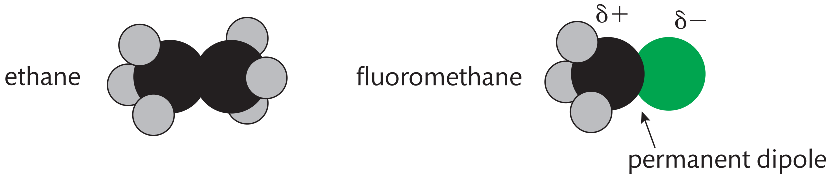


Figure 1.23: Fluoromethane and ethane

Explain why fluoromethane has a higher boiling point than ethane. (2)

Fluorine has the highest electronegativity of all elements.
This means there is a larger permanent dipole on the fluoromethane.

This is an 'explain' question. The examiner is looking for a fact and a reason or a 'because'. The question is worth two marks so two points must be made to gain these marks. The answer above would be worth both marks.

This is a 6 mark levelled question. It is worth 2 pass marks, 2 merit marks and 2 distinction marks. You gain marks for showing understanding rather than there being one mark per point. The more detailed and in-depth your discussion, the more likely you are to gain 6 marks. You would be expected to use all your knowledge about fluorine to answer this question. You should consider position in the periodic table, reactivity, bonding.

Q. Potassium fluoride has many uses in manufacturing, synthesising and refining. It converts chlorocarbons to fluorocarbons. It is used to etch glass and in the making of disinfectants. However, potassium fluoride is not made in the laboratory by reacting potassium with fluorine as this reaction is not safe.

Discuss why it is not safe to react fluorine with an alkali metal in a school laboratory. (6)

Question Number	**Answer**	**Mark**
	Indicative content ▸ Reactions between fluorine and alkali metals are vigorous and explosive ▸ Group 1 metals are the most reactive metals/as you go left to right on periodic table metals become less reactive ▸ As you go down group 1 the metals become more reactive ▸ Group 1 metals only have 1 electron in their outer shell ▸ Reactivity increases in the non-metals as you go from left to right across the period ▸ Reactivity decreases as you go down group 7/ fluorine is the most reactive non-metal ▸ Fluorine has a small radius because positive protons strongly attract negative electrons to it ▸ It has 5 electrons in its 2*p* sub-shell ▸ The *p* sub-shell can hold 6 electrons so fluorine is close to its ideal electron configuration ▸ Strong attraction by positive nucleus due to small size/one shell shielding ▸ Fluorine is a strong oxidising agent ▸ It forms strong ionic compounds with metals. ▸ Fluorine easily gains one electron and alkali metals easily lose an electron	(6)
0	No rewardable content.	0
Pass level	A simple description of why it is unsafe. Learner will use the periodic table to show that the reactants are very reactive.	1–2
Merit level	Links ideas of electronic structure to reaction.	3–4
Distinction level	Links ideas of electronic structure to reaction in detail. May discuss strength of oxidising agent to reaction.	5–6

Image credit: adapted from http://www.chemguide.co.uk/atoms/bonding/vdwstrengths.html

Ans 1. It is not safe to react fluorine with an alkali metal because alkali metals and fluorine are both very reactive meaning that there might be an explosion. This is because alkali metals are on the left hand side of the table and the metals get less reactive as they go across the periodic table.

This would be a pass-level answer. The candidate knows that an explosion will occur and knows how to use the periodic table to show this. The answer is quite simple and shows little understanding of why the elements are so reactive.

Ans 2 Group 1 metals are highly reactive as they only need to lose one metal to form an ionic bond. The repulsion from the positive nucleus makes this easy to react. Fluorine is the most reactive non- metal. It is at the top of group 7 and the trend is that reactivity decreases down the group. Fluorine has a small radius it only has 2 shells so it is already small and the protons in nucleus are able to draw the electrons to it strongly making it even smaller. It has outer electron configuration of $2s^2$ $2p^5$. This means it has only got to get one electron to get a full outer shell. The small radius and only one shell shielding the nucleus means the electron on the metal is strongly attracted to the charge on the positive. Fluorine is a very strong oxidising agent. All of these things means that when fluorine reacts with a group 1 metal the reaction will be vigorous and probably explosive so it is not safe to carry out.

This would be a distinction-level answer. The candidate has discussed reactivity of alkali metals in general in relation to their position on the periodic table. They have discussed the electronic structure of fluorine in detail and have started to relate this to reactivity. They then have said why this makes the reaction unsafe. The ideas are mostly quite detailed and are linked.