Chemistry Grade 10 [CAPS]

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C O N N E X I O N S

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Chapter 1

Classification of Matter[1]

1.1 Introduction

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

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

Some of the properties of matter that you should know are:

- Materials can be strong and resist bending (e.g. iron rods, cement) or weak (e.g. fabrics)
- Materials that conduct heat (e.g. metals) are called thermal conductors. Materials that conduct electricity are electrical conductors.
- Brittle materials break easily. Materials that are malleable can be easily formed into different shapes. Ductile materials are able to be formed into long wires.
- Magnetic materials have a magnetic field.
- Density is the mass per unit volume. An example of a dense material is concrete.
- The boiling and melting points of substance help us to classify substances as solids, liquids or gases at a specific temperature.

The diagram below shows one way in which matter can be classified (grouped) according to its different properties. As you read further in this chapter, you will see that there are also other ways of classifying materials, for example according to whether or not they are good electrical conductors.

[1]This content is available online at <http://cnx.org/content/m38118/1.5/>.

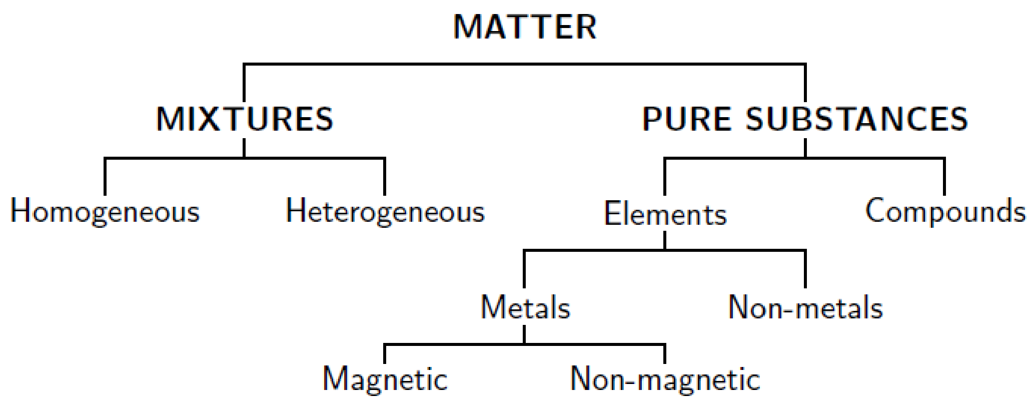

Figure 1.1: The classification of matter

Discussion: Everyday materials
In groups of 3 or 4 look at the labels of medicines, food items, and any other items that you use often. What can you tell about the material inside the container from the list of ingredients? Why is it important to have a list of ingredients on the materials that we use? Do some research on the safety data of the various compounds in the items that you looked at. Are the compounds in the items safe to use? In the food items, what preservatives and additives are there? Are these preservatives and additives good for you? Are there natural alternatives (natural alternatives are usually used by indigenous people groups)?

1.2 Mixtures

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

Definition 1.1: Mixture
A **mixture** is a combination of two or more substances, where these substances are not bonded (or joined) to each other.

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

- *are not in a fixed ratio* Imagine, for example, that you have a 250 ml beaker of water. It doesn't matter whether you add 20 g, 40 g, 100 g or any other mass of sand to the water; it will still be called a mixture of sand and water.
- *keep their physical properties* In the example we used of the sand and water, neither of these substances has changed in any way when they are mixed together. Even though the sand is in water, it still has the same properties as when it was out of the water.
- *can be separated by mechanical means* To separate something by 'mechanical means', means that there is no chemical process involved. In our sand and water example, it is possible to separate the mixture by simply pouring the water through a filter. Something *physical* is done to the mixture, rather than something *chemical*.

Some other examples of mixtures include blood (a mixture of blood cells, platelets and plasma), steel (a mixture of iron and other materials) and the gold that is used to make jewellery. The gold in jewellery is not pure gold but is a mixture of metals. The amount of gold in the jewellery is measured in *karats* (24 karat would be pure gold, while 18 karat is only 75% gold).

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

1.2.1 Heterogeneous mixtures

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

Definition 1.2: Heterogeneous mixture
A heterogeneous mixture is one that is non-uniform and the different components of the mixture can be seen.

1.2.2 Homogeneous mixtures

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

Definition 1.3: Homogeneous mixture
A homogeneous mixture is one that is uniform, and where the different components of the mixture cannot be seen.

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

Activity: Classifying materials
Look around your classroom or school. Make a list of all the different materials that you see around you. Try to work out why a particular material was used. Can you classify all the different materials used according to their properties? On your way to school or at home or in the shops, look at the different materials that are used. Why are these materials chosen over other materials?

Activity: Making mixtures
Make mixtures of sand and water, potassium dichromate and water, iodine and ethanol, iodine and water. Classify these as heterogeneous or homogeneous. Try to make mixtures using other substances. Are the mixtures that you have made heterogeneous or homogeneous? Give reasons for your choice.

1.2.3 Mixtures

1. Which of the following substances are *mixtures*?
   a. tap water

b. brass (an alloy of copper and zinc)
c. concrete
d. aluminium
e. Coca cola
f. distilled water

2. In each of the examples above, say whether the mixture is homogeneous or heterogeneous.
Click here for the solution[2]

1.3 Pure Substances: Elements and Compounds

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

One way to determine if a substance is pure is to look at its melting or boiling point. Pure substances will have a sharply defined melting or boiling point (i.e. the melting or boiling point will be a single temperature rather than a range of temperatures.) Impure substances have a temperature range over which they melt or boil. We can also use chromatography to determine if a substance is pure or not. Chromatography is the process of separating substances into their individual components. If a substance is pure then chromatography will only produce one substance at the end of the process. If a substance is impure then several substances will be seen at the end of the process.

1.3.1 Activity: Smartie Chromatography

You will need filter paper (or chromatography paper), some smarties in different colours, water and an eye dropper.

Place a smartie in the center of a piece of filter paper. Carefully drop a few drops of water onto the smartie. You should see rings of different colour forming around the smartie. Each colour is one of the individual colours that are used to make up the colour of the smartie.

1.3.2 Elements

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

Definition 1.4: Element
An element is a substance that cannot be broken down into other substances through chemical means.

There are 112 officially named elements and about 118 known elements. Most of these are natural, but some are man-made. The elements we know are represented in the **Periodic Table of the Elements**, where each element is abbreviated to a **chemical symbol**. Examples of elements are magnesium (Mg), hydrogen (H), oxygen (O) and carbon (C). On the Periodic Table you will notice that some of the abbreviations do not seem to match the elements they represent. The element iron, for example, has the chemical formula Fe. This is because the elements were originally given Latin names. Iron has the abbreviation Fe because its Latin name is 'ferrum'. In the same way, sodium's Latin name is 'natrium' (Na) and gold's is 'aurum' (Au).

NOTE: Recently it was agreed that two more elements would be added to the list of officially named elements. These are elements number 114 and 116. The proposed name for element 114 is flerovium and for element 116 it is moscovium. This brings the total number of officially named elements to 114.

[2] http://www.fhsst.org/llm

1.3.3 Compounds

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

Definition 1.5: Compound

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

Figure 1.2 might help you to understand the difference between the terms *element*, *mixture* and *compound*. Iron (Fe) and sulphur (S) are two elements. When they are added together, they form a *mixture* of iron and sulphur. The iron and sulphur are not joined together. However, if the mixture is heated, a new *compound* is formed, which is called iron sulphide (FeS). In this compound, the iron and sulphur are joined to each other in a ratio of 1:1. In other words, one atom of iron is joined to one atom of sulphur in the compound iron sulphide.

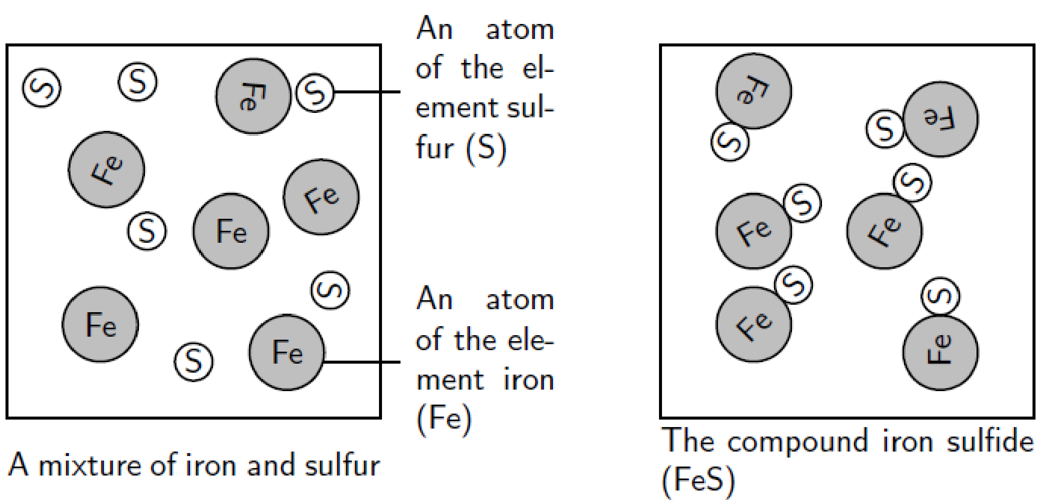

Figure 1.2: Understanding the difference between a mixture and a compound

Figure 1.2 shows the microscopic representation of mixtures and compounds. In a microscopic representation we use circles to represent different elements. To show a compound, we draw several circles joined together. Mixtures are simply shown as two or more individual elements in the same box. The circles are not joined for a mixture.

We can also use symbols to represent elements, mixtures and compounds. The symbols for the elements are all found on the periodic table. Compounds are shown as two or more element names written right next to each other. Subscripts may be used to show that there is more than one atom of a particular element. (e.g. H_2O or $NaCl$). Mixtures are written as: a mixture of element (or compound) A and element (or compound) B. (e.g. a mixture of Fe and S).

One way to think of mixtures and compounds is to think of buildings. The building is a mixture of different building materials (e.g. glass, bricks, cement, etc.). The building materials are all compounds. You

can also think of the elements as Lego blocks. Each Lego block can be added to other Lego blocks to make new structures, in the same way that elements can combine to make compounds.
Activity: Using models to represent substances
Use coloured balls and sticks to represent elements and compounds. Think about the way that we represent substances microscopically. Would you use just one ball to represent an element or many? Why?

1.3.3.1 Elements, mixtures and compounds

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

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

Table 1.1

Click here for the solution[3]

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

   Click here for the solution[4]

1.4 Giving names and formulae to substances

Think about what you call your friends. Their full name is like the substances name and their nickname is like the substances formulae. Without these names your friends would have no idea which of them you are referring to. In the same way scientists like to have a consistent way of naming things and a short way of describing the thing being named. This helps scientists to communicate efficiently.

It is easy to describe elements and mixtures. We simply use the names that we find on the periodic table for elements and we use words to describe mixtures. But how are compounds named? In the example of iron sulphide that was used earlier, which element is named first, and which 'ending' is given to the compound name (in this case, the ending is -ide)?

The following are some guidelines for naming compounds:

1. The compound name will always include the **names of the elements** that are part of it.
   - A compound of **iron** (Fe) and *sulphur* (S) is **iron** *sulph*ide (FeS)
   - A compound of **potassium** (K) and *bromine* (Br) is **potassium** *brom*ide (KBr)
   - A compound of **sodium** (Na) and *chlorine* (Cl) is **sodium** *chlor*ide ($NaCl$)

[3]http://www.fhsst.org/lly
[4]http://www.fhsst.org/llV

2. In a compound, the element that is on the left of the Periodic Table, is used *first* when naming the compound. In the example of $NaCl$, sodium is a group 1 element on the left hand side of the table, while chlorine is in group 7 on the right of the table. Sodium therefore comes first in the compound name. The same is true for FeS and KBr.
3. The **symbols** of the elements can be used to represent compounds e.g. FeS, $NaCl$, KBr and H_2O. These are called **chemical formulae**. In the first three examples, the ratio of the elements in each compound is 1:1. So, for FeS, there is one atom of iron for every atom of sulphur in the compound. In the last example (H_2O) there are two atoms of hydrogen for every atom of oxygen in the compound.
4. A compound may contain **compound ions**. An ion is an atom that has lost (positive ion) or gained (negative ion) electrons. Some of the more common compound ions and their formulae are given below.

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

Table 1.2

5. When there are only two elements in the compound, the compound is often given a **suffix** (ending) of -ide. You would have seen this in some of the examples we have used so far. For compound ions, when a non-metal is combined with oxygen to form a negative ion (anion) which then combines with a positive ion (cation) from hydrogen or a metal, then the suffix of the name will be ...ate or ...ite. $\mathrm{NO_3^-}$ for example, is a negative ion, which may combine with a cation such as hydrogen (HNO_3) or a metal like potassium ($\mathrm{KNO_3}$). The $\mathrm{NO_3^-}$ anion has the name nitr**ate**. $\mathrm{SO_3^{2-}}$ in a formula is sulph**ite**, e.g. sodium sulphite (Na_2SO_3).
   SO_4^{2-} is sulph**ate** and PO_4^{3-} is phosph**ate**.
6. **Prefixes** can be used to describe the ratio of the elements that are in the compound. You should know the following prefixes: 'mono' (one), 'di' (two) and 'tri' (three).
   - CO (carbon monoxide) - There is one atom of oxygen for every one atom of carbon
   - NO_2 (nitrogen dioxide) - There are two atoms of oxygen for every one atom of nitrogen
   - SO_3 (sulphur trioxide) - There are three atoms of oxygen for every one atom of sulphur

The above guidelines also help us to work out the formula of a compound from the name of the compound.

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

We can use these rules to help us name both ionic compounds and covalent compounds (more on these compounds will be covered in a later chapter). However, covalent compounds are often given other names by scientists to simplify the name. For example, if we have 2 hydrogen atoms and one oxygen atom the above naming rules would tell us that the substance is dihydrogen monoxide. But this compound is better known as water! Or if we had 1 carbon atom and 4 hydrogen atoms then the name would be carbon tetrahydride, but scientists call this compound methane.

Exercise 1.1: Naming compounds ***(Solution on p. 21.)***

What is the chemical name for

a. $K\mathrm{MnO}_4$
b. NH_4Cl

Exercise 1.2 ***(Solution on p. 21.)***

Write the chemical formulae for:

a. sodium sulphate
b. potassium chromate

1.4.1 Naming compounds

1. The formula for calcium carbonate is $CaCO_3$.
   a. Is calcium carbonate a mixture or a compound? Give a reason for your answer.
   b. What is the ratio of $Ca : C : O$ atoms in the formula?

   Click here for the solution[5]
2. Give the name of each of the following substances.
   a. KBr
   b. HCl
   c. KMnO_4
   d. NO_2
   e. $\mathrm{NH}_4\mathrm{OH}$
   f. $Na_2\mathrm{SO}_4$

   Click here for the solution[6]
3. Give the chemical formula for each of the following compounds.
   a. potassium nitrate
   b. sodium iodide
   c. barium sulphate
   d. nitrogen dioxide
   e. sodium monosulphate

   Click here for the solution[7]

[5]http://www.fhsst.org/llp
[6]http://www.fhsst.org/lld
[7]http://www.fhsst.org/llv

4. Refer to the diagram below, showing sodium chloride and water, and then answer the questions that follow.

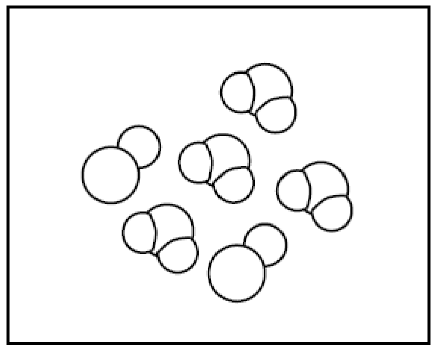

Figure 1.3

   a. What is the chemical formula for water?
   b. What is the chemical formula for sodium chloride?
   c. Label the water and sodium chloride in the diagram.
   d. Give a description of the picture. Focus on whether there are elements or compounds and if it is a mixture or not.

   Click here for the solution[8]
5. What is the formula of this molecule?

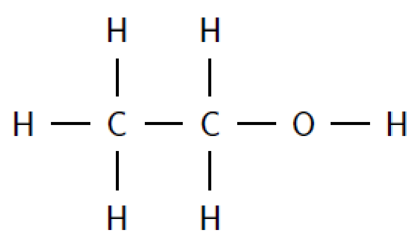

Figure 1.4

   a. C_6H_2O
   b. C_2H_6O
   c. $2C$6HO
   d. $_2CH_6O$

   Click here for the solution[9]

[8] http://www.fhsst.org/llL
[9] http://www.fhsst.org/llf

1.5 Metals, Metalloids and Non-metals

The elements in the Periodic Table can also be divided according to whether they are **metals**, **metalloids** or **non-metals**. On the right hand side of the Periodic Table you can draw a 'zigzag' line (This line starts with Boron (B) and goes down to Polonium (Po). This line separates all the elements that are metals from those that are non-metals. Metals are found on the left of the line, and non-metals are those on the right. Along the line you find the metalloids. You should notice that there are more metals then non-metals. Metals, metalloids and non-metals all have their own specific properties.

1.5.1 Metals

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

- *Thermal conductors* Metals are good conductors of heat. This makes them useful in cooking utensils such as pots and pans.
- *Electrical conductors* Metals are good conductors of electricity. Metals can be used in electrical conducting wires.
- *Shiny metallic lustre* Metals have a characteristic shiny appearance and so are often used to make jewellery.
- *Malleable* This means that they can be bent into shape without breaking.
- *Ductile* Metals (such as copper) can be stretched into thin wires, which can then be used to conduct electricity.
- *Melting point* Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged.

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

1.5.1.1 Group Work : Looking at metals

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

1.5.2 Non-metals

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

1.5.3 Metalloids

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

You should now be able to take any material and determine whether it is a metal, non-metal or metalloid simply by using its properties.

1.6 Electrical conductors, semi-conductors and insulators

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

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

Definition 1.6: Conductors and insulators
A conductor allows the easy movement or flow of something such as heat or electrical charge through it. Insulators are the opposite to conductors because they *inhibit* or reduce the flow of heat, electrical charge, sound etc through them.

Think about the materials around you. Are they electrical conductors or not? Why are different materials used? Think about the use of semiconductors in electronics? Can you think of why they are used there?

1.6.1 Experiment : Electrical conductivity

Aim:

To investigate the electrical conductivity of a number of substances
Apparatus:

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

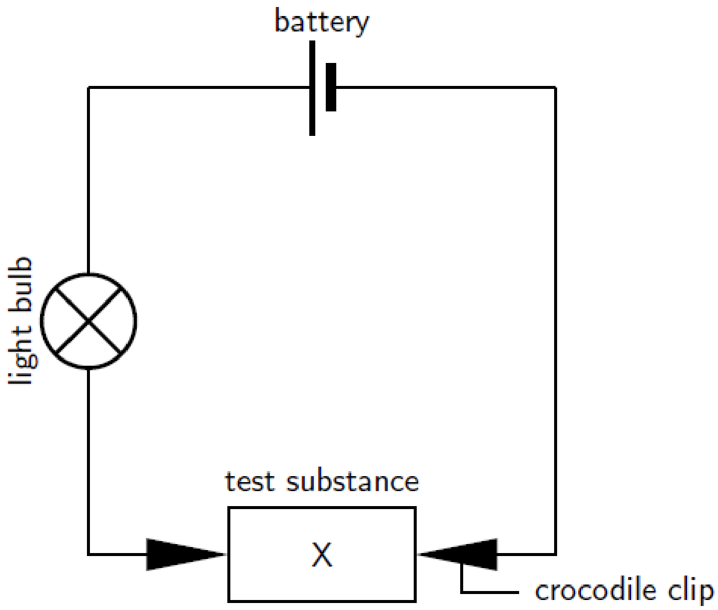

Figure 1.5

Method:

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

Results:

Record your results in the table below:

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

Table 1.3

Conclusions:

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

The following simulation allows you to work through the above activity. For this simulation use the grab bag option to get materials to test. Set up the circuit as described in the activity.

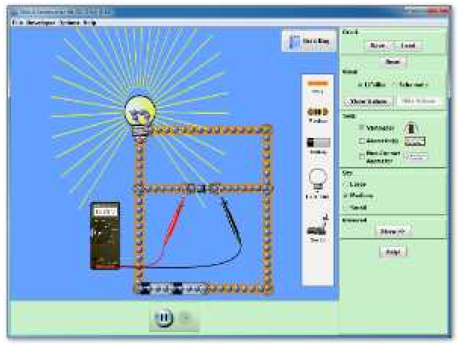

Figure 1.6

run demo[10]

1.7 Thermal Conductors and Insulators

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

1.7.1 Demonstration : Thermal conductivity

Aim:

To demonstrate the ability of different substances to conduct heat.
Apparatus:

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

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

Results:

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

[10]http://phet.colorado.edu/sims/circuit-construction-kit/circuit-construction-kit-dc_en.jnlp

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

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

1.7.2 Investigation : A closer look at thermal conductivity

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

Material	**Thermal Conductivity** ($W \cdot m^{-1} \cdot K^{-1}$)
Silver	429
Stainless steel	16
Standard glass	1.05
Concrete	0.9 - 2
Red brick	0.69
Water	0.58
Snow	0.25 - 0.5
Wood	0.04 - 0.12
Polystyrene	0.03
Air	0.024

Table 1.4

Use this information to answer the following questions:

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

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

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

1.8 Magnetic and Non-magnetic Materials

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

Definition 1.7: Magnetism
Magnetism is one of the phenomena by which materials exert attractive or repulsive forces on other materials.

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

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

1.8.1 Investigation : Magnetism

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

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

Object	Magnetic or non-magnetic

Table 1.5

1.8.2 Group Discussion : Properties of materials

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

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

1.9 Seperating mixtures - Not in CAPS - included for completeness

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

- *Filtration* A piece of filter paper in a funnel can be used to separate a mixture of sand and water.
- *Heating / evaporation* Heating a solution causes the liquid (normally water) to evaporate, leaving the other (solid) part of the mixture behind. You can try this using a salt solution.
- *Centrifugation* This is a laboratory process which uses the centrifugal force of spinning objects to separate out the heavier substances from a mixture. This process is used to separate the cells and plasma in blood. When the test tubes that hold the blood are spun round in the machine, the heavier cells sink to the bottom of the test tube. Can you think of a reason why it might be important to have a way of separating blood in this way?
- *Dialysis* This is an interesting way of separating a mixture because it can be used in some important applications. Dialysis works using a process called *diffusion*. Diffusion takes place when one substance in a mixture moves from an area where it has a high concentration to an area where its concentration is lower. When this movement takes place across a semi-permeable membrane it is called osmosis. A semi-permeable membrane is a barrier that lets some things move across it, but not others. This process is very important for people whose kidneys are not functioning properly, an illness called *renal failure*.

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

1.9.1 Investigation : The separation of a salt solution

Aim:

To demonstrate that a homogeneous salt solution can be separated using physical methods.
Apparatus:

glass beaker, salt, water, retort stand, bunsen burner.
Method:

1. Pour a small amount of water (about 20 ml) into a beaker.
2. Measure a teaspoon of salt and pour this into the water.

3. Stir until the salt dissolves completely. This is now called a *salt solution*. This salt solution is a homogeneous mixture.
4. Place the beaker on a retort stand over a bunsen burner and heat gently. You should increase the heat until the water almost boils.
5. Watch the beaker until all the water has evaporated. What do you see in the beaker?

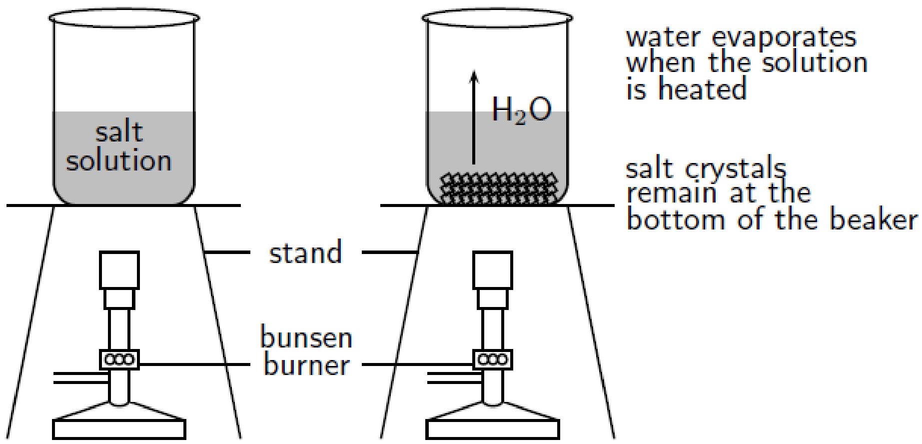

Figure 1.7

Results:

The water evaporates from the beaker and tiny grains of salt remain at the bottom. (You may also observe grains of salt on the walls of the beaker.)
Conclusion:

The salt solution, which is a homogeneous mixture of salt and water, has been separated using heating and evaporation.

1.9.2 Discussion : Separating mixtures

Work in groups of 3-4

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

The following presentation provides a summary of the classification of matter.

This media object is a Flash object. Please view or download it at
<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=classificationofsubstances-100511050657-phpapp02&stripped_title=classification-of-substances-4047520&userName=kwarne>

Figure 1.8

1.10 Summary

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

- Materials may also be classified as **thermal conductors** or **thermal insulators** depending on whether or not they are able to conduct heat.
- Materials may also be either **magnetic** or **non-magnetic**.

1.10.1 Summary

1. For each of the following **multiple choice** questions, choose *one* correct answer from the list provided.

    a. Which of the following can be classified as a mixture:
        a. sugar
        b. table salt
        c. air
        d. iron

        Click here for the solution[11]

    b. An element can be defined as:
        a. A substance that cannot be separated into two or more substances by ordinary chemical (or physical) means
        b. A substance with constant composition
        c. A substance that contains two or more substances, in definite proportion by weight
        d. A uniform substance

        Click here for the solution[12]

2. Classify each of the following substances as an *element*, a *compound*, a *solution* (homogeneous mixture), or a *heterogeneous mixture*: salt, pure water, soil, salt water, pure air, carbon dioxide, gold and bronze. Click here for the solution[13]
3. Look at the table below. In the first column (A) is a list of substances. In the second column (B) is a description of the group that each of these substances belongs in. Match up the *substance* in Column A with the *description* in Column B.

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

Table 1.6

Click here for the solution[14]

4. You are given a test tube that contains a mixture of iron filings and sulphur. You are asked to weigh the amount of iron in the sample.
    a. Suggest one method that you could use to separate the iron filings from the sulphur.
    b. What property of metals allows you to do this?

    Click here for the solution[15]

[11] http://www.fhsst.org/ll6
[12] http://www.fhsst.org/llF
[13] http://www.fhsst.org/llG
[14] http://www.fhsst.org/ll7
[15] http://www.fhsst.org/llA

5. Given the following descriptions, write the chemical formula for each of the following substances:
    a. silver metal
    b. a compound that contains only potassium and bromine
    c. a gas that contains the elements carbon and oxygen in a ratio of 1:2

    Click here for the solution[16]
6. Give the names of each of the following compounds:
    a. $NaBr$
    b. $BaSO_4$
    c. SO_2

    Click here for the solution[17]
7. For each of the following materials, say what properties of the material make it important in carrying out its particular function.
    a. **tar** on roads
    b. **iron** burglar bars
    c. **plastic** furniture
    d. **metal** jewellery
    e. **clay** for building
    f. **cotton** clothing

    Click here for the solution[18]

[16]http://www.fhsst.org/llo
[17]http://www.fhsst.org/lls
[18]http://www.fhsst.org/llH

Solutions to Exercises in Chapter 1

Solution to Exercise 1.1 (p. 8)

Step 1. For a) we have potassium and the permanganate ion. For b) we have the ammonium ion and chlorine.
Step 2. For a) we list the potassium first and the permanganate ion second. So a) is potassium permanganate. For b) we list the ammonium ion first and change the ending of chlorine to -ide. So b) is ammonium chloride.

Solution to Exercise 1.2 (p. 8)

Step 1. In part a) we have Na^{+} (sodium) and SO_4^{2-} (sulphate). In part b) we have K^{+} (potassium) and CrO_4^{2-} (chromate)
Step 2. In part a) the charge on sodium is $+1$ and the charge on sulphate is -2, so we must have two sodiums for every sulphate. In part b) the charge on potassium is $+1$ and the charge on chromate is -2, so we must have two potassiums for every chromate.
Step 3. a) is $Na_2\mathrm{SO}_4$ and b) is K_2CrO_4

Chapter 2

States of matter and the kinetic molecular theory[1]

2.1 Introduction

In this chapter we will explore the states of matter and then look at the kinetic molecular theory. Matter exists in three states: solid, liquid and gas. We will also examine how the kinetic theory of matter helps explain boiling and melting points as well as other properties of matter.

> NOTE: When a gas is heated above a certain temperature the electrons in the atoms start to leave the atoms. The gas is said to be ionised. When a gas is ionised it is known as a plasma. Plasmas share many of the properties of gases (they have no fixed volume and fill the space they are in). This is a very high energy state and plasmas often glow. Ionisation is the process of moving from a gas to a plasma and deionisation is the reverse process. We will not consider plasmas further in this chapter.

2.2 States of matter

All matter is made up of particles. We can see this when we look at diffusion. Diffusion is the movement of particles from a high concentration to a low concentration. Diffusion can be seen as a spreading out of particles resulting in an even distribution of the particles. You can see diffusion when you place a drop of food colouring in water. The colour slowly spreads out through the water. If matter were not made of particles then we would only see a clump of colour when we put the food colouring in water, as there would be nothing that could move about and mix in with the water. The composition of matter will be looked at in What are the objects around us made of? (Chapter 6).

Diffusion is a result of the constant thermal motion of particles. In Section 2.3 (The Kinetic Theory of Matter) we will talk more about the thermal motion of particles.

In 1828 Robert Brown observed that pollen grains suspended in water moved about in a rapid, irregular motion. This motion has since become known as Brownian motion. Brownian motion is essentially diffusion of many particles.

Matter exists in one of three states, namely solid, liquid and gas. Matter can change between these states by either adding heat or removing heat. This is known as a change of state. As we heat an object (e.g. water) it goes from a solid to a liquid to a gas. As we cool an object it goes from a gas to a liquid to a solid. The changes of state that you should know are:

- Melting is the process of going from solid to liquid.

[1]This content is available online at <http://cnx.org/content/m38210/1.7/>.

- Boiling (or evaporation) is the process of going from liquid to gas.
- Freezing is the process of going from liquid to solid.
- Condensation is the process of going from gas to liquid.
- Occasionally (e.g. for carbon dioxide) we can go directly from solid to gas in a process called sublimation.

A solid has a fixed shape and volume. A liquid takes on the shape of the container that it is in. A gas completely fills the container that it is in. See Section 2.3 (The Kinetic Theory of Matter) for more on changes of state.

If we know the melting and boiling point of a substance then we can say what state (solid, liquid or gas) it will be in at any temperature.

2.2.1 Experiment: States of matter

Aim

To investigate the heating and cooling curve of water.

Apparatus

beakers, ice, bunsen burner, thermometer, water.

Method

- Place some ice in a beaker
- Measure the temperature of the ice and record it.
- After 10 s measure the temperature again and record it. Repeat every 10 s, until at least 1 minute after the ice has melted.
- Heat some water in a beaker until it boils. Measure and record the temperature of the water.
- Remove the water from the heat and measure the temperature every 10 s, until the beaker is cool to touch

WARNING: Be careful when handling the beaker of hot water. Do not touch the beaker with your hands, you will burn yourself.

Results

Record your results in the following table:

Temperature of ice	**Time (s)**	**Temperature of water**	**Time (s)**

Table 2.1: Table of results

Plot a graph of temperature against time for the ice melting and the boiling water cooling.

Discussion and conclusion

Discuss your results with others in your class. What conclusions can you draw? You should find that the temperature of the ice increases until the first drops of liquid appear and then the temperature remains the same, until all the ice is melted. You should also find that when you cool water down from boiling, the temperature remains constant for a while, then starts decreasing.

In the above experiment, you investigated the heating and cooling curves of water. We can draw heating and cooling curves for any substance. A heating curve of a substance gives the changes in temperature as we move from a solid to a liquid to a gas. A cooling curve gives the changes in temperature as we move from gas to liquid to solid. An important observation is that as a substance melts or boils, the temperature remains constant until the substance has changed state. This is because all the heat energy goes into breaking or forming the forces between the molecules.

The above experiment is one way of demonstrating the changes of state of a substance. Ice melting or water boiling should be very familiar to you.

2.3 The Kinetic Theory of Matter

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

Broadly, the Kinetic Theory of Matter says that:

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

Table 2.2 summarises the characteristics of the particles that are in each phase of matter.

Property of matter	**Solid**	**Liquid**	**Gas**
Particles	Atoms or molecules	Atoms or molecules	Atoms or molecules
Energy and movement of particles	Low energy - particles vibrate around a fixed point	Particles have less energy than in the gas phase	Particles have high energy and are constantly moving
Spaces between particles	Very little space between particles. Particles are tightly packed together	Smaller spaces than in gases, but larger spaces than in solids	Large spaces because of high energy

continued on next page

Attractive forces between particles	Very strong forces. Solids have a fixed volume.	Stronger forces than in gas. Liquids can be poured.	Weak forces because of the large distance between particles
Changes in phase	Solids become liquids if their temperature is increased. In some cases a solid may become a gas if the temperature is increased.	A liquid becomes a gas if its temperature is increased. It becomes a solid if its temperature decreases.	In general a gas becomes a liquid when it is cooled. (In a few cases a gas becomes a solid when cooled). Particles have less energy and therefore move closer together so that the attractive forces become stronger, and the gas becomes a liquid (or a solid.)

Table 2.2: Table summarising the general features of solids, liquids and gases.

The following presentation is a brief summary of the above. Try to fill in the blank spaces before clicking onto the next slide.

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<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=kinetictheory-done-100510040904-phpapp02&stripped_title=kinetic-theory-done>

Figure 2.1

Let's look at an example that involves the three phases of water: ice (solid), water (liquid) and water vapour (gas). Note that in the Figure 2.2 below the molecules in the solid phase are represented by single spheres, but they would in reality look like the molecules in the liquid and gas phase. Sometimes we represent molecules as single spheres in the solid phase to emphasise the small amount of space between them and to make the drawing simpler.

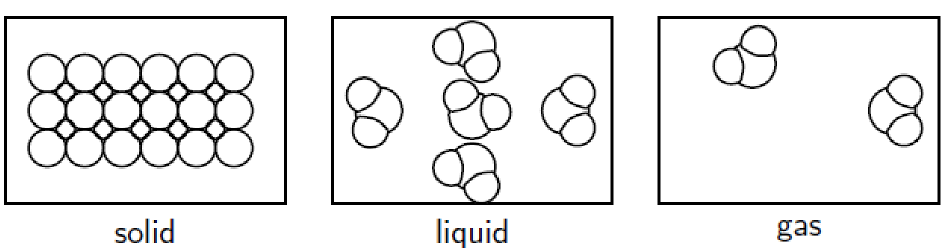

Figure 2.2: The three phases of matter

Taking water as an example we find that in the solid phase the water molecules have very little energy and can't move away from each other. The molecules are held closely together in a regular pattern called a *lattice*. If the ice is heated, the energy of the molecules increases. This means that some of the water molecules are able to overcome the intermolecular forces that are holding them together, and the molecules move further apart to form *liquid water*. This is why liquid water is able to flow, because the molecules are more free to move than they were in the solid lattice. If the molecules are heated further, the liquid water will become water vapour, which is a gas. Gas particles have lots of energy and are far away from each other. That is why it is difficult to keep a gas in a specific area! The attractive forces between the particles are very weak and they are only loosely held together. Figure 2.3 shows the changes in phase that may occur in matter, and the names that describe these processes.

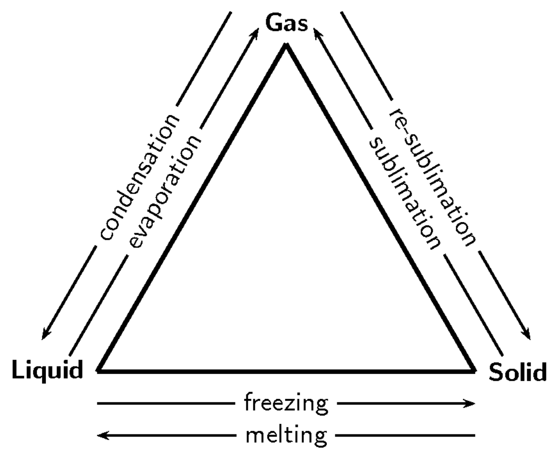

Figure 2.3: Changes in phase

2.4 Intramolecular and intermolecular forces (Not in CAPS - Included for Completeness)

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

Definition 2.1: Intramolecular force

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

Examples of the types of chemical bonds that can exist between atoms inside a molecule are shown below. These will be looked at in more detail in Chemical bonding (Chapter 5).

- *Covalent bond* Covalent bonds exist between non-metal atoms e.g. There are covalent bonds between the carbon and oxygen atoms in a molecule of carbon dioxide.
- *Ionic bond* Ionic bonds occur between non-metal and metal atoms e.g. There are ionic bonds between the sodium and chlorine atoms in a molecule of sodium chloride.

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

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

Definition 2.2: Intermolecular force

A force between molecules, which holds them together.

The following diagram may help you to understand the difference between intramolecular forces and intermolecular forces.

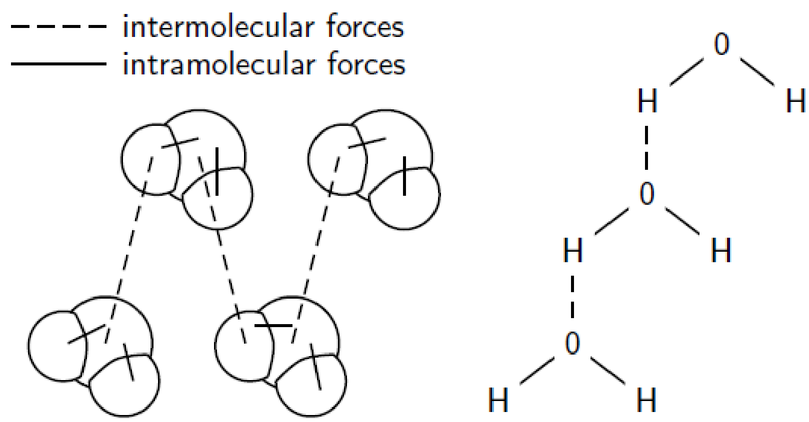

Figure 2.4: Two representations showing the intermolecular and intramolecular forces in water: space-filling model and structural formula.

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

TIP: To help you remember that intermolecular means between molecules, remember that international means between nations.

2.4.1 Intramolecular and intermolecular forces

1. Using ammonia gas as an example...
    a. Explain what is meant by an *intramolecular* force or *chemical bond*.

b. Explain what is meant by an *intermolecular* force.

Click here for the solution[2]

2. Draw a diagram showing three molecules of carbon dioxide. On the diagram, show where the intramolecular and intermolecular forces are.
Click here for the solution[3]
3. Why is it important to understand the types of forces that exist between atoms and between molecules? Try to use some practical examples in your answer.
Click here for the solution[4]

2.5 The Properties of Matter

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

1. **Melting point**

   Definition 2.3: Melting point
   The temperature at which a *solid* changes its phase or state to become a *liquid*. The process is called melting and the reverse process (change in phase from liquid to solid) is called **freezing**.

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

   Definition 2.4: Boiling point
   The temperature at which a *liquid* changes its phase to become a *gas*. The process is called evaporation and the reverse process is called condensation

   When the temperature of a liquid increases, the average kinetic energy of the particles also increases and they are able to overcome the bonding forces that are holding them in the liquid. When boiling point is reached, *evaporation* takes place and some particles in the liquid become a gas. In other words, the energy of the particles is too great for them to be held in a liquid anymore. The stronger the bonds within a liquid, the higher the boiling point needs to be in order to break these bonds. Metallic and ionic compounds have high boiling points while the boiling point for molecular liquids is lower. The data in Table 2.3 below may help you to understand some of the concepts we have explained. Not all of the substances in the table are solids at room temperature, so for now, let's just focus on the *boiling points* for each of these substances. What do you notice?

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

[2]http://www.fhsst.org/lin
[3]http://www.fhsst.org/liQ
[4]http://www.fhsst.org/liU

Table 2.3: The melting and boiling points for a number of substances

You will have seen that substances such as ethanol, with relatively weak intermolecular forces, have the lowest boiling point, while substances with stronger intermolecular forces such as sodium chloride and mercury, must be heated much more if the particles are to have enough energy to overcome the forces that are holding them together in the liquid. See the section (Section 2.5.1: Exercise: Forces and boiling point) below for a further exercise on boiling point.

3. **Density and viscosity**

    NOTE: Density and viscosity is not in CAPS - Included for Completeness

    Definition 2.5: Density
    Density is a measure of the mass of a substance per unit volume.

    The density of a solid is generally higher than that of a liquid because the particles are held much more closely together and therefore there are more particles packed together in a particular volume. In other words, there is a greater mass of the substance in a particular volume. In general, density increases as the strength of the intermolecular forces increases.

    Definition 2.6: Viscosity
    Viscosity is a measure of how resistant a liquid is to flowing (in other words, how easy it is to pour the liquid from one container to another).

    Viscosity is also sometimes described as the 'thickness' of a fluid. Think for example of syrup and how slowly it pours from one container into another. Now compare this to how easy it is to pour water. The viscosity of syrup is greater than the viscosity of water. Once again, the stronger the intermolecular forces in the liquid, the greater its viscosity.

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

2.5.1 Exercise: Forces and boiling point

The table below gives the molecular formula and the boiling point for a number of organic compounds called *alkanes* (more on these compounds in grade 12). Refer to the table and then answer the questions that follow.

Organic compound	Molecular formula	Boiling point ($^\circ C$)
Methane	CH_4	-161.6
Ethane	C_2H_6	- 88.6
Propane	C_3H_8	-45
Butane	C_4H_{10}	-0.5
Pentane	C_5H_{12}	36.1
Hexane	C_6H_{14}	69
Heptane	C_7H_{16}	98.42
Octane	C_8H_{18}	125.52

Table 2.4

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

1. Draw a graph to show the relationship between the number of carbon atoms in each alkane and its boiling point. (Number of carbon atoms will go on the x-axis and boiling point on the y-axis).
2. Describe what you see.
3. Suggest a reason for what you have observed.
4. Why was it enough for us to use 'number of carbon atoms' as a measure of the molecular weight of the molecules?

Click here for the solution[5]

2.5.2 Investigation : Determining the density of liquids:

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

The equation for density is:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \tag{2.1}$$

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

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

Apparatus:
Laboratory mass balance, 10 ml and 100 ml graduated cylinders, thread, distilled water, two different liquids.
Method:
Determine the density of the distilled water and two liquids as follows:

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

Liquid	**Mass (g)**	**Volume (ml)**	**Density ($g \cdot \text{ml}^{-1}$)**
Distilled water			
Liquid 1			
Liquid 2			

Table 2.5

2.5.3 Investigation : Determining the density of irregular solids:

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

[5] http://www.fhsst.org/liP

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

Solid	**Mass (g)**	**Volume (ml)**	**Density** $(g \cdot \mathrm{ml}^{-1})$
Solid 1			
Solid 2			
Solid 3			

Table 2.6

This media object is a Flash object. Please view or download it at <http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=phasesofmatter-100510045221-phpapp01&stripped_title=phases-of-matter&userName=kwarne>

Figure 2.5

2.6 Summary

- There are three states of matter: solid, liquid and gas.
- Diffusion is the movement of particles from a high concentration to a low concentration. Brownian motion is the diffusion of many particles.
- The **kinetic theory of matter** attempts to explain the behaviour of matter in different phases.
- The kinetic theory of matter says that all matter is composed of **particles** which have a certain amount of **energy** which allows them to **move** at different speeds depending on the temperature (energy). There are **spaces** between the particles and also **attractive forces** between particles when they come close together.
- Intramolecular force is the force between the atoms of a molecule, which holds them together. Intermolecular force is a force between molecules, which holds them together.
- Understanding chemical bonds, intermolecular forces and the kinetic theory of matter can help to explain many of the **macroscopic properties** of matter.
- **Melting point** is the temperature at which a *solid* changes its phase to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the melting point will be.
- **Boiling point** is the temperature at which a liquid changes phase to become a gas. The reverse process (change in phase from gas to liquid) is called **condensing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the boiling point will be.
- **Density** is a measure of the mass of a substance per unit volume.
- **Viscosity** is a measure of how resistant a liquid is to flowing.

2.7 End of chapter exercises

1. Give one word or term for each of the following descriptions.
   a. The property that determines how easily a liquid flows.
   b. The change in phase from liquid to gas.

   Click here for the solution[6]
2. If one substance A has a melting point that is *lower* than the melting point of substance B, this suggests that...
   a. A will be a liquid at room temperature.
   b. The chemical bonds in substance A are weaker than those in substance B.
   c. The chemical bonds in substance A are stronger than those in substance B.
   d. B will be a gas at room temperature.

   Click here for the solution[7]
3. Boiling point is an important concept to understand.
   a. Define 'boiling point'.
   b. What change in phase takes place when a liquid reaches its boiling point?
   c. What is the boiling point of water?
   d. Use the kinetic theory of matter and your knowledge of intermolecular forces to explain why water changes phase at this temperature.

   Click here for the solution[8]
4. Describe a solid in terms of the kinetic molecular theory.
   Click here for the solution[9]
5. Refer to the table below which gives the melting and boiling points of a number of elements and then answer the questions that follow. (*Data from http://www.chemicalelements.com*)

Element	**Melting point**	**Boiling point** ($^\circ C$)
copper	1083	2567
magnesium	650	1107
oxygen	-218,4	-183
carbon	3500	4827
helium	-272	-268,6
sulphur	112,8	444,6

Table 2.7

   a. What state of matter (i.e. solid, liquid or gas) will each of these elements be in at room temperature?
   b. Which of these elements has the strongest forces between its atoms? Give a reason for your answer.
   c. Which of these elements has the weakest forces between its atoms? Give a reason for your answer.

   Click here for the solution[10]

[6]http://www.fhsst.org/l2t
[7]http://www.fhsst.org/lip
[8]http://www.fhsst.org/lim
[9]http://www.fhsst.org/lgf
[10]http://www.fhsst.org/liy

Chapter 3

The Atom[1]

3.1 Introduction

The following video covers some of the properties of an atom.

Veritasium video on the atom - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/SeDaOigLBTU;rel=0>

Figure 3.1

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

3.2 Project: Models of the atom

Our current understanding of the atom came about over a long period of time, with many different people playing a role. Conduct some research into the development of the different ideas of the atom and the people who contributed to it. Some suggested people to look at are: JJ Thomson, Ernest Rutherford, Marie Curie, JC Maxwell, Max Planck, Albert Einstein, Niels Bohr, Lucretius, LV de Broglie, CJ Davisson, LH Germer, Chadwick, Werner Heisenberg, Max Born, Erwin Schrodinger, John Dalton, Empedocles, Leucippus, Democritus, Epicurus, Zosimos, Maria the Jewess, Geber, Rhazes, Robert Boyle, Henry Cavendish, A Lavoisier and H Becquerel. You do not need to find information on all these people, but try to find information about as many of them as possible.

Make a list of the key contributions to a model of the atom that each of these people made and then make a timeline of this information. (You can use an online tool such as Dipity[2] to make a timeline.) Try to get a feel for how it all eventually fit together into the modern understanding of the atom.

[1]This content is available online at <http://cnx.org/content/m38126/1.5/>.
[2]http://www.dipity.com/

3.3 Models of the Atom

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

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

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

Definition 3.1: Model
A model is a representation of a system in the real world. Models help us to understand systems and their properties. For example, an *atomic model* represents what the structure of an atom *could* look like, based on what we know about how atoms behave. It is not necessarily a true picture of the exact structure of an atom.

3.3.1 The Plum Pudding Model

After the electron was discovered by J.J. Thomson in 1897, people realised that atoms were made up of even smaller particles than they had previously thought. However, the atomic nucleus had not been discovered yet and so the 'plum pudding model' was put forward in 1904. In this model, the atom is made up of negative electrons that float in a soup of positive charge, much like plums in a pudding or raisins in a fruit cake (Figure 3.2). In 1906, Thomson was awarded the Nobel Prize for his work in this field. However, even with the Plum Pudding Model, there was still no understanding of how these electrons in the atom were arranged.

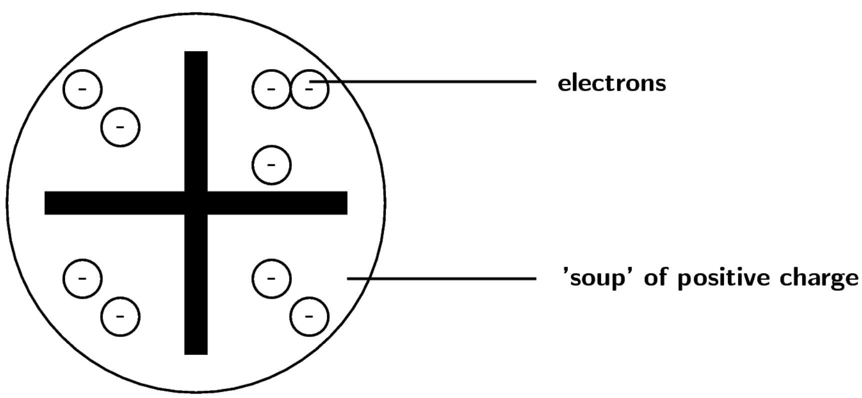

Figure 3.2: A schematic diagram to show what the atom looks like according to the Plum Pudding model

The discovery of **radiation** was the next step along the path to building an accurate picture of atomic structure. In the early twentieth century, Marie Curie and her husband Pierre, discovered that some elements (the *radioactive* elements) emit particles, which are able to pass through matter in a similar way to X-rays (read more about this in Grade 11). It was Ernest Rutherford who, in 1911, used this discovery to revise the model of the atom.

3.3.2 Rutherford's model of the atom

Radioactive elements emit different types of particles. Some of these are positively charged alpha (α) particles. Rutherford carried out a series of experiments where he bombarded sheets of gold foil with these particles, to try to get a better understanding of where the positive charge in the atom was. A simplified diagram of his experiment is shown in Figure 3.3.

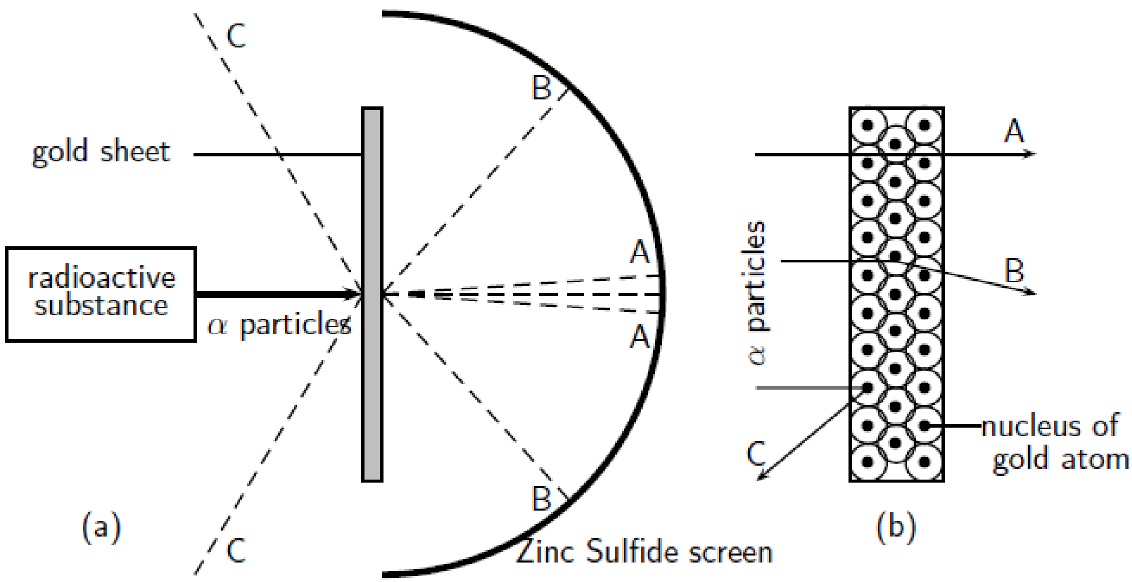

Figure 3.3: Rutherford's gold foil experiment. Figure (a) shows the path of the α particles after they hit the gold sheet. Figure (b) shows the arrangement of atoms in the gold sheets and the path of the α particles in relation to this.

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

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

Rutherford's work led to a change in ideas around the atom. His new model described the atom as a tiny, dense, positively charged core called a nucleus surrounded by lighter, negatively charged electrons. Another way of thinking about this model was that the atom was seen to be like a mini solar system where the electrons orbit the nucleus like planets orbiting around the sun. A simplified picture of this is shown in Figure 3.4. This model is sometimes known as the planetary model of the atom.

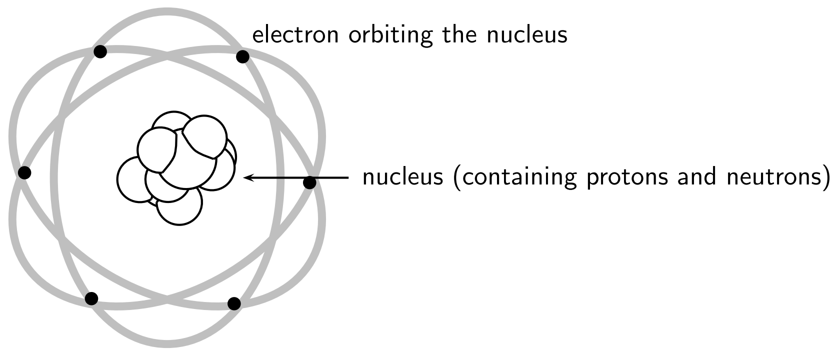

Figure 3.4: Rutherford's model of the atom

3.3.3 The Bohr Model

There were, however, some problems with this model: for example it could not explain the very interesting observation that atoms only emit light at certain wavelengths or frequencies. Niels Bohr solved this problem by proposing that the electrons could only orbit the nucleus in certain special orbits at different energy levels around the nucleus. The exact energies of the orbitals in each energy level depends on the type of atom. Helium for example, has different energy levels to Carbon. If an electron jumps down from a higher energy level to a lower energy level, then light is emitted from the atom. The energy of the light emitted is the same as the gap in the energy between the two energy levels. You can read more about this in "Energy quantisation and electron configuration" (Section 3.8: Electron configuration). The distance between the nucleus and the electron in the lowest energy level of a hydrogen atom is known as the **Bohr radius**.

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

3.3.4 Other models of the atom

Although the most common model of the atom is the Bohr model, scientists have not stopped thinking about other ways to describe atoms. One of the most important contributions to atomic theory (the field of science that looks at atoms) was the development of quantum theory. Schrodinger, Heisenberg, Born and many others have had a role in developing quantum theory. The description of an atom by quantum theory is very complex and is only covered at university level.

3.3.5 Models of the atom

Match the information in column A, with the key discoverer in column B.

Column A	Column B
Discovery of electrons and the plum pudding model	Niels Bohr
Arrangement of electrons	Marie Curie and her husband, Pierre
Atoms as the smallest building block of matter	Ancient Greeks
Discovery of the nucleus	JJ Thomson
Discovery of radiation	Rutherford

Table 3.1

Click here for the solution[3]

3.4 The size of atoms

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

3.4.1 How heavy is an atom?

It is possible to determine the mass of a single atom in kilograms. But to do this, you would need very modern *mass spectrometers* and the values you would get would be very clumsy and difficult to use. The mass of a carbon atom, for example, is about $1,99 \times 10^{-26}$ kg, while the mass of an atom of hydrogen is about $1,67 \times 10^{-27}$ kg. Looking at these very small numbers makes it difficult to compare how much bigger the mass of one atom is when compared to another.

To make the situation simpler, scientists use a different unit of mass when they are describing the mass of an atom. This unit is called the **atomic mass unit** (amu). We can abbreviate (shorten) this unit to just 'u'. Scientists use the **carbon standard** to determine amu. The carbon standard assigns carbon an atomic mass of 12 u. Using the carbon standard the mass of an atom of hydrogen will be 1 u. You can check this by dividing the mass of a carbon atom in kilograms (see above) by the mass of a hydrogen atom in kilograms (you will need to use a calculator for this!). If you do this calculation, you will see that the mass of a carbon atom is twelve times greater than the mass of a hydrogen atom. When we use atomic mass units instead of kilograms, it becomes easier to see this. Atomic mass units are therefore not giving us the *actual* mass of an atom, but rather its mass *relative* to the mass of one (carefully chosen) atom in the Periodic Table. Although carbon is the usual element to compare other elements to, oxygen and hydrogen have also been used. The important thing to remember here is that the atomic mass unit is relative to one (carefully chosen) element. The atomic masses of some elements are shown in the table (Table 3.2) below.

[3]http://www.fhsst.org/l4g

Element	Atomic mass (u)
Carbon (C)	12
Nitrogen (N)	14
Bromine (Br)	80
Magnesium (Mg)	24
Potassium (K)	39
Calcium (Ca)	40
Oxygen (O)	16

Table 3.2: The atomic mass number of some of the elements

The actual value of 1 atomic mass unit is $1,67 \times 10^{-24}\,g$ or $1,67 \times 10^{-27}$ kg. This is a very tiny mass!

3.4.2 How big is an atom?

TIP: *pm* stands for *picometres*. $1\,\text{pm} = 10^{-12}\,m$

Atomic radius also varies depending on the element. On average, the radius of an atom ranges from 32 pm (Helium) to 225 pm (Caesium). Using different units, 100 pm = 1 Angstrom, and 1 Angstrom $= 10^{-10}\,m$. That is the same as saying that 1 Angstrom $= 0,0000000010\,m$ or that 100 pm $= 0,0000000010\,m$! In other words, the diameter of an atom ranges from $0,0000000010\,m$ to $0,0000000067\,m$. This is very small indeed.

The atomic radii given above are for the whole atom (nucleus and electrons). The nucleus itself is even smaller than this by a factor of about 23 000 in uranium and 145 000 in hydrogen. If the nucleus were the size of a golf ball, then the nearest electrons would be about one kilometer away! This should give help you realise that the atom is mostly made up of empty space.

3.5 Atomic structure

As a result of the work done by previous scientists on atomic models (that we discussed in "Models of the Atom" (Section 3.3: Models of the Atom)), scientists now have a good idea of what an atom looks like. This knowledge is important because it helps us to understand why materials have different properties and why some materials bond with others. Let us now take a closer look at the microscopic structure of the atom.

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

3.5.1 The Electron

The electron is a very light particle. It has a mass of $9,11 \times 10^{-31}$ kg. Scientists believe that the electron can be treated as a **point particle** or **elementary particle** meaning that it can't be broken down into anything smaller. The electron also carries one unit of **negative** electric charge which is the same as $1,6 \times 10^{-19}\,C$ (Coulombs).

The electrons determine the charge on an atom. If the number of electrons is the same as the number of protons then the atom will be neutral. If the number of electrons is greater than the number of protons then the atom will be negatively charged. If the number of electrons is less than the number of protons then the atom will be positively charged. Atoms that are not neutral are called ions. Ions will be covered in more detail in a later chapter. For now all you need to know is that for each electron you remove from an atom you loose -1 of charge and for each electron that you add to an atom you gain $+1$ of charge. For example, the charge on an atom of sodium after removing one electron is -1.

3.5.2 The Nucleus

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

3.5.2.1 The Proton

Each proton carries one unit of **positive** electric charge. Since we know that atoms are **electrically neutral**, i.e. do not carry any extra charge, then the number of protons in an atom has to be the same as the number of electrons to balance out the positive and negative charge to zero. The total positive charge of a nucleus is equal to the number of protons in the nucleus. The proton is much heavier than the electron (10 000 times heavier!) and has a mass of $1,6726 \times 10^{-27}$ kg. When we talk about the atomic mass of an atom, we are mostly referring to the combined mass of the protons and neutrons, i.e. the nucleons.

3.5.2.2 The Neutron

The neutron is electrically neutral i.e. it carries no charge at all. Like the proton, it is much heavier than the electron and its mass is $1,6749 \times 10^{-27}$ kg (slightly heavier than the proton).

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

	proton	**neutron**	**electron**
Mass (kg)	$1,6726 \times 10^{-27}$	$1,6749 \times 10^{-27}$	$9,11 \times 10^{-31}$
Units of charge	$+1$	0	-1
Charge (C)	$1,6 \times 10^{-19}$	0	$-1,6 \times 10^{-19}$

Table 3.3: Summary of the particles inside the atom

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

3.6 Atomic number and atomic mass number

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

Definition 3.2: Atomic number (Z)
The number of protons in an atom

You can find the atomic number on the periodic table. The atomic number is an integer and ranges from 1 to about 118.

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

Definition 3.3: Atomic mass number (A)
The number of protons and neutrons in the nucleus of an atom

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

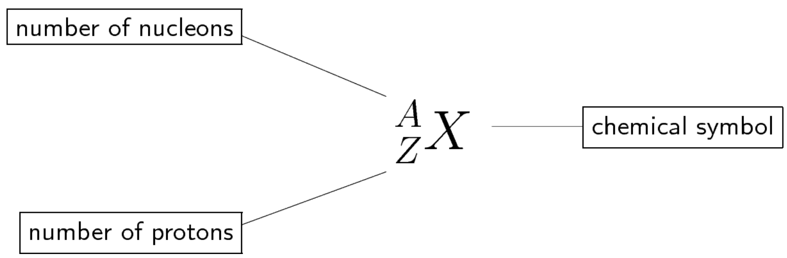

Figure 3.5

NOTE: A nuclide is a distinct kind of atom or nucleus characterized by the number of protons and neutrons in the atom. To be absolutely correct, when we represent atoms like we do here, then we should call them nuclides.

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

$$^{56}_{26}Fe \tag{3.1}$$

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

TIP: Don't confuse the notation we have used above with the way this information appears on the Periodic Table. On the Periodic Table, the atomic number usually appears in the top lefthand corner of the block or immediately above the element's symbol. The number below the element's symbol is its **relative atomic mass**. This is not exactly the same as the atomic mass number. This will be explained in "Isotopes" (Section 3.7: Isotopes). The example of iron is shown below.

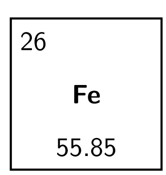

Figure 3.6

You will notice in the example of iron that the atomic mass number is more or less the same as its atomic mass. Generally, an atom that contains n nucleons (protons and neutrons), will have a mass approximately equal to nu. For example the mass of a $C-12$ atom which has 6 protons, 6 neutrons and 6 electrons is 12u, where the protons and neutrons have about the same mass and the electron mass is negligible.

Exercise 3.1 *(Solution on p. 66.)*

Use standard notation to represent sodium and give the number of protons, neutrons and electrons in the element.

3.6.1 The structure of the atom

1. Explain the meaning of each of the following terms:
   a. nucleus
   b. electron
   c. atomic mass

   Click here for the solution[4]
2. Complete the following table: (Note: You will see that the atomic masses on the Periodic Table are not *whole numbers*. This will be explained later. For now, you can round off to the nearest whole number.)

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

Table 3.4

   Click here for the solution[5]
3. Use standard notation to represent the following elements:
   a. potassium
   b. copper
   c. chlorine

   Click here for the solution[6]
4. For the element $^{35}_{17}Cl$, give the number of ...
   a. protons
   b. neutrons
   c. electrons

   ... in the atom.
   Click here for the solution[7]

[4]http://cnx.org/content/m38126/latest/ http://www.fhsst.org/ll0
[5]http://cnx.org/content/m38126/latest/ http://www.fhsst.org/ll8
[6]http://cnx.org/content/m38126/latest/ http://www.fhsst.org/ll9
[7]http://cnx.org/content/m38126/latest/ http://www.fhsst.org/llX

5. Which of the following atoms has 7 electrons?
   a. $^{5}_{2}He$
   b. $^{13}_{6}C$
   c. $^{7}_{3}Li$
   d. $^{15}_{7}N$

   Click here for the solution[8]
6. In each of the following cases, give the number or the element symbol represented by 'X'.
   a. $^{40}_{18}X$
   b. $^{x}_{20}Ca$
   c. $^{31}_{x}P$

   Click here for the solution[9]
7. Complete the following table:

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

Table 3.5

In these two different forms of Uranium...

   a. What is the *same*?
   b. What is *different*?

   Uranium can occur in different forms, called *isotopes*. You will learn more about isotopes in "Isotopes" (Section 3.7: Isotopes).
   Click here for the solution[10]

3.7 Isotopes

3.7.1 What is an isotope?

The chemical properties of an element depend on the number of protons and electrons inside the atom. So if a neutron or two is added or removed from the nucleus, then the chemical properties will not change. This means that such an atom would remain in the same place in the Periodic Table. For example, no matter how many neutrons we add or subtract from a nucleus with 6 protons, that element will **always** be called carbon and have the element symbol C (see the Table of Elements). Atoms which have the same number of protons, but a different number of neutrons, are called **isotopes**.

Definition 3.4: Isotope
The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons.

The different isotopes of an element have the same atomic number Z but different mass numbers A because they have a different number of neutrons N. The chemical properties of the different isotopes of an element are the same, but they might vary in how stable their nucleus is. Note that we can also write elements as X $-$ A where the X is the element symbol and the A is the atomic mass of that element. For example, C$-$12 has an atomic mass of 12 and Cl$-$35 has an atomic mass of 35 u, while Cl$-$37 has an atomic mass of 37 u.

[8] http://cnx.org/content/m38126/latest/ http://www.fhsst.org/llk
[9] http://cnx.org/content/m38126/latest/ http://www.fhsst.org/llK
[10] http://cnx.org/content/m38126/latest/ http://www.fhsst.org/llB

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

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

Exercise 3.2: Isotopes *(Solution on p. 66.)*

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

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

Exercise 3.3: Isotopes *(Solution on p. 66.)*

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

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

Exercise 3.4: Isotopes *(Solution on p. 66.)*

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

a. protons
b. nucleons
c. electrons
d. neutrons

3.7.1.1 Isotopes

1. Atom A has 5 protons and 5 neutrons, and atom B has 6 protons and 5 neutrons. These atoms are...
   a. allotropes
   b. isotopes
   c. isomers
   d. atoms of different elements

   Click here for the solution[11]
2. For the sulphur isotopes, $^{32}_{16}S$ and $^{34}_{16}S$, give the number of...
   a. protons
   b. nucleons
   c. electrons
   d. neutrons

   Click here for the solution[12]
3. Which of the following are isotopes of $^{35}_{17}Cl$?
   a. $^{17}_{35}Cl$
   b. $^{35}_{17}Cl$
   c. $^{37}_{17}Cl$

   Click here for the solution[13]

[11] http://www.fhsst.org/ll4
[12] http://www.fhsst.org/llZ
[13] http://www.fhsst.org/llW

4. Which of the following are isotopes of U−235? (X represents an element symbol)
   a. $^{238}_{92}X$
   b. $^{238}_{90}X$
   c. $^{235}_{92}X$

   Click here for the solution[14]

3.7.2 Relative atomic mass

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

Definition 3.5: Relative atomic mass
Relative atomic mass is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units.

NOTE: The relative atomic mass of some elements depends on where on Earth the element is found. This is because the isotopes can be found in varying ratios depending on certain factors such as geological composition, etc. The International Union of Pure and Applied Chemistry (IUPAC) has decided to give the relative atomic mass of some elements as a range to better represent the varying isotope ratios on the Earth. For the calculations that you will do at high school, it is enough to simply use one number without worrying about these ranges.

Exercise 3.5: The relative atomic mass of an isotopic element ***(Solution on p. 66.)***
The element chlorine has two isotopes, chlorine-35 and chlorine-37. The abundance of these isotopes when they occur naturally is 75% chlorine-35 and 25% chlorine-37. Calculate the *average* relative atomic mass for chlorine.

This simulation allows you to see how isotopes and relative atomic mass are inter related.

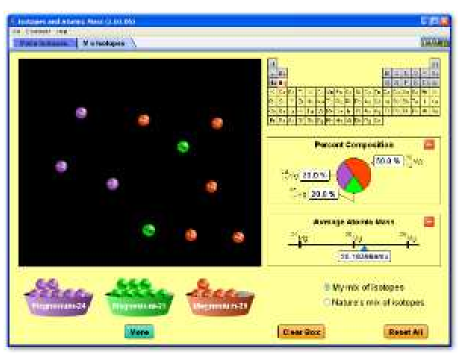

Figure 3.7

[14]http://www.fhsst.org/llD

run demo[15]

3.7.2.1 Isotopes

1. Complete the table below:

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

Table 3.6

   Click here for the solution[16]
2. If a sample contains 90% carbon-12 and 10% carbon-14, calculate the relative atomic mass of an atom in that sample. Click here for the solution[17]
3. If a sample contains 22,5% Cl-37 and 77,5% Cl-35, calculate the relative atomic mass of an atom in that sample. Click here for the solution[18]

3.7.2.2 Group Discussion : The changing nature of scientific knowledge

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

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

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

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

[15] http://phet.colorado.edu/sims/build-an-atom/isotopes-and-atomic-mass_en.jnlp
[16] http://www.fhsst.org/llj
[17] http://www.fhsst.org/llb
[18] http://www.fhsst.org/llT

3.8 Electron configuration

3.8.1 The energy of electrons

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

3.8.2 Energy quantisation and line emission spectra (Not in CAPS, included for completeness)

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

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

3.8.3 Electron configuration

We will start with a very simple view of the arrangement or configuration of electrons around an atom. This view simply states that electrons are arranged in energy levels (or shells) around the nucleus of an atom. These energy levels are numbered 1, 2, 3, etc. Electrons that are in the first energy level (energy level 1) are closest to the nucleus and will have the lowest energy. Electrons further away from the nucleus will have a higher energy.

In the following examples, the energy levels are shown as concentric circles around the central nucleus. The important thing to know for these diagrams is that the first energy level can hold 2 electrons, the second energy level can hold 8 electrons and the third energy level can hold 8 electrons.

1. **Lithium** Lithium (Li) has an atomic number of 3, meaning that in a neutral atom, the number of electrons will also be 3. The first two electrons are found in the first energy level, while the third electron is found in the second energy level (Figure 3.8).

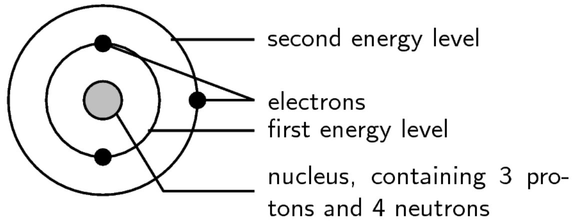

Figure 1: The arrangement of electrons in a lithium atom.

Figure 3.8: The arrangement of electrons in a lithium atom.

2. **Fluorine** Fluorine (F) has an atomic number of 9, meaning that a neutral atom also has 9 electrons. The first 2 electrons are found in the first energy level, while the other 7 are found in the second energy level (Figure 3.9).

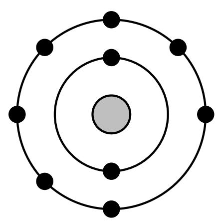

Figure 3.9: The arrangement of electrons in a fluorine atom.

3. **Argon** Argon has an atomic number of 18, meaning that a neutral atom also has 18 electrons. The first 2 electrons are found in the first energy level, the next 8 are found in the second energy level, and the last 8 are found in the third energy level (Figure 3.10).

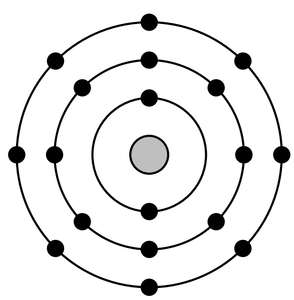

Figure 3.10: The arrangement of electrons in an argon atom.

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

Definition 3.6: Atomic orbital
An atomic orbital is the region in which an electron may be found around a single atom.

There are different orbital shapes, but we will be mainly dealing with only two. These are the 's' and 'p' orbitals (there are also 'd' and 'f' orbitals). The 's' orbitals are spherical and the 'p' orbitals are dumbbell shaped.

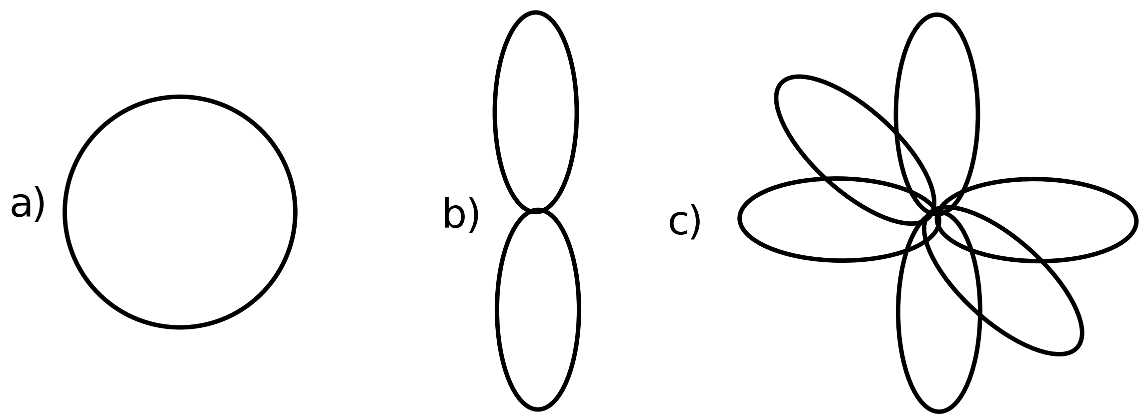

Figure 3.11: The shapes of orbitals. a) shows an 's' orbital, b) shows a single 'p' orbital and c) shows the three 'p' orbitals.

The first energy level contains only one 's' orbital, the second energy level contains one 's' orbital and three 'p' orbitals and the third energy level contains one 's' orbital and three 'p' orbitals (as well as 5 'd' orbitals). Within each energy level, the 's' orbital is at a lower energy than the 'p' orbitals. This arrangement is shown in Figure 3.12.

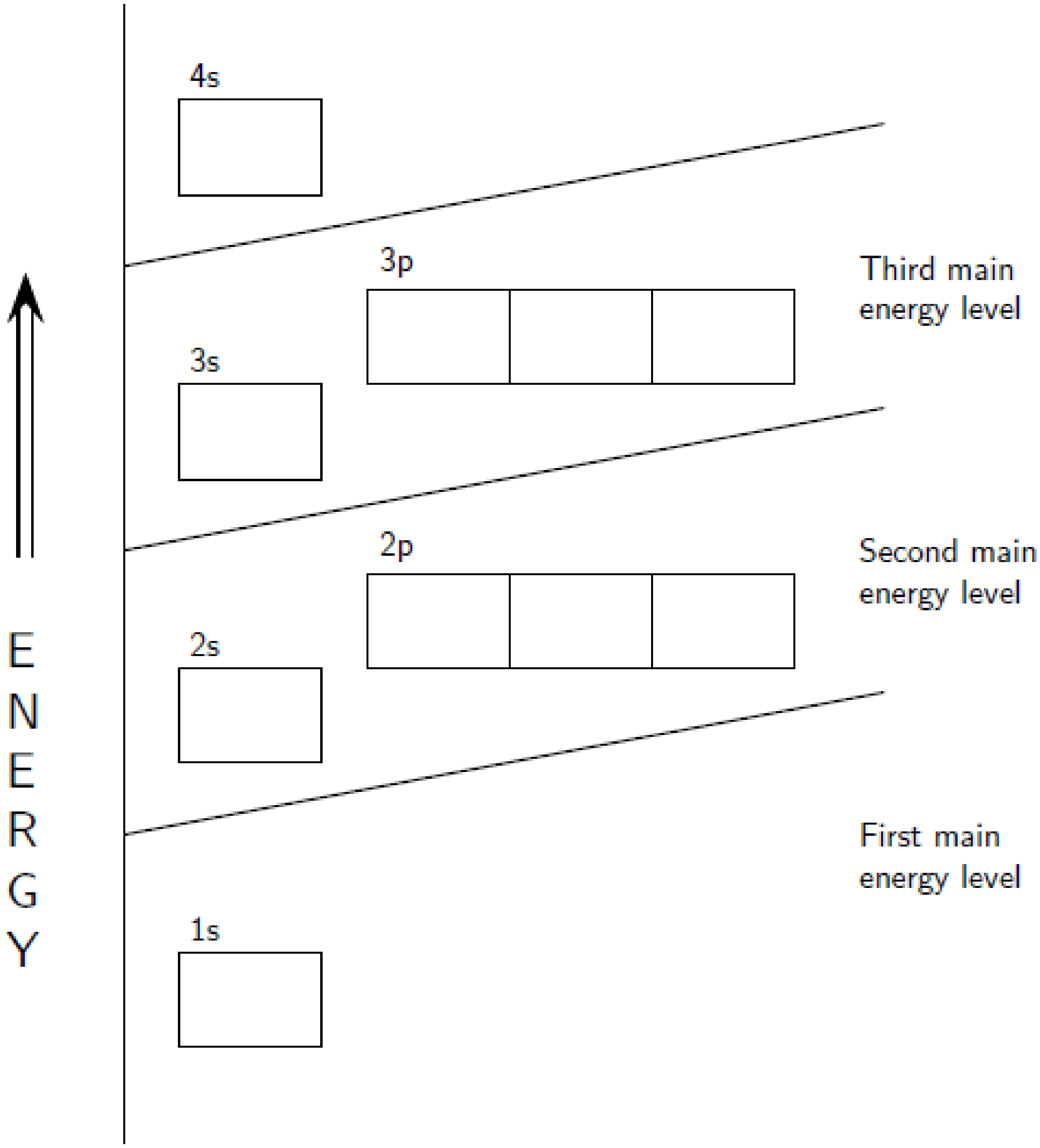

Figure 3.12: The positions of the first ten orbitals of an atom on an energy diagram. Note that each block is able to hold two electrons.

This diagram also helps us when we are working out the electron configuration of an element. The electron configuration of an element is the arrangement of the electrons in the shells and subshells. There are a few guidelines for working out the electron configuration. These are:

- Each orbital can only hold **two electrons**. Electrons that occur together in an orbital are called an **electron pair**.

- An electron will always try to enter an orbital with the lowest possible energy.
- An electron will occupy an orbital on its own, rather than share an orbital with another electron. An electron would also rather occupy a lower energy orbital *with* another electron, before occupying a higher energy orbital. In other words, within one energy level, electrons will fill an 's' orbital before starting to fill 'p' orbitals.
- The s subshell can hold 2 electrons
- The p subshell can hold 6 electrons

In the examples you will cover, you will mainly be filling the s and p subshells. Occasionally you may get an example that has the d subshell. The f subshell is more complex and is not covered at this level.

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

Definition 3.7: Electron configuration

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

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

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

TIP: When there are two electrons in an orbital, the electrons are called an **electron pair**. If the orbital only has one electron, this electron is said to be an **unpaired electron**. Electron pairs are shown with arrows pointing in opposite directions. You may hear people talking of the Pauli exclusion principle. This principle says that electrons have a property known as spin and two electrons in an orbital will not spin the same way. This is why we use arrows pointing in opposite directions. An arrow pointing up denotes an electron spinning one way and an arrow pointing downwards denotes an electron spinning the other way.

NOTE: Aufbau is the German word for 'building up'. Scientists used this term since this is exactly what we are doing when we work out electron configuration, we are building up the atoms structure.

Sometimes people refer to Hund's rule for electron configuration. This rule simply says that electrons would rather be in a subshell on it's own then share a subshell. This is why, when you are filling the subshells you put one electron in each subshell and only if there are extra electrons do you go back and fill the subshell, before moving onto the next energy level.

An Aufbau diagram for the element Lithium is shown in Figure 3.13.

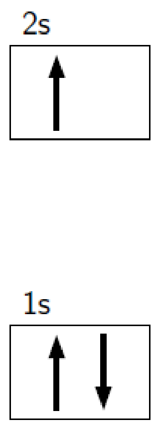

Figure 3.13: The electron configuration of Lithium, shown on an Aufbau diagram

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

Aufbau diagrams for the elements fluorine and argon are shown in Figure 3.14 and Figure 3.15 respectively. Using standard notation, the electron configuration of fluorine is $1s^2 2s^2 2p^5$ and the electron configuration of argon is $1s^2 2s^2 2p^6$.

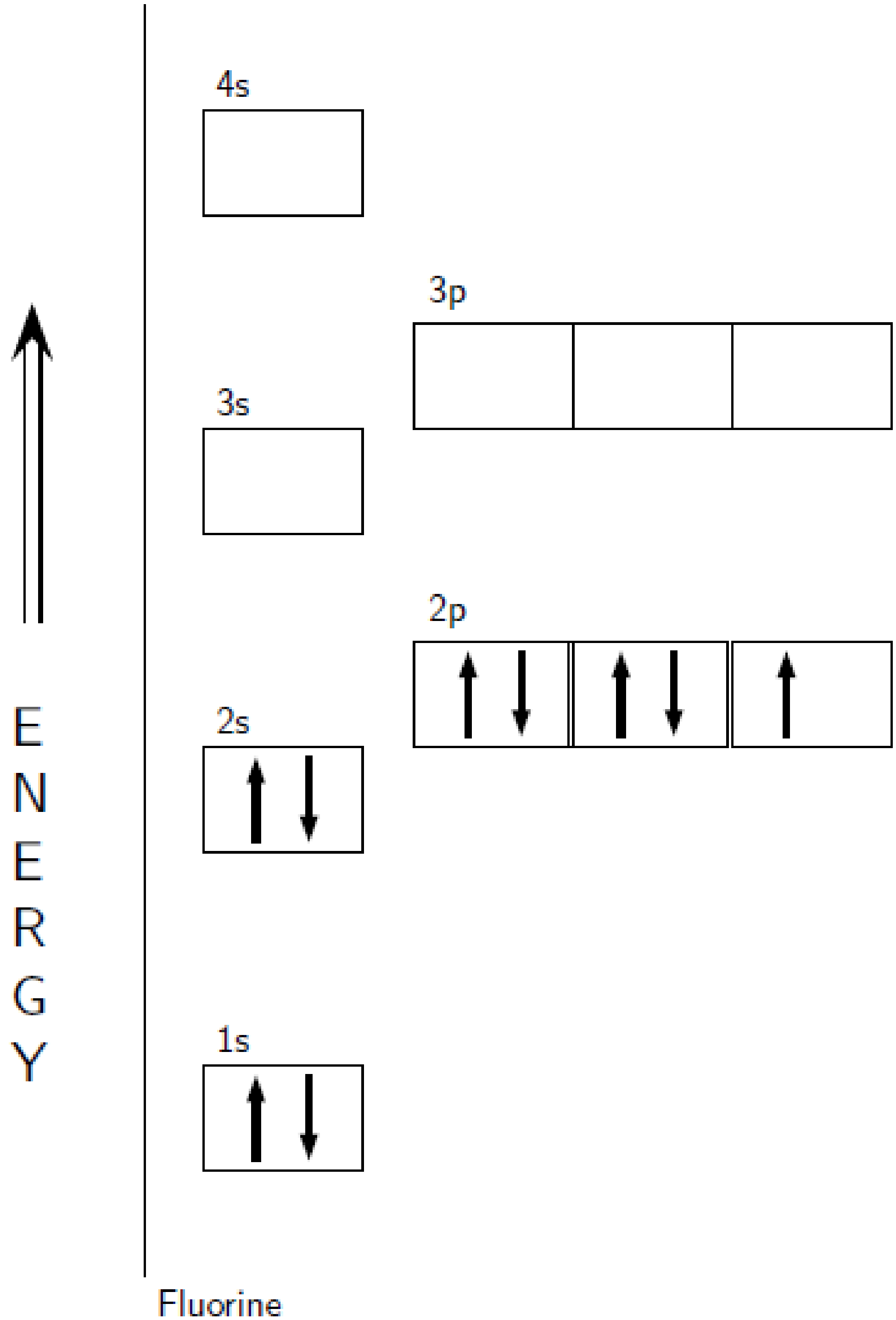

Figure 3.14: An Aufbau diagram showing the electron configuration of fluorine

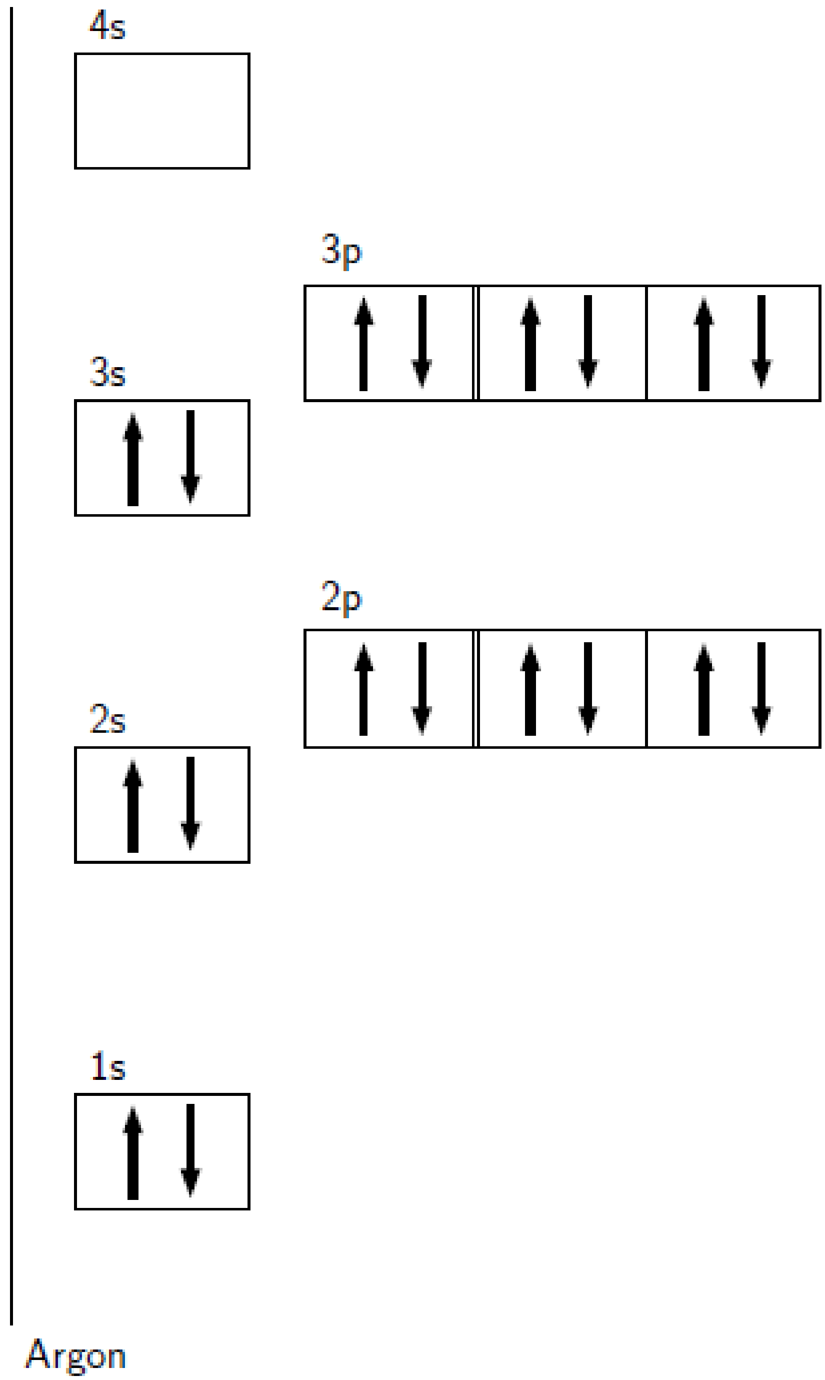

Figure 3.15: An Aufbau diagram showing the electron configuration of argon

Exercise 3.6: Aufbau diagrams *(Solution on p. 67.)*

Give the electron configuration for sodium (Na) and draw an aufbau diagram.

3.8.4 Core and valence electrons

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

Definition 3.8: Valence electrons

The electrons in the outer energy level of an atom

Definition 3.9: Core electrons

All the electrons in an atom, excluding the valence electrons

3.8.5 The importance of understanding electron configuration

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

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

One final point to note about electron configurations is stability. Which configurations are stable and which are not? Very simply, the most stable configurations are the ones that have full energy levels. These configurations occur in the noble gases. The noble gases are very stable elements that do not react easily (if at all) with any other elements. This is due to the full energy levels. All elements would like to reach the most stable electron configurations, i.e. all elements want to be noble gases. This principle of stability is sometimes referred to as the octet rule. An octet is a set of 8, and the number of electrons in a full energy level is 8.

3.8.5.1 Experiment: Flame tests

Aim:

To determine what colour a metal cation will cause a flame to be.

Apparatus:

Watch glass, bunsen burner, methanol, bamboo sticks, metal salts (e.g. $NaCl$, $CuCl_2$, $CaCl_2$, KCl, etc.) and metal powders (e.g. copper, magnesium, zinc, iron, etc.)

Method:

For each salt or powder do the following:

1. Dip a clean bamboo stick into the methanol
2. Dip the stick into the salt or powder
3. Wave the stick through the flame from the bunsen burner. DO NOT hold the stick in the flame, but rather wave it back and forth through the flame.

4. Observe what happens

Results:

Record your results in a table, listing the metal salt and the colour of the flame.
Conclusion:

You should have observed different colours for each of the metal salts and powders that you tested.

The above experiment on flame tests relates to the line emission spectra of the metals. These line emission spectra are a direct result of the arrangement of the electrons in metals.

3.8.5.2 Energy diagrams and electrons

1. Draw Aufbau diagrams to show the electron configuration of each of the following elements:
   a. magnesium
   b. potassium
   c. sulphur
   d. neon
   e. nitrogen
2. Use the Aufbau diagrams you drew to help you complete the following table:

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

Table 3.7

3. Rank the elements used above in order of *increasing reactivity*. Give reasons for the order you give. Click here for the answer[19]

3.8.5.3 Group work : Building a model of an atom

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

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

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

[19]http://www.fhsst.org/ll2

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

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

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

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

The following simulation allows you to build an atom
run demo[20]

[20] http://phet.colorado.edu/sims/build-an-atom/build-an-atom_en.jnlp

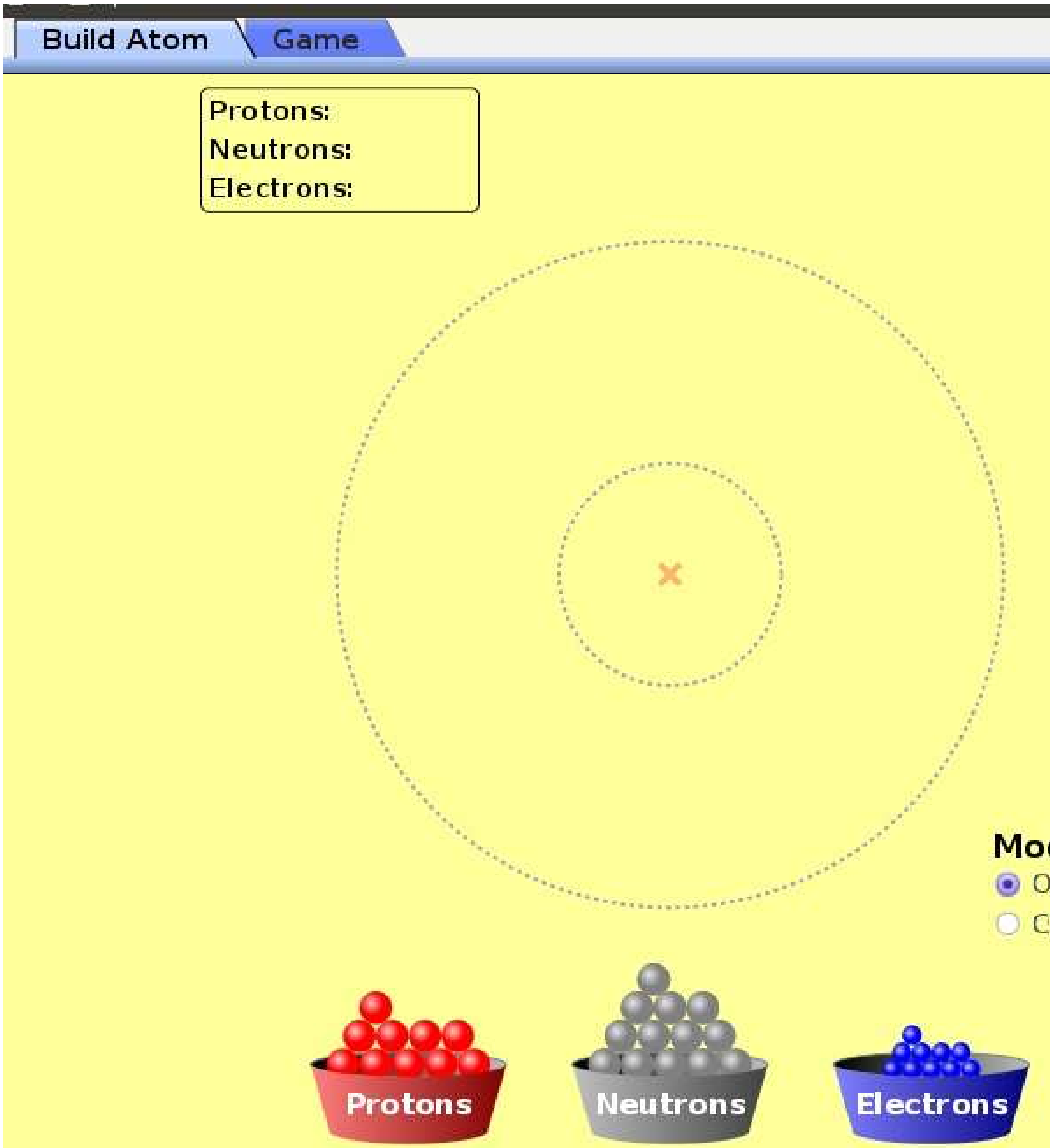

Figure 3.16: Build an atom simulation

This is another simulation that allows you to build an atom. This simulation also provides a summary of what you have learnt so far.
Run demo[21]

[21] http://scratch.mit.edu/projects/beckerr/867623

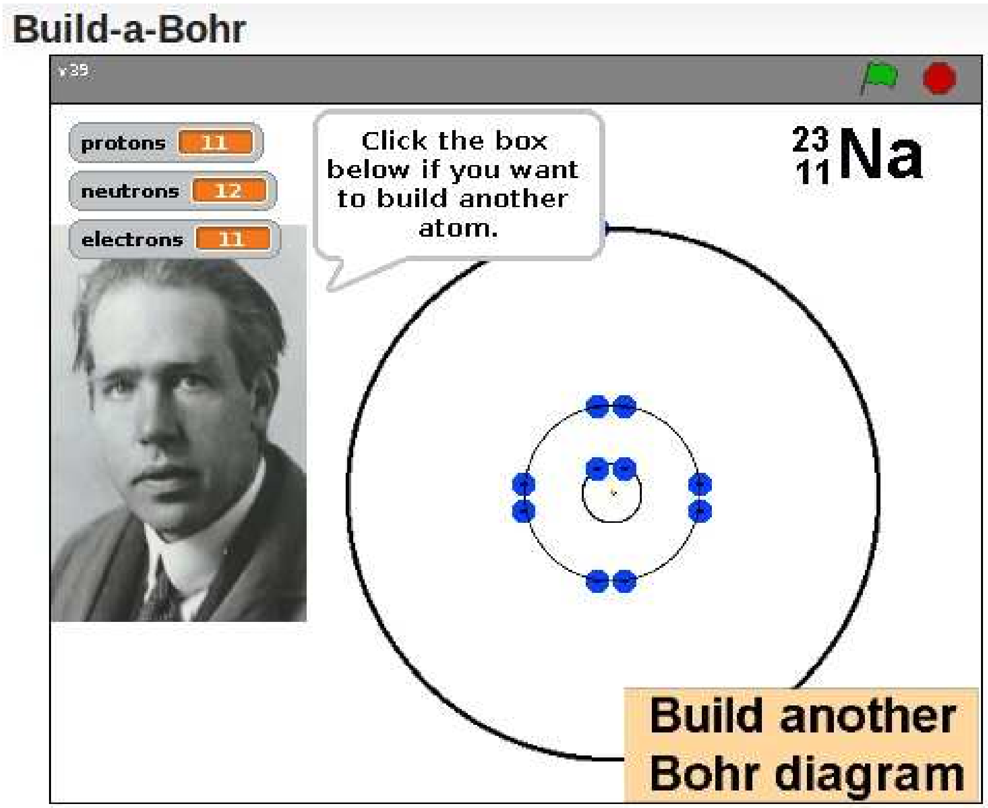

Figure 3.17: Build an atom simulation 2

3.9 Summary

- Much of what we know today about the atom, has been the result of the work of a number of scientists who have added to each other's work to give us a good understanding of atomic structure.
- Some of the important scientific contributors include **J.J.Thomson** (discovery of the electron, which led to the Plum Pudding Model of the atom), **Ernest Rutherford** (discovery that positive charge is concentrated in the centre of the atom) and **Niels Bohr** (the arrangement of electrons around the nucleus in energy levels).
- Because of the very small mass of atoms, their mass is measured in **atomic mass units** (u). $1\,u = 1,67 \times 10^{-24}\,g$.
- An atom is made up of a central **nucleus** (containing **protons** and **neutrons**), surrounded by **electrons**.
- The **atomic number** (Z) is the number of protons in an atom.
- The **atomic mass number** (A) is the number of protons and neutrons in the nucleus of an atom.
- The **standard notation** that is used to write an element, is $^A_Z X$, where X is the element symbol, A is the atomic mass number and Z is the atomic number.
- The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons. This means that not all atoms of an element will have the same atomic mass.
- The **relative atomic mass** of an element is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units. The relative atomic mass is written under the elements' symbol on the Periodic Table.
- The energy of electrons in an atom is **quantised**. Electrons occur in specific energy levels around an atom's nucleus.
- Within each energy level, an electron may move within a particular shape of **orbital**. An orbital defines the space in which an electron is most likely to be found. There are different orbital shapes, including s, p, d and f orbitals.
- Energy diagrams such as **Aufbau diagrams** are used to show the electron configuration of atoms.
- The electrons in the outermost energy level are called **valence electrons**.
- The electrons that are not valence electrons are called **core electrons**.
- Atoms whose outermost energy level is full, are less chemically reactive and therefore more stable, than those atoms whose outer energy level is not full.

This media object is a Flash object. Please view or download it at
<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=atomictheoryrevision-100512033251-phpapp01&stripped_title=atomic-theory-revision>

Figure 3.18

3.9.1 End of chapter exercises

1. Write down only the word/term for each of the following descriptions.
   a. The sum of the number of protons and neutrons in an atom
   b. The defined space around an atom's nucleus, where an electron is most likely to be found

   Click here for the solution[22]

[22] http://www.fhsst.org/lif

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

   Click here for the solution[23]
3. Multiple choice questions: In each of the following, choose the **one** correct answer.
   a. The three basic components of an atom are:
      a. protons, neutrons, and ions
      b. protons, neutrons, and electrons
      c. protons, neutrinos, and ions
      d. protium, deuterium, and tritium

      Click here for the solution[24]
   b. The charge of an atom is...
      a. positive
      b. neutral
      c. negative

      Click here for the solution[25]
   c. If Rutherford had used neutrons instead of alpha particles in his scattering experiment, the neutrons would...
      a. not deflect because they have no charge
      b. have deflected more often
      c. have been attracted to the nucleus easily
      d. have given the same results

      Click here for the solution[26]
   d. Consider the isotope $^{234}_{92}U$. Which of the following statements is *true*?
      a. The element is an isotope of $^{234}_{94}Pu$
      b. The element contains 234 neutrons
      c. The element has the same electron configuration as $^{238}_{92}U$
      d. The element has an atomic mass number of 92

      Click here for the solution[27]
   e. The electron configuration of an atom of chlorine can be represented using the following notation:
      a. $1s^2 2s^8 3s^7$
      b. $1s^2 2s^2 2p^6 3s^2 3p^5$
      c. $1s^2 2s^2 2p^6 3s^2 3p^6$
      d. $1s^2 2s^2 2p^5$

      Click here for the solution[28]
4. Give the standard notation for the following elements:
   a. beryllium
   b. carbon-12
   c. titanium-48

[23]http://www.fhsst.org/liG
[24]http://www.fhsst.org/li7
[25]http://www.fhsst.org/liA
[26]http://www.fhsst.org/lio
[27]http://www.fhsst.org/lis
[28]http://www.fhsst.org/liH

d. fluorine

Click here for the solution[29]

5. Give the electron configurations and aufbau diagrams for the following elements:
   a. aluminium
   b. phosphorus
   c. carbon

   Click here for the solution[30]
6. Use standard notation to represent the following elements:
   a. argon
   b. calcium
   c. silver-107
   d. bromine-79

   Click here for the solution[31]
7. For each of the following elements give the number of protons, neutrons and electrons in the element:
   a. $^{195}_{78}Pt$
   b. $^{40}_{18}Ar$
   c. $^{59}_{27}Co$
   d. $^{7}_{3}Li$
   e. $^{11}_{5}B$

   Click here for the solution[32]
8. For each of the following elements give the element or number represented by 'x':
   a. $^{103}_{45}X$
   b. $^{35}_{x}Cl$
   c. $^{x}_{4}Be$

   Click here for the solution[33]
9. Which of the following are isotopes of $^{24}_{12}Mg$:
   a. $^{12}_{25}Mg$
   b. $^{26}_{12}Mg$
   c. $^{24}_{13}Al$

   Click here for the solution[34]
10. If a sample contains 69% of copper-63 and 31% of copper-65, calculate the relative atomic mass of an atom in that sample.
    Click here for the solution[35]
11. Complete the following table:

[29] http://www.fhsst.org/lg7
[30] http://www.fhsst.org/lgG
[31] http://cnx.org/content/m38126/latest/ http://www.fhsst.org/l44
[32] http://cnx.org/content/m38126/latest/ http://www.fhsst.org/l42
[33] http://www.fhsst.org/l4T
[34] http://www.fhsst.org/l4b
[35] http://www.fhsst.org/l4j

Element	Electron configuration	Core electrons	Valence electrons
Boron (B)			
Calcium (Ca)			
Silicon (Si)			
Lithium (Li)			
Neon (Ne)			

Table 3.8

Click here for the solution[36]

12. Draw aufbau diagrams for the following elements:
    a. beryllium
    b. sulphur
    c. argon

    Click here for the solution[37]

[36]http://www.fhsst.org/l4D
[37]http://www.fhsst.org/lgW

Solutions to Exercises in Chapter 3

Solution to Exercise 3.1 (p. 43)

Step 1. Sodium is given by Na
Step 2. Sodium has 11 protons, so we have: $_{11}Na$
Step 3. Sodium has 12 neutrons.
Step 4. $A = N + Z = 12 + 11 = 23$
Step 5. In standard notation sodium is given by: $^{23}_{11}Na$. The number of protons is 11, the number of neutrons is 12 and the number of electrons is 11.

Solution to Exercise 3.2 (p. 45)

Step 1. We know that isotopes of any element have the **same** number of protons (same atomic number) in each atom, which means that they have the same chemical symbol. However, they have a different number of neutrons, and therefore a different mass number.
Step 2. Therefore, any isotope of uranium will have the symbol:

$$U \tag{3.2}$$

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

$$_{92}U \tag{3.3}$$

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

$$^{232}_{92}U \tag{3.4}$$

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

$$^{238}_{92}U \tag{3.5}$$

Solution to Exercise 3.3 (p. 45)

Step 1. We know that isotopes have the same atomic number but different mass numbers.
Step 2. You need to look for the element that has the same atomic number but a different atomic mass number. The only element is $^{42}_{20}Ca$. What is different is that there are 2 more neutrons than in the original element.

Solution to Exercise 3.4 (p. 45)

Step 1. $Z = 16$, therefore the number of protons is 16 (answer to (a)).
Step 2. $A = 33$, therefore the number of nucleons is 33 (answer to (b)).
Step 3. The atom is neutral, and therefore the number of electrons is the same as the number of protons. The number of electrons is 16 (answer to (c)).
Step 4.

$$N = A - Z = 33 - 16 = 17 \tag{3.6}$$

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

Solution to Exercise 3.5 (p. 46)

Step 1. Contribution of Cl−35 $= \left(\frac{75}{100} \times 35\right) = 26,25\,u$
Step 2. Contribution of Cl−37 $= \left(\frac{25}{100} \times 37\right) = 9,25\,u$

Step 3. Relative atomic mass of chlorine $= 26,25\,u + 9,25\,u = 35,5\,u$
If you look on the periodic table, the average relative atomic mass for chlorine is $35,5\,u$. You will notice that for many elements, the relative atomic mass that is shown is not a whole number. You should now understand that this number is the *average* relative atomic mass for those elements that have naturally occurring isotopes.

Solution to Exercise 3.6 (p. 55)

Step 1. Sodium has 11 electrons.
Step 2. We start by placing two electrons in the $1s$ orbital: $1s^2$. Now we have 9 electrons left to place in orbitals, so we put two in the $2s$ orbital: $2s^2$. There are now 7 electrons to place in orbitals so we place 6 of them in the $2p$ orbital: $2p^6$. The last electron goes into the $3s$ orbital: $3s^1$.
Step 3. The electron configuration is: $1s^2 2s^2 2p^6 3s^1$
Step 4. Using the electron configuration we get the following diagram:

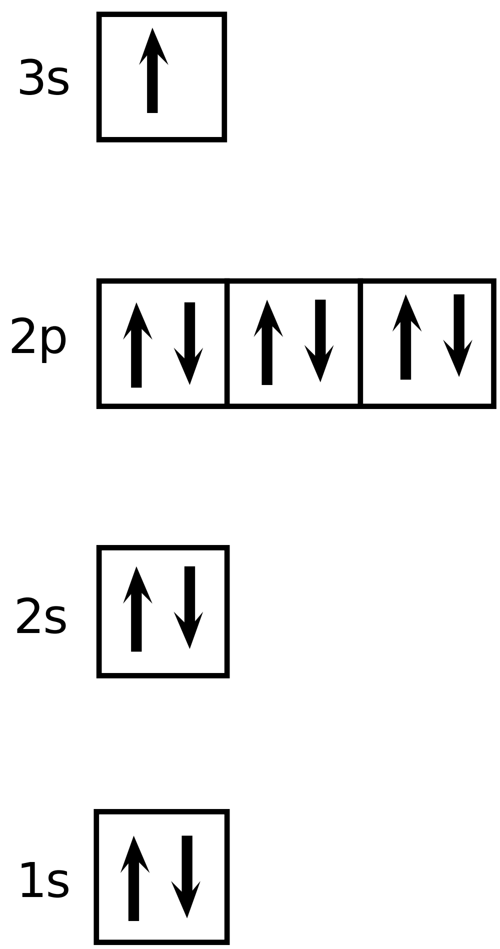

Figure 3.19

Chapter 4

The Periodic Table[1]

4.1 The arrangement of atoms in the periodic table

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

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

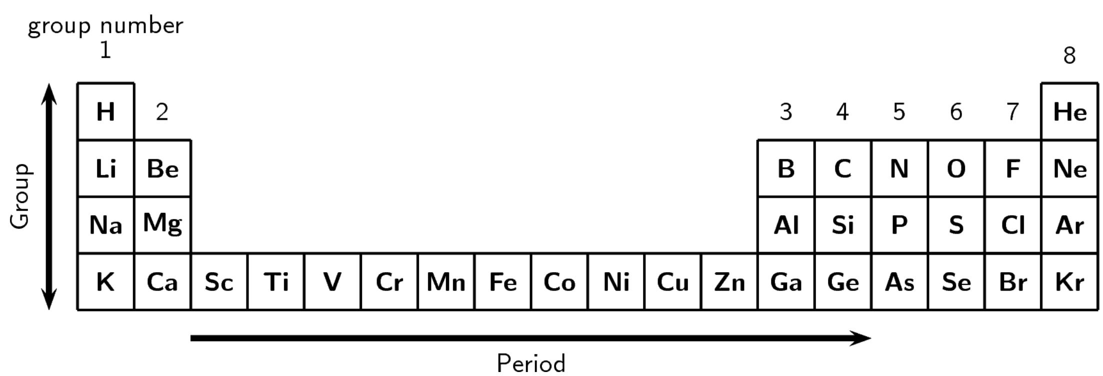

Figure 4.1: A simplified diagram showing part of the Periodic Table

[1] This content is available online at <http://cnx.org/content/m38133/1.6/>.

You can view an online periodic table at Periodic table[2] . The full periodic table is also reproduced at the front of this book.

4.1.1 Activity: Inventing the periodic table

You are the official chemist for the planet Zog. You have discovered all the same elements that we have here on Earth, but you don't have a periodic table. The citizens of Zog want to know how all these elements relate to each other. How would you invent the periodic table? Think about how you would organize the data that you have and what properties you would include. Do not simply copy Mendeleev's ideas, be creative and come up with some of your own. Research other forms of the periodic table and make one that makes sense to you. Present your ideas to your class.

4.1.2 Groups in the periodic table

A *group* is a vertical column in the periodic table and is considered to be the most important way of classifying the elements. If you look at a periodic table, you will see the groups numbered at the top of each column. The groups are numbered from left to right starting with 1 and ending with 18. This is the convention that we will use in this book. On some periodic tables you may see that the groups are numbered from left to right as follows: 1, 2, then an open space which contains the **transition elements**, followed by groups 3 to 8. Another way to label the groups is using Roman numerals. In some groups, the elements display very similar chemical properties and the groups are even given separate names to identify them.

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

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

Image not finished

Figure 4.2: Electron diagrams for some of the Group 1 elements, with sodium and potasium incomplete; to be completed as an excersise.

- *Group 2:* These elements are known as the **alkali earth metals**. Each element only has two valence electrons and so in chemical reactions, the group 2 elements tend to *lose* these electrons so that the energy shells are complete. These elements are less reactive than those in group 1 because it is more difficult to lose two electrons than it is to lose one.
- *Group 13* elements have three valence electrons.
- *Group 16:* These elements are sometimes known as the chalcogens. These elements are fairly reactive and tend to gain electrons to fill their outer shell.
- *Group 17:* These elements are known as the **halogens**. Each element is missing just one electron from its outer energy shell. These elements tend to *gain* electrons to fill this shell, rather than losing them. These elements are also very reactive.
- *Group 18:* These elements are the **noble gases**. All of the energy shells of the halogens are full and so these elements are very unreactive.

[2]http://periodictable.com/

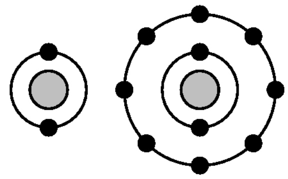

Figure 4.3: Electron diagrams for two of the noble gases, helium (He) and neon (Ne).

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

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

NOTE: Group 15 on the periodic table is sometimes called the pnictogens.

4.1.2.1 Investigation : The properties of elements

Refer to Figure 4.2.

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

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

Khan academy video on the periodic table - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/LDHg7Vgzses&rel=0>

Figure 4.4

4.1.3 Periods in the periodic table

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

- As you move from one group to the next within a period, the number of valence electrons increases by one each time.

- Within a single period, all the valence electrons occur in the same energy shell. If the period increases, so does the energy shell in which the valence electrons occur.
- In general, the diameter of atoms decreases as one moves from left to right across a period. Consider the attractive force between the positively charged nucleus and the negatively charged electrons in an atom. As you move across a period, the number of protons in each atom increases. The number of electrons also increases, but these electrons will still be in the same energy shell. As the number of protons increases, the force of attraction between the nucleus and the electrons will increase and the atomic diameter will decrease.
- Ionisation energy increases as one moves from left to right across a period. As the valence electron shell moves closer to being full, it becomes more difficult to remove electrons. The opposite is true when you move down a *group* in the table because more energy shells are being added. The electrons that are closer to the nucleus 'shield' the outer electrons from the attractive force of the positive nucleus. Because these electrons are not being held to the nucleus as strongly, it is easier for them to be removed and the ionisation energy decreases.
- In general, the reactivity of the elements decreases from left to right across a period.
- The formation of halides follows the general pattern: XCl_n (where X is any element in a specific group and n is the number of that specific group.). For example, the formula for the halides of group 1 will be XCl, for the second group the halides have the formula XCl_2 and in the third group the halides have the formula XCl_3. This should be easy to see if you remember the valency of the group and of the halides.

The formation of oxides show a trend as you move across a period. This should be easy to see if you think about valency. In the first group all the elements lose an electron to form a cation. So the formula for an oxide will be X_2O. In the second group (moving from left to right across a period) the oxides have the formula XO. In the third group the oxides have the formula X_2O_3.

Several other trends may be observed across a period such as density, melting points and boiling points. These trends are not as obvious to see as the above trends and often show variations to the general trend.

Electron affinity and electronegativity also show some general trends across periods. Electron affinity can be thought of as how much an element wants electrons. Electron affinity generally increases from left to right across a period. Electronegativity is the tendency of atoms to attract electrons. The higher the electronegativity, the greater the atom attracts electrons. Electronegativity generally increases across a period (from left to right). Electronegativity and electron affinity will be covered in more detail in a later grade.

You may see periodic tables labeled with s-block, p-block, d-block and f-block. This is simply another way to group the elements. When we group elements like this we are simply noting which orbitals are being filled in each block. This method of grouping is not very useful to the work covered at this level.

Using the properties of the groups and the trends that we observe in certain properties (ionization energy, formation of halides and oxides, melting and boiling points, atomic diameter) we can predict the the properties of unknown elements. For example, the properties of the unfamiliar elements Francium (Fr), Barium (Ba), Astatine (At), and Xenon (Xe) can be predicted by knowing their position on the periodic table. Using the periodic table we can say: Francium (Group 1) is an alkali metal, very reactive and needs to lose 1 electron to obtain a full outer energy shell; Barium (Group 2) is an alkali earth metal and needs to lose 2 electrons to achieve stability; Astatine (Group 7) is a halogen, very reactive and needs to gain 1 electron to obtain a full outer energy shell; and Xenon (Group 8) is a noble gas and thus stable due to its full outer energy shell. This is how scientists are able to say what sort of properties the atoms in the last period have. Almost all of the elements in this period do not occur naturally on earth and are made in laboratories. These atoms do not exist for very long (they are very unstable and break apart easily) and so measuring their properties is difficult.

4.1.3.1 Exercise: Elements in the periodic table

Refer to the elements listed below:

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

Which of the elements listed above:

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

Click here for the solution[3]

4.2 Ionisation Energy and the Periodic Table

4.2.1 Ions

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

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

Definition 4.1: Ion
An ion is a charged atom. A positively charged ion is called a **cation** e.g. Na^{+}, and a negatively charged ion is called an **anion** e.g. F^{-}. The charge on an ion depends on the number of electrons that have been lost or gained.

But how do we know how many electrons an atom will gain or lose? Remember what we said about stability? We said that all atoms are trying to get a full outer shell. For the elements on the left hand side of the periodic table the easiest way to do this is to lose electrons and for the elements on the right of the periodic table the easiest way to do this is to gain electrons. So the elements on the left of the periodic table will form cations and the elements on the right hand side of the periodic table will form anions. By doing this the elements can be in the most stable electronic configuration and so be as stable as the noble gases.

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

Lithium
A lithium atom loses one electron to form a positive ion:

[3]See the file at <http://cnx.org/content/m38133/latest/http://www.fhsst.org/liw>

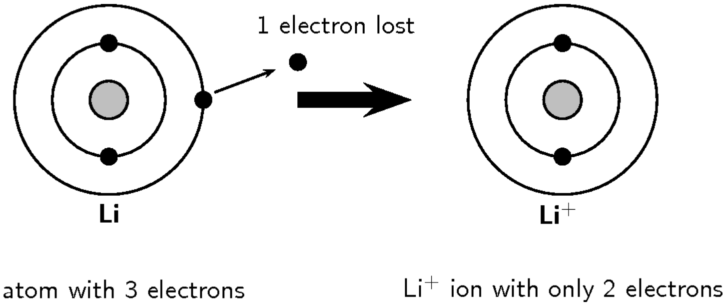

Figure 4.5: The arrangement of electrons in a lithium ion.

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

Fluorine

A fluorine atom gains one electron to form a negative ion:

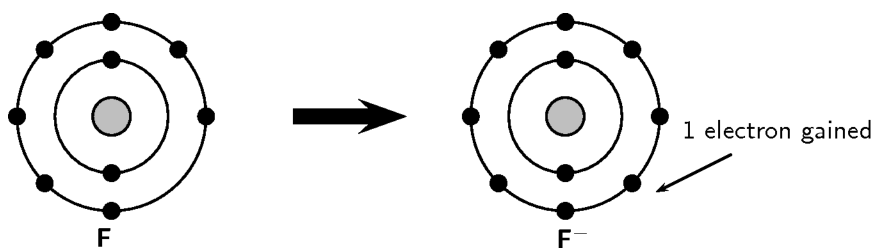

Figure 4.6: The arrangement of electrons in a fluorine ion.

You should have noticed in both these examples that each element lost or gained electrons to make a full outer shell.

4.2.1.1 Investigation : The formation of ions

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

Observations:

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

- In each case the number of electrons that is either gained or lost, is the same as the number of electrons that are needed for the atoms to achieve a full outer energy level.

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

4.2.1.2 Exercise: The formation of ions

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

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

Table 4.1

Click here for the solution[4]

4.2.2 Ionisation Energy

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

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

4.2.2.1 Trends in ionisation energy

Refer to the data table below which gives the ionisation energy (in $\text{kJ} \cdot \text{mol}^{-1}$) and atomic number (Z) for a number of elements in the periodic table:

[4]See the file at <http://cnx.org/content/m38133/latest/http://www.fhsst.org/lid>

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

Table 4.2

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

Click here for the solution[5]

Khan academy video on periodic table - 2

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/ywqg9PorTAw&rel=0>

Figure 4.7

By now you should have an appreciation of what the periodic table can tell us. The periodic table does not just list the elements, but tells chemists what the properties of elements are, how the elements will combine and many other useful facts. The periodic table is truly an amazing resource. Into one simple table, chemists have packed so many facts and data that can easily be seen with a glance. The periodic table is a crucial part of chemistry and you should never go to science class without it.

The following presentation provides a summary of the periodic table

This media object is a Flash object. Please view or download it at
<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=periodictableintroduction-100510044643-phpapp01&stripped_title=periodic-table-introduction>

Figure 4.8

[5] See the file at <http://cnx.org/content/m38133/latest/http://www.fhsst.org/liv>

4.3 Summary

- Elements are arranged in periods and groups on the periodic table. The elements are arranged according to increasing atomic number.
- A **group** is a column on the periodic table containing elements with similar properties. A **period** is a row on the periodic table.
- The groups on the periodic table are labeled from 1 to 8. The first group is known as the alkali metals, the second group is known as the alkali earth metals, the seventh group is known as the halogens and the eighth group is known as the noble gases. Each group has the same properties.
- Several trends such as ionisation energy and atomic diameter can be seen across the periods of the periodic table
- An **ion** is a charged atom. A **cation** is a positively charged ion and an **anion** is a negatively charged ion.
- When forming an ion, an atom will lose or gain the number of electrons that will make its valence energy level full.
- An element's **ionisation energy** is the energy that is needed to remove one electron from an atom.
- Ionisation energy increases across a **period** in the periodic table.
- Ionisation energy decreases down a **group** in the periodic table.

4.4 End of chapter exercises

1. For the following questions state whether they are true or false. If they are false, correct the statement.

   a. The group 1 elements are sometimes known as the alkali earth metals.
   b. The group 2 elements tend to lose 2 electrons to form cations.
   c. The group 8 elements are known as the noble gases.
   d. Group 7 elements are very unreactive.
   e. The transition elements are found between groups 3 and 4.

   Click here for the solution[6]
2. Give one word or term for each of the following:

   a. A positive ion
   b. The energy that is needed to remove one electron from an atom
   c. A horizontal row on the periodic table
   d. A very reactive group of elements that is missing just one electron from their outer shells.

   Click here for the solution[7]
3. For each of the following elements give the ion that will be formed:

   a. sodium
   b. bromine
   c. magnesium
   d. oxygen

   Click here for the solution[8]
4. The following table shows the first ionisation energies for the elements of period 1 and 2.

[6]http://www.fhsst.org/l4Z
[7]http://www.fhsst.org/l4K
[8]http://cnx.org/content/m38133/latest/ http://www.fhsst.org/l4B

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

Table 4.3

a. What is the meaning of the term *first ionisation energy*?
b. Identify the pattern of first ionisation energies in a period.
c. Which TWO elements exert the strongest attractive forces on their electrons? Use the data in the table to give a reason for your answer.
d. Draw Aufbau diagrams for the TWO elements you listed in the previous question and explain why these elements are so stable.
e. It is safer to use helium gas than hydrogen gas in balloons. Which property of helium makes it a safer option?
f. 'Group 1 elements readily form positive ions'. Is this statement correct? Explain your answer by referring to the table.

Click here for the solution[9]

[9] See the file at <http://cnx.org/content/m38133/latest/http://www.fhsst.org/li6>

Chapter 5

Chemical Bonding[1]

5.1 Chemical Bonding

When you look at the matter, or physical substances, around you, you will realise that atoms seldom exist on their own. More often, the things around us are made up of different atoms that have been joined together. This is called **chemical bonding**. Chemical bonding is one of the most important processes in chemistry because it allows all sorts of different molecules and combinations of atoms to form, which then make up the objects in the complex world around us.

5.2 What happens when atoms bond?

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

Definition 5.1: Chemical bond

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

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

You need to remember that it is the *valence electrons* that are involved in bonding and that atoms will try to fill their outer energy levels so that they are more stable (or are more like the noble gases which are very stable).

5.3 Covalent Bonding

5.3.1 The nature of the covalent bond

Covalent bonding occurs between the atoms of **non-metals**. The outermost orbitals of the atoms overlap so that unpaired electrons in each of the bonding atoms can be shared. By overlapping orbitals, the outer energy shells of all the bonding atoms are filled. The shared electrons move in the orbitals around *both*

[1] This content is available online at <http://cnx.org/content/m38131/1.6/>.

atoms. As they move, there is an attraction between these negatively charged electrons and the positively charged nuclei, and this force holds the atoms together in a covalent bond.

Definition 5.2: Covalent bond
Covalent bonding is a form of chemical bonding where pairs of electrons are shared between atoms.

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

Exercise 5.1: Covalent bonding ***(Solution on p. 95.)***
How do hydrogen and chlorine atoms bond covalently in a molecule of hydrogen chloride?

Exercise 5.2: Covalent bonding involving multiple bonds ***(Solution on p. 95.)***
How do nitrogen and hydrogen atoms bond to form a molecule of ammonia (NH_3)?

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

Exercise 5.3: Covalent bonding involving a double bond ***(Solution on p. 96.)***
How do oxygen atoms bond covalently to form an oxygen molecule?

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

Definition 5.3: Valency
The number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.

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

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

5.3.1.1 Covalent bonding and valency

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

[2] http://www.fhsst.org/lOY

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

Table 5.1

Click here for the solution.[3]

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

   Click here for the solution.[4]

5.3.2 Properties of covalent compounds

Covalent compounds have several properties that distinguish them from ionic compounds and metals. These properties are:

1. The melting and boiling points of covalent compounds is generally lower than that for ionic compounds.
2. Covalent compounds are generally more flexible than ionic compounds. The molecules in covalent compounds are able to move around to some extent and can sometimes slide over each other (as is the case with graphite, this is why the lead in your pencil feels slightly slippery). In ionic compounds all the ions are tightly held in place.
3. Covalent compounds generally are not very soluble in water.
4. Covalent compounds generally do not conduct electricity when dissolved in water. This is because they do not dissociate as ionic compounds do.

5.4 Lewis notation and molecular structure

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

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

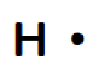

Figure 5.1

[3] http://www.fhsst.org/lOr
[4] http://www.fhsst.org/lO1

A chlorine atom would look like this:

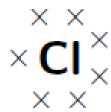

Figure 5.2

A molecule of hydrogen chloride would be shown like this:

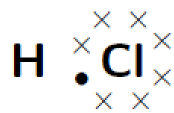

Figure 5.3

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

Exercise 5.4: Lewis notation: Simple molecules ***(Solution on p. 96.)***

Represent the molecule H_2O using Lewis notation

Exercise 5.5: Lewis notation: Molecules with multiple bonds ***(Solution on p. 97.)***

Represent the molecule HCN (hydrogen cyanide) using Lewis notation

Exercise 5.6: Lewis notation: Atoms with variable valencies ***(Solution on p. 97.)***

Represent the molecule H_2S (hydrogen sulphide) using Lewis notation

Another way of representing molecules is using **Couper notation**. In this case, only the electrons that are involved in the bond between the atoms are shown. A line is used for each covalent bond. Using Couper notation, a molecule of water and a molecule of HCN would be represented as shown in figures Figure 5.4 and Figure 5.5 below.

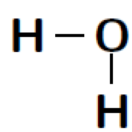

Figure 5.4: A water molecule represented using Couper notation

Figure 5.5: A molecule of HCN represented using Couper notation

5.4.1 Atomic bonding and Lewis notation

1. Represent each of the following *atoms* using Lewis notation:
    a. beryllium
    b. calcium
    c. lithium

   Click here for the solution.[5]
2. Represent each of the following *molecules* using Lewis notation:
    a. bromine gas (Br_2)
    b. carbon dioxide (CO_2)

   Click here for the solution.[6]
3. Which of the two molecules listed above contains a double bond?
   Click here for the solution.[7]
4. Two chemical reactions are described below.
    - nitrogen and hydrogen react to form NH_3
    - carbon and hydrogen bond to form a molecule of CH_4

   For each reaction, give:
    a. the valency of each of the atoms involved in the reaction
    b. the Lewis structure of the product that is formed
    c. the chemical formula of the product
    d. the name of the product

   Click here for the solution.[8]
5. A chemical compound has the following Lewis notation:

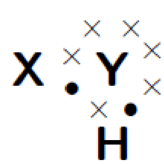

Figure 5.6

[5]http://www.fhsst.org/lOC
[6]http://www.fhsst.org/lOa
[7]http://www.fhsst.org/lOa
[8]http://www.fhsst.org/lOx

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

Click here for the solution.[9]

5.5 Ionic Bonding

5.5.1 The nature of the ionic bond

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

Ionic bonding takes place when the difference in electronegativity between the two atoms is more than 1,7. This usually happens when a metal atom bonds with a non-metal atom. When the difference in electronegativity is large, one atom will attract the shared electron pair much more strongly than the other, causing electrons to be transferred from one atom to the other.

Definition 5.4: Ionic bond
An ionic bond is a type of chemical bond based on the electrostatic forces between two oppositely-charged ions. When ionic bonds form, a metal donates one or more electrons, due to having a low electronegativity, to form a positive ion or cation. The non-metal atom has a high electronegativity, and therefore readily gains electrons to form a negative ion or anion. The two ions are then attracted to each other by electrostatic forces.

Example 1:
In the case of $NaCl$, the difference in electronegativity is 2,1. Sodium has only one valence electron, while chlorine has seven. Because the electronegativity of chlorine is higher than the electronegativity of sodium, chlorine will attract the valence electron of the sodium atom very strongly. This electron from sodium is transferred to chlorine. Sodium loses an electron and forms an Na^+ ion. Chlorine gains an electron and forms an Cl^- ion. The attractive force between the positive and negative ion holds the molecule together.

The balanced equation for the reaction is:

$$Na + Cl \rightarrow NaCl \tag{5.1}$$

This can be represented using Lewis notation:

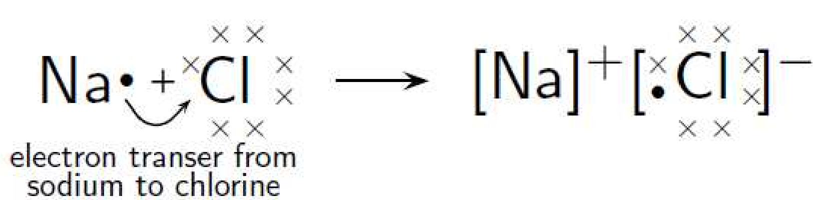

Figure 5.7: Ionic bonding in sodium chloride

[9] http://www.fhsst.org/lOc

Example 2:

Another example of ionic bonding takes place between magnesium (Mg) and oxygen (O) to form magnesium oxide (MgO). Magnesium has two valence electrons and an electronegativity of 1,2, while oxygen has six valence electrons and an electronegativity of 3,5. Since oxygen has a higher electronegativity, it attracts the two valence electrons from the magnesium atom and these electrons are transferred from the magnesium atom to the oxygen atom. Magnesium loses two electrons to form Mg^{2+}, and oxygen gains two electrons to form O^{2-}. The attractive force between the oppositely charged ions is what holds the molecule together.

The balanced equation for the reaction is:

$$2Mg + O_2 \rightarrow 2MgO \tag{5.2}$$

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

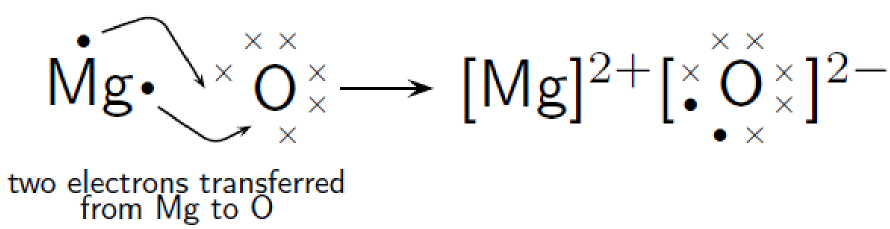

Figure 5.8: Ionic bonding in magnesium oxide

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

5.5.1.1 Ionic compounds

1. Explain the difference between a *covalent* and an *ionic* bond.
   Click here for the solution[10]
2. Magnesium and chlorine react to form magnesium chloride.
   a. What is the difference in electronegativity between these two elements?
   b. Give the chemical formula for:
      i. a magnesium ion
      ii. a chloride ion
      iii. the ionic compound that is produced during this reaction
   c. Write a balanced chemical equation for the reaction that takes place.

   Click here for the solution[11]
3. Draw Lewis diagrams to represent the following ionic compounds:
   a. sodium iodide (NaI)
   b. calcium bromide ($CaBr_2$)
   c. potassium chloride (KCl)

[10]http://www.fhsst.org/lOq
[11]See the file at <http://cnx.org/content/m38131/latest/http://www.fhsst.org/lOl>

Click here for the solution[12]

5.5.2 The crystal lattice structure of ionic compounds

Ionic substances are actually a combination of lots of ions bonded together into a giant molecule. The arrangement of ions in a regular, geometric structure is called a **crystal lattice**. So in fact $NaCl$ does not contain one Na and one Cl ion, but rather a lot of these two ions arranged in a crystal lattice where the ratio of Na to Cl ions is 1:1. The structure of a crystal lattice is shown in Figure 5.9.

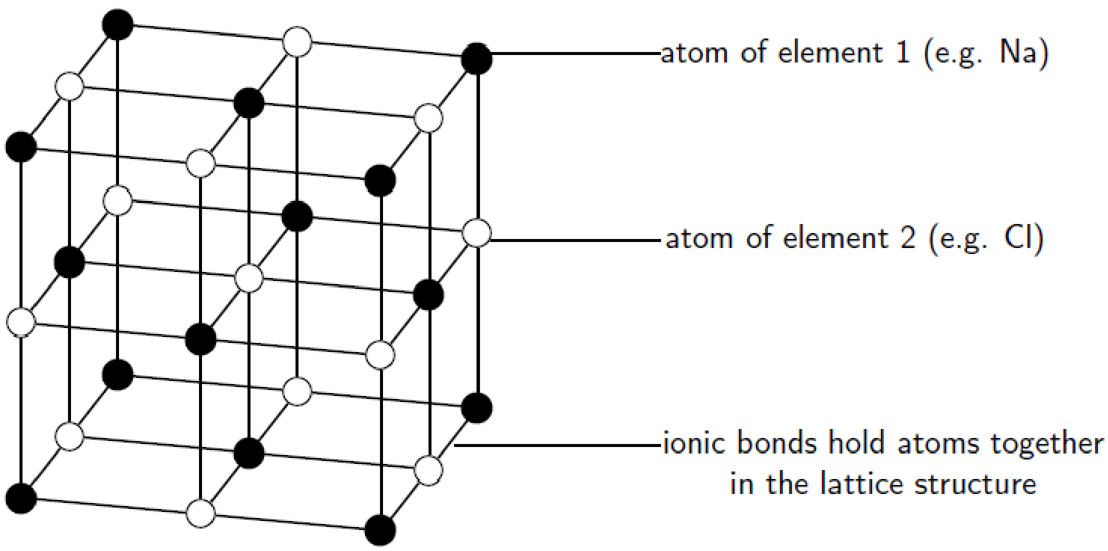

Figure 5.9: The crystal lattice arrangement in an ionic compound (e.g. $NaCl$)

5.5.3 Properties of Ionic Compounds

Ionic compounds have a number of properties:

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

[12]See the file at <http://cnx.org/content/m38131/latest/http://www.fhsst.org/lOi>

5.6 Metallic bonds

5.6.1 The nature of the metallic bond

The structure of a metallic bond is quite different from covalent and ionic bonds. In a metal bond, the valence electrons are *delocalised*, meaning that an atom's electrons do not stay around that one nucleus. In a metallic bond, the positive atomic nuclei (sometimes called the 'atomic kernels') are surrounded by a sea of delocalised electrons which are attracted to the nuclei (Figure 5.10).

Definition 5.5: Metallic bond
Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal.

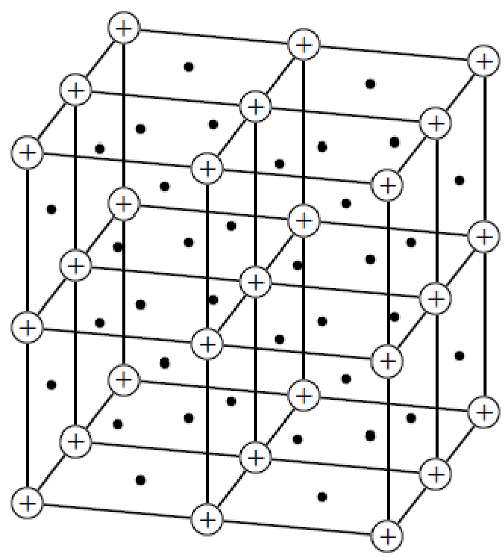

Figure 5.10: Positive atomic nuclei (+) surrounded by delocalised electrons (•)

5.6.2 The properties of metals

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

- *Thermal conductors* Metals are good conductors of heat and are therefore used in cooking utensils such as pots and pans. Because the electrons are loosely bound and are able to move, they can transport heat energy from one part of the material to another.
- *Electrical conductors* Metals are good conductors of electricity, and are therefore used in electrical conducting wires. The loosely bound electrons are able to move easily and to transfer charge from one part of the material to another.
- *Shiny metallic lustre* Metals have a characteristic shiny appearance and are often used to make jewellery. The loosely bound electrons are able to absorb and reflect light at all frequencies, making metals look polished and shiny.

- *Malleable and ductile* This means that they can be bent into shape without breaking (malleable) and can be stretched into thin wires (ductile) such as copper, which can then be used to conduct electricity. Because the bonds are not fixed in a particular direction, atoms can slide easily over one another, making metals easy to shape, mould or draw into threads.
- *Melting point* Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged. The high melting point is due to the high strength of metallic bonds.
- *Density* Metals have a high density because their atoms are packed closely together.

5.6.3 Activity: Building models

Using coloured balls and sticks (or any other suitable materials) build models of each type of bonding. Think about how to represent each kind of bonding. For example, covalent bonding could be represented by simply connecting the balls with sticks to represent the molecules, while for ionic bonding you may wish to construct part of the crystal lattice. Do some research on types of crystal lattices (although the section on ionic bonding only showed the crystal lattice for sodium chloride, many other types of lattices exist) and try to build some of these. Share your findings with your class and compare notes to see what types of crystal lattices they found. How would you show metallic bonding?

You should spend some time doing this activity as it will really help you to understand how atoms combine to form molecules and what the differences are between the types of bonding.

Khan academy video on bonding - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/CGA8sRwqIFg&rel=0>

Figure 5.11

5.6.4 Chemical bonding

1. Give two examples of everyday objects that contain..
   a. covalent bonds
   b. ionic bonds
   c. metallic bonds

   Click here for the solution[13]
2. Complete the table which compares the different types of bonding:

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

[13]http://www.fhsst.org/l3h

Table 5.2

Click here for the solution[14]

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

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

Table 5.3

Click here for the solution[15]

4. Which of these substances will conduct electricity most effectively? Give a reason for your answer. Click here for the solution[16]
5. Use your knowledge of the different types of bonding to explain the following statements:

   a. Swimming during an electric storm (i.e. where there is lightning) can be very dangerous.
   b. Most jewellery items are made from metals.
   c. Plastics are good insulators.

   Click here for the solution[17]

5.7 Writing chemical formulae

5.7.1 The formulae of covalent compounds

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

Exercise 5.7: Formulae of covalent compounds ***(Solution on p. 98.)***
Write the chemical formula for water

Exercise 5.8: Formulae of covalent compounds ***(Solution on p. 98.)***
Write the chemical formula for magnesium oxide

Exercise 5.9: Formulae of covalent compounds ***(Solution on p. 98.)***
Write the chemical formula for copper (II) chloride.

[14]http://www.fhsst.org/l3u
[15]http://www.fhsst.org/l3J
[16]http://www.fhsst.org/l3J
[17]http://www.fhsst.org/l3S

5.7.2 The formulae of ionic compounds

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

Some ions are made up of groups of atoms, and these are called **compound ions**. It is a good idea to learn the compound ions that are shown in Table 5.4

Name of compound ion	**formula**	**Name of compound ion**	**formula**
Carbonate	CO_3^{2-}	Nitrate	NO_2^-
Sulphate	SO_4^{2-}	Hydrogen sulphite	HSO_3^-
Hydroxide	OH^-	Hydrogen sulphate	HSO_4^-
Ammonium	NH_4^+	Dihydrogen phosphate	$H_2PO_4^-$
Nitrate	NO_3^-	Hypochlorite	ClO^-
Hydrogen carbonate	HCO_3^-	Acetate (ethanoate)	CH_3COO^-
Phosphate	PO_4^{3-}	Oxalate	$C_2O_4^{2-}$
Chlorate	ClO_3^-	Oxide	O^{2-}
Cyanide	CN^-	Peroxide	O_2^{2-}
Chromate	CrO_4^{2-}	Sulphide	S^{2-}
Permanganate	MnO_4^-	Sulphite	SO_3^{2-}
Thiosulphate	$S_2O_3^{2-}$	Manganate	MnO_4^{2-}
Phosphide	P^{3-}	Hydrogen phosphate	HPO_4^{3-}

Table 5.4: Table showing common compound ions and their formulae

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

Exercise 5.10: Formulae of ionic compounds *(Solution on p. 98.)*
Write the chemical formula for potassium iodide.

Exercise 5.11: Formulae of ionic compounds *(Solution on p. 99.)*
Write the chemical formula for sodium sulphate.

Exercise 5.12: Formulae of ionic compounds *(Solution on p. 99.)*
Write the chemical formula for calcium hydroxide.

NOTE: Notice how in the last example we wrote OH^- inside brackets. We do this to indicate that OH^- is a complex ion and that there are two of these ions bonded to one calcium ion.

5.7.2.1 Chemical formulae

1. Copy and complete the table below:

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

Table 5.5

   Click here for the solution[18]

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

   Click here for the solution[19]

5.8 Chemical compounds: names and masses

In Giving names and formulae to substances (Section 1.4: Giving names and formulae to substances) the names of chemical compounds was revised. The relative molecular mass for covalent molecules is simply the sum of the relative atomic masses of each of the individual atoms in that compound. For ionic compounds we use the formula of the compound to work out a relative formula mass. We ignore the fact that there are many molecules linked together to form a crystal lattice. For example $NaCl$ has a relative formula mass of $58\ g \cdot \text{mol}^{-1}$.

This media object is a Flash object. Please view or download it at <http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=bonding-100511043651-phpapp02&stripped_title=bonding-4047232&userName=kwarne>

Figure 5.12

5.9 Summary

- A **chemical bond** is the physical process that causes atoms and molecules to be attracted together and to be bound in new compounds.

[18] http://www.fhsst.org/l3t
[19] http://www.fhsst.org/l3z

- Atoms are more **reactive**, and therefore more likely to bond, when their outer electron orbitals are not full. Atoms are less reactive when these outer orbitals contain the maximum number of electrons. This explains why the noble gases do not combine to form molecules.
- When atoms bond, electrons are either shared or exchanged.
- **Covalent bonding** occurs between the atoms of non-metals and involves a sharing of electrons so that the orbitals of the outermost energy levels in the atoms are filled.
- The **valency** of an atom is the number of electrons in the outer shell of that atom and valence electrons are able to form bonds with other atoms.
- A **double** or **triple bond** occurs if there are two or three electron pairs that are shared between the same two atoms.
- A **dative covalent bond** is a bond between two atoms in which both the electrons that are shared in the bond come from the same atom.
- **Lewis** and **Couper** notation are two ways of representing molecular structure. In Lewis notation, dots and crosses are used to represent the valence electrons around the central atom. In Couper notation, lines are used to represent the bonds between atoms.
- An **ionic bond** occurs between atoms where the difference in electronegativity is greater than 1,7. An exchange of electrons takes place and the atoms are held together by the electrostatic force of attraction between oppositely-charged ions.
- Ionic solids are arranged in a **crystal lattice** structure.
- Ionic compounds have a number of specific **properties**, including their high melting and boiling points, brittle nature, the lattice structure of solids and the ability of ionic solutions to conduct electricity.
- A **metallic bond** is the electrostatic attraction between the positively charged nuclei of metal atoms and the delocalised electrons in the metal.
- Metals also have a number of properties, including their ability to conduct heat and electricity, their metallic lustre, the fact that they are both malleable and ductile, and their high melting point and density.
- The valency of atoms, and the way they bond, can be used to determine the **chemical formulae** of compounds.

5.9.1 End of chapter exercises

1. Explain the meaning of each of the following terms
   a. Valency
   b. Covalent bond

   Click here for the solution[20]
2. Which ONE of the following best describes the bond formed between an H^+ ion and the NH_3 molecule?
   a. Covalent bond
   b. Dative covalent (coordinate covalent) bond
   c. Ionic Bond
   d. Hydrogen Bond

   Click here for the solution[21]
3. Which of the following reactions will *not* take place? Explain your answer.
   a. $H + H \rightarrow H_2$
   b. $\mathrm{Ne} + \mathrm{Ne} \rightarrow \mathrm{Ne}_2$
   c. $\mathrm{Cl} + \mathrm{Cl} \rightarrow \mathrm{Cl}_2$

[20] http://www.fhsst.org/lgV
[21] http://www.fhsst.org/l37

Click here for the solution[22]

4. Draw the Lewis structure for each of the following:
   a. calcium
   b. iodine (Hint: Which group is it in? It will be similar to others in that group)
   c. hydrogen bromide (HBr)
   d. nitrogen dioxide (NO_2)

   Click here for the solution[23]

5. Given the following Lewis structure, where X and Y each represent a different element...

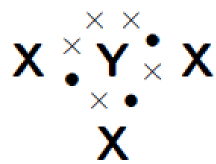

Figure 5.13

   a. What is the valency of X?
   b. What is the valency of Y?
   c. Which elements could X and Y represent?

   Click here for the solution[24]

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

[22]http://www.fhsst.org/l36
[23]http://www.fhsst.org/l3H
[24]http://www.fhsst.org/l3s

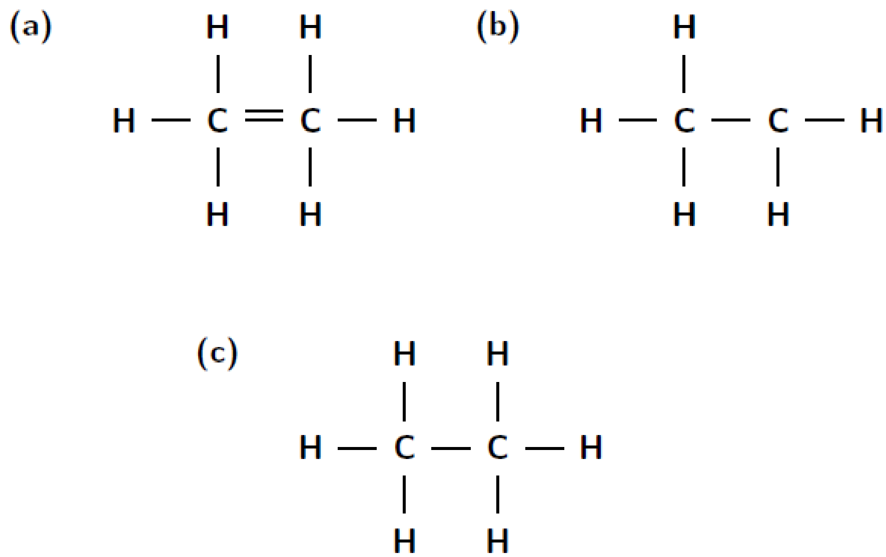

Figure 5.14

Click here for the solution[25]

7. Potassium dichromate is dissolved in water.
   a. Give the name and chemical formula for each of the ions in solution.
   b. What is the chemical formula for potassium dichromate?

   Click here for the solution[26]

[25]http://www.fhsst.org/l3o
[26]http://www.fhsst.org/l3A

Solutions to Exercises in Chapter 5

Solution to Exercise 5.1 (p. 80)

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

Step 2. Chlorine has 7 valence electrons. One of these electrons is unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3. The hydrogen atom needs one more electron to complete its valence shell. The chlorine atom also needs one more electron to complete its shell. Therefore *one pair of electrons* must be shared between the two atoms. In other words, one electron from the chlorine atom will spend some of its time orbiting the hydrogen atom so that hydrogen's valence shell is full. The hydrogen electron will spend some of its time orbiting the chlorine atom so that chlorine's valence shell is also full. A molecule of hydrogen chloride is formed (Figure 5.15). Notice the shared electron pair in the overlapping orbitals.

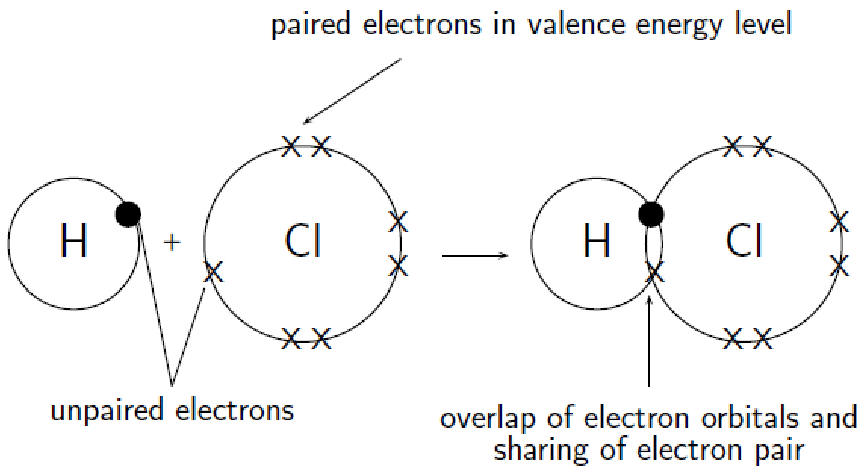

Figure 5.15: Covalent bonding in a molecule of hydrogen chloride

Solution to Exercise 5.2 (p. 80)

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

Step 2. Nitrogen has 5 valence electrons meaning that 3 electrons are unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3. Each hydrogen atom needs one more electron to complete its valence energy shell. The nitrogen atom needs three more electrons to complete its valence energy shell. Therefore *three pairs of electrons* must be shared between the four atoms involved. The nitrogen atom will share three of its electrons so that each of the hydrogen atoms now have a complete valence shell. Each of the hydrogen atoms will share its electron with the nitrogen atom to complete its valence shell (Figure 5.16).

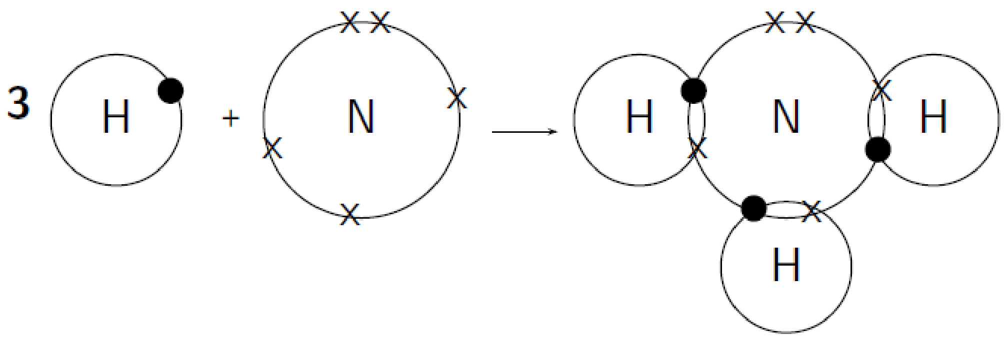

Figure 5.16: Covalent bonding in a molecule of ammonia

Solution to Exercise 5.3 (p. 80)

Step 1. Each oxygen atom has 8 electrons, and their electron configuration is $1s^2 2s^2 2p^4$.
Step 2. Each oxygen atom has 6 valence electrons, meaning that each atom has 2 unpaired electrons.
Step 3. Each oxygen atom needs two more electrons to complete its valence energy shell. Therefore *two pairs of electrons* must be shared between the two oxygen atoms so that both valence shells are full. Notice that the two electron pairs are being shared between *the same two* atoms, and so we call this a **double bond** (Figure 5.17).

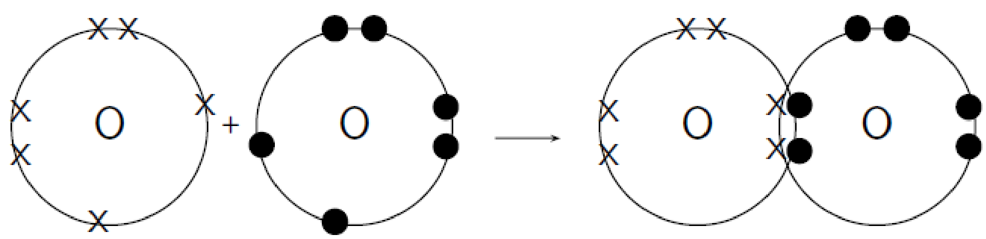

Figure 5.17: A double covalent bond in an oxygen molecule

Solution to Exercise 5.4 (p. 82)

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

Figure 5.18

Step 2. The water molecule is represented below.

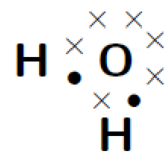

Figure 5.19

Solution to Exercise 5.5 (p. 82)

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

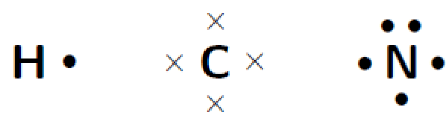

Figure 5.20

Step 2. The HCN molecule is represented below. Notice the three electron pairs between the nitrogen and carbon atom. Because these three covalent bonds are between the same two atoms, this is a *triple* bond.

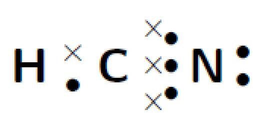

Figure 5.21

Solution to Exercise 5.6 (p. 82)

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

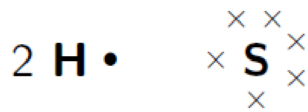

Figure 5.22

Step 2. The H_2S molecule is represented below.

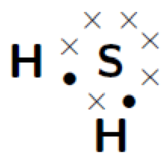

Figure 5.23

Solution to Exercise 5.7 (p. 89)

Step 1. A molecule of water contains the elements *hydrogen* and *oxygen*.
Step 2. The valency of hydrogen is 1 and the valency of oxygen is 2. This means that oxygen can form two bonds with other elements and each of the hydrogen atoms can form one.
Step 3. Using the valencies of hydrogen and oxygen, we know that in a single water molecule, two hydrogen atoms will combine with one oxygen atom. The chemical formula for water is therefore:
H_2O.

Solution to Exercise 5.8 (p. 89)

Step 1. A molecule of magnesium oxide contains the elements *magnesium* and *oxygen*.
Step 2. The valency of magnesium is 2, while the valency of oxygen is also 2. In a molecule of magnesium oxide, one atom of magnesium will combine with one atom of oxygen.
Step 3. The chemical formula for magnesium oxide is therefore:
MgO

Solution to Exercise 5.9 (p. 89)

Step 1. A molecule of copper (II) chloride contains the elements *copper* and *chlorine*.
Step 2. The valency of copper is 2, while the valency of chlorine is 1. In a molecule of copper (II) chloride, two atoms of chlorine will combine with one atom of copper.
Step 3. The chemical formula for copper (II) chloride is therefore:
$CuCl_2$

Solution to Exercise 5.10 (p. 90)

Step 1. Potassium iodide contains potassium and iodide ions.
Step 2. Potassium iodide contains the ions K^+ (valency = 1; charge = +1) and I^- (valency = 1; charge = −1). In order to balance the charge in a single molecule, one atom of potassium will be needed for every one atom of iodine.
Step 3. The chemical formula for potassium iodide is therefore:
KI

Solution to Exercise 5.11 (p. 90)

Step 1. Sodium sulphate contains sodium ions and sulphate ions.
Step 2. Na^+ (valency = 1; charge = +1) and SO_4^{2-}(valency = 2; charge = −2).
Step 3. Two sodium ions will be needed to balance the charge of the sulphate ion. The chemical formula for sodium sulphate is therefore:
Na_2SO_4

Solution to Exercise 5.12 (p. 90)

Step 1. Calcium hydroxide contains calcium ions and hydroxide ions.
Step 2. Calcium hydroxide contains the ions Ca^{2+} (charge = +2) and OH^- (charge = −1). In order to balance the charge in a single molecule, two hydroxide ions will be needed for every ion of calcium.
Step 3. The chemical formula for calcium hydroxide is therefore:
$Ca\,(OH)_2$

Chapter 6

What are the objects around us made of[1]

6.1 Introduction: The atom as the building block of matter

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

The answer lies in the smallest building block of matter: the **atom**. It is the *type* of atoms, and the way in which they are *arranged* in a material, that affects the properties of that substance. This is similar to building materials. We can use bricks, steel, cement, wood, straw (thatch), mud and many other things to build structures from. In the same way that the choice of building material affects the properties of the structure, so the atoms that make up matter affect the properties of matter.

It is not often that substances are found in atomic form (just as you seldom find a building or structure made from one building material). Normally, atoms are bonded (joined) to other atoms to form **compounds** or **molecules**. It is only in the *noble gases* (e.g. helium, neon and argon) that atoms are found individually and are not bonded to other atoms. We looked at some of the reasons for this in earlier chapters.

6.2 Compounds

Definition 6.1: Compound
A compound is a group of two or more different atoms that are attracted to each other by relatively strong forces or bonds.

Almost everything around us is made up of molecules. The only substances that are not made of molecules, but instead are individual atoms are the noble gases. *Water* is made up of molecules, each of which has two hydrogen atoms joined to one oxygen atom. *Oxygen* is a molecule that is made up of two oxygen atoms that are joined to one another. Even the food that we eat is made up of molecules that contain atoms of elements such as carbon, hydrogen and oxygen that are joined to one another in different ways. All of these are known as **small molecules** because there are only a few atoms in each molecule. **Giant molecules** are those where there may be millions of atoms per molecule. Examples of giant molecules are *diamonds*, which are made up of millions of carbon atoms bonded to each other and *metals*, which are made up of millions of metal atoms bonded to each other.

As we learnt in Chapter 5 atoms can share electrons to form covalent bonds or exchange electrons to form ionic bonds. Covalently bonded substances are known as molecular compounds. Ionically bonded

[1]This content is available online at <http://cnx.org/content/m38120/1.6/>.

substances are known as ionic compounds. We also learnt about metallic bonding. In a metal the atoms lose their outermost electrons to form positively charged ions that are arranged in a lattice, while the outermost electrons are free to move amongst the spaces of the lattice.

We can classify covalent molecules into covalent molecular structures and covalent network structures. Covalent molecular structures are simply individual covalent molecules and include water, oxygen, sulphur (S_8) and buckminsterfullerene (C_{60}). All covalent molecular structures are simple molecules. Covalent network structures are giant lattices of covalently bonded molecules, similar to the ionic lattice. Examples include diamond, graphite and silica (SiO_2). All covalent network structures are giant molecules.

Examples of ionic substances are sodium chloride ($NaCl$) and potassium permanganate ($KMnO_4$). Examples of metals are copper, zinc, titanium, gold, etc.

6.2.1 Representing molecules

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

1. **Using formulae to show the structure of a molecule.** A **chemical formula** is an abbreviated (shortened) way of describing a molecule, or some other chemical substance. In the chapter on classification of matter, we saw how chemical compounds can be represented using element symbols from the Periodic Table. A chemical formula can also tell us the *number* of atoms of each element that are in a molecule and their *ratio* in that molecule. For example, the chemical formula for a molecule of carbon dioxide is CO_2 The formula above is called the **molecular formula** of that compound. The formula tells us that in one molecule of carbon dioxide, there is one atom of carbon and two atoms of oxygen. The ratio of carbon atoms to oxygen atoms is 1:2.

   Definition 6.2: Molecular formula
   This is a concise way of expressing information about the atoms that make up a particular chemical compound. The molecular formula gives the exact number of each type of atom in the molecule.

   A molecule of glucose has the molecular formula: $C_6H_{12}O_6$. In each glucose molecule, there are six carbon atoms, twelve hydrogen atoms and six oxygen atoms. The ratio of carbon:hydrogen:oxygen is 6:12:6. We can simplify this ratio to write 1:2:1, or if we were to use the element symbols, the formula would be written as CH_2O. This is called the **empirical formula** of the molecule.

   Definition 6.3: Empirical formula
   This is a way of expressing the *relative* number of each type of atom in a chemical compound. In most cases, the empirical formula does not show the exact number of atoms, but rather the simplest *ratio* of the atoms in the compound.

   The empirical formula is useful when we want to write the formula for a *giant molecule*. Since giant molecules may consist of millions of atoms, it is impossible to say exactly how many atoms are in each molecule. It makes sense then to represent these molecules using their empirical formula. So, in the case of a metal such as copper, we would simply write Cu, or if we were to represent a molecule of sodium chloride, we would simply write NaCl. Chemical formulae therefore tell us something about the *types* of atoms that are in a molecule and the *ratio* in which these atoms occur in the molecule, but they don't give us any idea of what the molecule actually looks like, in other words its *shape*. To show the shape of molecules we can represent molecules using diagrams. Another type of formula that can be used to describe a molecule is its **structural formula**. A structural formula uses a graphical representation to show a molecule's structure (Figure 6.1).

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Figure 6.1: Diagram showing (a) the molecular, (b) the empirical and (c) the structural formula of isobutane

2. **Using diagrams to show the structure of a molecule** Diagrams of molecules are very useful because they help us to picture how the atoms are arranged in the molecule and they help us to see the shape of the molecule. There are two types of diagrams that are commonly used:
    - *Ball and stick models* This is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them. The centres of the atoms (the balls) are connected by straight lines which represent the bonds between them. A simplified example is shown in Figure 6.2.

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Figure 6.2: A ball and stick model of a water molecule

    - *Space-filling model* This is also a 3-dimensional molecular model. The atoms are represented by spheres. Figure 6.3 and Figure 6.4 are some examples of **simple molecules** that are represented in different ways.

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Figure 6.3: A space-filling model and structural formula of a water molecule. Each molecule is made up of two hydrogen atoms that are attached to one oxygen atom. This is a simple molecule.

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Figure 6.4: A space-filling model and structural formula of a molecule of ammonia. Each molecule is made up of one nitrogen atom and three hydrogen atoms. This is a simple molecule.

Figure 6.5 shows the bonds between the carbon atoms in diamond, which is a **giant molecule**. Each carbon atom is joined to four others, and this pattern repeats itself until a complex *lattice* structure is formed. Each black ball in the diagram represents a carbon atom, and each line represents the bond between two carbon atoms. Note that the carbon atoms on the edges are actually bonded to four carbon atoms, but some of these carbon atoms have been omitted.

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Figure 6.5: Diagrams showing the microscopic structure of diamond. The diagram on the left shows part of a diamond lattice, made up of numerous carbon atoms. The diagram on the right shows how each carbon atom in the lattice is joined to four others. This forms the basis of the lattice structure. Diamond is a giant molecule.

NOTE: Diamonds are most often thought of in terms of their use in the jewellery industry. However, about 80% of mined diamonds are unsuitable for use as gemstones and are therefore used in industry because of their strength and hardness. These properties of diamonds are due to the strong covalent bonds (covalent bonding will be explained later) between the carbon atoms in diamond. The most common uses for diamonds in industry are in cutting, drilling, grinding, and polishing.

This website[2] allows you to view several molecules. You do not need to know these molecules, this is simply to allow you to see one way of representing molecules.

[2] http://alteredqualia.com/canvasmol/

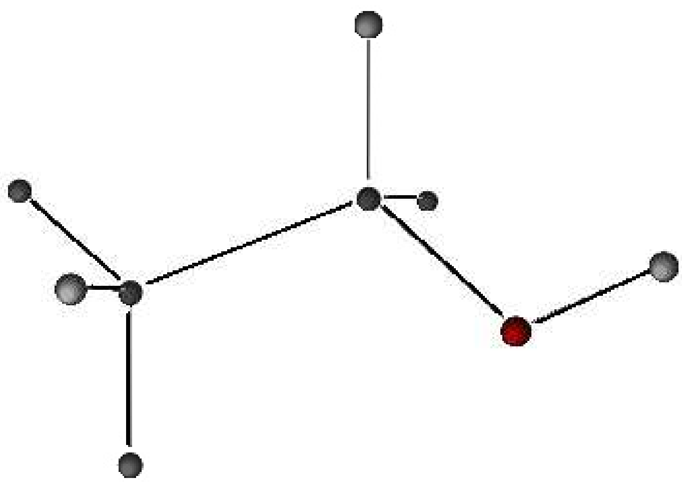


Figure 6.6: Ball-and-stick view of ethanol from http://alteredqualia.com/canvasmol/[3]

6.2.1.1 Atoms and molecules

1. In each of the following, say whether the chemical substance is made up of single atoms, simple molecules or giant molecules.
   a. ammonia gas (NH_3)
   b. zinc metal (Zn)
   c. graphite (C)
   d. nitric acid (HNO_3)
   e. neon gas (He)

   Click here for the solution[4]
2. Refer to the diagram below and then answer the questions that follow:

   Image not finished

   Figure 6.7

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

   Click here for the solution[5]
3. Represent each of the following molecules using its *chemical formula*, *structural formula* and *ball and stick model*.
   a. Hydrogen
   b. Ammonia
   c. sulphur dioxide

   Click here for the solution[6]

6.3 Summary

- The smallest unit of matter is the **atom**. Atoms can combine to form **molecules**.
- A **compound** is a group of two or more atoms that are attracted to each other by chemical bonds.
- A **small molecule** consists of a few atoms per molecule. A **giant molecule** consists of millions of atoms per molecule, for example metals and diamonds.
- The structure of a molecule can be represented in a number of ways.
- The **chemical formula** of a molecule is an abbreviated way of showing a molecule, using the symbols for the elements in the molecule. There are two types of chemical formulae: molecular and empirical formula.
- The **molecular formula** of a molecule gives the exact number of atoms of each element that are in the molecule.
- The **empirical formula** of a molecule gives the relative number of atoms of each element in the molecule.
- Molecules can also be represented using **diagrams**.

[3] http://alteredqualia.com/canvasmol/
[4] http://www.fhsst.org/li5
[5] http://www.fhsst.org/liN
[6] http://www.fhsst.org/liR

- A **ball and stick** diagram is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them.
- A **space-filling model** is also a 3-dimensional molecular model. The atoms are represented by spheres.

- In a molecule, atoms are held together by **chemical bonds** or **intramolecular forces**. Covalent bonds, ionic bonds and metallic bonds are examples of chemical bonds.
- A **covalent bond** exists between non-metal atoms. An **ionic bond** exists between non-metal and metal atoms and a **metallic bond** exists between metal atoms.
- **Intermolecular forces** are the bonds that hold *molecules* together.

6.3.1 End of chapter exercises

1. Give one word or term for each of the following descriptions.
   a. A composition of two or more atoms that act as a unit.
   b. Chemical formula that gives the relative number of atoms of each element that are in a molecule.

   Click here for the solution[7]
2. Give a definition for each of the following terms: descriptions.
   a. molecule
   b. Ionic compound
   c. Covalent network structure
   d. Empirical formula
   e. Ball-and-stick model

   Click here for the solution[8]
3. Ammonia, an ingredient in household cleaners, can be broken down to form one part nitrogen (N) and three parts hydrogen (H). This means that ammonia...
   a. is a colourless gas
   b. is not a compound
   c. cannot be an element
   d. has the formula N_3H

   Click here for the solution[9]
4. Represent each of the following molecules using its chemical formula, its structural formula and the ball-and-stick model:
   a. nitrogen
   b. carbon dioxide
   c. methane
   d. argon

   Click here for the solution[10]

[7] http://www.fhsst.org/l2e
[8] http://www.fhsst.org/l2M
[9] http://www.fhsst.org/liV
[10] http://www.fhsst.org/l2L

Chapter 7

Physical and Chemical Change[1]

7.1 Physical and Chemical Change - Grade 10

Matter is all around us. The desks we sit at, the air we breathe and the water we drink are all examples of matter. But matter doesn't always stay the same. It can change in many different ways. In this chapter, we are going to take a closer look at **physical** and **chemical** changes that occur in matter.

7.2 Physical changes in matter

A **physical change** is one where the particles of the substances that are involved in the change are not broken up in any way. When water is heated for example, the temperature and energy of the water molecules increases and the liquid water evaporates to form water vapour. When this happens, some kind of change has taken place, but the molecular structure of the water has not changed. This is an example of a *physical change*.

$H_2O\,(l) \rightarrow H_2O\,(g)$

Conduction (the transfer of energy through a material) is another example of a physical change. As energy is transferred from one material to another, the *energy* of each material is changed, but not its chemical makeup. Dissolving one substance in another is also a physical change.

Definition 7.1: Physical change

A change that can be seen or felt, but that doesn't involve the break up of the particles in the reaction. During a physical change, the *form* of matter may change, but not its *identity*. A change in temperature is an example of a physical change.

You can think of a physical change as a person who is standing still. When they start to move (start walking) then a change has occurred and this is similar to a physical change.

There are some important things to remember about physical changes in matter:

1. *Arrangement of particles*
   When a physical change occurs, the particles (e.g. atoms, molecules) may re-arrange themselves without actually breaking up in any way. In the example of evaporation that we used earlier, the water molecules move further apart as their temperature (and therefore energy) increases. The same would be true if ice were to melt. In the solid phase, water molecules are packed close together in a very ordered way, but when the ice is heated, the molecules overcome the forces holding them together and they move apart. Once again, the particles have re-arranged themselves, but have not broken up.

$$H_2O\,(s) \rightarrow H_2O\,(l) \tag{7.1}$$

[1] This content is available online at <http://cnx.org/content/m38141/1.8/>.

Figure 7.1 shows this more clearly. In each phase of water, the water molecule itself stays the same, but the way the molecules are arranged has changed. Note that in the solid phase, we simply show the water molecules as spheres. This makes it easier to see how tightly packed the molecules are. In reality the water molecules would all look the same.

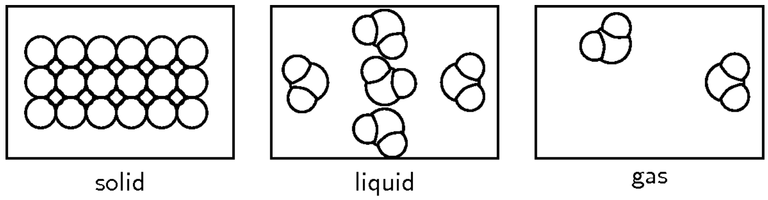


Figure 7.1: The arrangement of water molecules in the three phases of matter

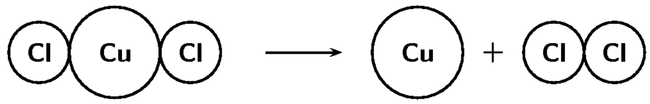
2. *Conservation of mass*
   In a physical change, the total mass, the number of atoms and the number of molecules will always stay the same. In other words you will always have the same number of molecules or atoms at the end of the change as you had at the beginning.
3. *Energy changes*
   Energy changes may take place when there is a physical change in matter, but these energy changes are normally smaller than the energy changes that take place during a chemical change.
4. *Reversibility*
   Physical changes in matter are usually easier to reverse than chemical changes. Water vapour for example, can be changed back to liquid water if the temperature is lowered. Liquid water can be changed into ice by simply decreasing the temperature.

Activity: Physical change
Use plastic pellets or marbles to represent water in the solid state. What do you need to do to the pellets to represent the change from solid to liquid?

7.3 Chemical Changes in Matter

When a **chemical change** takes place, new substances are formed in a chemical reaction. These new products may have very different properties from the substances that were there at the start of the reaction.

The breakdown of copper (II) chloride to form copper and chlorine is an example of chemical change. A simplified diagram of this reaction is shown in Figure 7.2. In this reaction, the initial substance is copper (II) chloride, but once the reaction is complete, the products are copper and chlorine.



$$CuCl_2 \rightarrow Cu + Cl_2$$

Figure 7.2: The decomposition of copper(II) chloride to form copper and chlorine. We write this as: $CuCl_2 \rightarrow Cu + Cl_2$

Definition 7.2: Chemical change
The formation of new substances in a chemical reaction. One type of matter is changed into something different.

There are some important things to remember about chemical changes:

1. *Arrangement of particles*
   During a chemical change, the particles themselves are changed in some way. In the example of copper (II) chloride that was used earlier, the $CuCl_2$ molecules were split up into their component atoms. The number of particles will change because each $CuCl_2$ molecule breaks down into one copper atom (Cu) and one chlorine molecule ($CuCl_2$). However, what you should have noticed, is that the number of atoms of each element stays the same, as does the total mass of the atoms. This will be discussed in more detail in a later section.
2. *Energy changes*
   The energy changes that take place during a chemical reaction are much greater than those that take place during a physical change in matter. During a chemical reaction, energy is used up in order to break bonds, and then energy is released when the new product is formed. This will be discussed in more detail in "Energy changes in chemical reactions" (Section 7.4: Energy changes in chemical reactions).
3. *Reversibility*
   Chemical changes are far more difficult to reverse than physical changes.

We will consider two types of chemical reactions: **decomposition reactions** and **synthesis reactions**.

7.3.1 Decomposition reactions

A **decomposition reaction** occurs when a chemical compound is broken down into elements or smaller compounds. The generalised equation for a decomposition reaction is:
$AB \rightarrow A + B$

One example of such a reaction is the decomposition of mercury (II) oxide (Figure 7.3) to form mercury and oxygen according to the following equation:

$$2HgO \rightarrow 2Hg + O_2 \qquad (7.2)$$

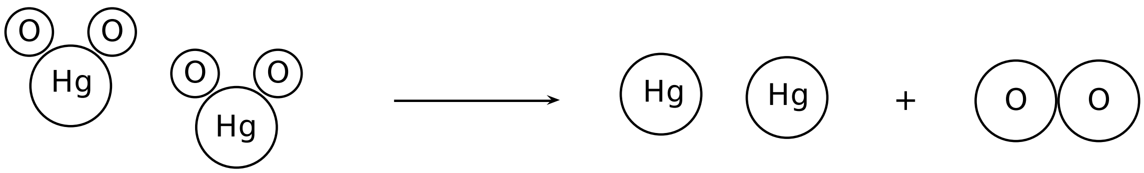


Figure 7.3: The decomposition of HgO to form Hg and O_2

The decomposition of hydrogen peroxide is another example.

7.3.1.1 Experiment : The decomposition of hydrogen peroxide

Aim:

To observe the decomposition of hydrogen peroxide when it is heated.
Apparatus:

Dilute hydrogen peroxide (about 3%); manganese dioxide; test tubes; a water bowl; stopper and delivery tube

Figure 7.4

Method:

1. Put a small amount (about 5 ml) of hydrogen peroxide in a test tube.
2. Set up the apparatus as shown in Figure 7.4
3. Very carefully add a small amount (about 0,5 g) of manganese dioxide to the test tube containing hydrogen peroxide.

Results:

You should observe a gas bubbling up into the second test tube. This reaction happens quite rapidly.
Conclusions:

When hydrogen peroxide is added to manganese dioxide it decomposes to form oxygen and water. The chemical decomposition reaction that takes place can be written as follows:

$$2H_2O_2 \rightarrow 2H_2O + O_2 \quad (7.3)$$

Note that the manganese dioxide is a catalyst and is not shown in the reaction. (A catalyst helps speed up a chemical reaction.)

> NOTE: The previous experiment used the downward displacement of water to collect a gas. This is a very common way to collect a gas in chemistry. The oxygen that is evolved in this reaction moves along the delivery tube and then collects in the top of the test tube. It does this because it

is lighter than water and so displaces the water downwards. If you use a test tube with an outlet attached, you could collect the oxygen into jars and store it for use in other experiments.

The above experiment can be very vigourous and produce a lot of oxygen very rapidly. For this reason you use dilute hydrogen peroxide and only a small amount of manganese dioxide.

7.3.2 Synthesis reactions

During a **synthesis reaction**, a new product is formed from elements or smaller compounds. The generalised equation for a synthesis reaction is as follows:

$$A + B \rightarrow AB \tag{7.4}$$

One example of a synthesis reaction is the burning of magnesium in oxygen to form magnesium oxide(Figure 7.5). The equation for the reaction is:

$$2Mg + O_2 \rightarrow 2MgO \tag{7.5}$$

Image not finished

Figure 7.5: The synthesis of magnesium oxide (MgO) from magnesium and oxygen

7.3.2.1 Experiment: Chemical reactions involving iron and sulphur

Aim:

To demonstrate the synthesis of iron sulphide from iron and sulphur.
Apparatus:

5,6 g iron filings and 3,2 g powdered sulphur; porcelain dish; test tube; Bunsen burner

Image not finished

Figure 7.6

Method:

1. Measure the quantity of iron and sulphur that you need and mix them in a porcelain dish.
2. Take some of this mixture and place it in the test tube. The test tube should be about 1/3 full.
3. This reaction should ideally take place in a fume cupboard. Heat the test tube containing the mixture over the Bunsen burner. Increase the heat if no reaction takes place. Once the reaction begins, you will need to remove the test tube from the flame. Record your observations.

4. Wait for the product to cool before breaking the test tube with a hammer. Make sure that the test tube is rolled in paper before you do this, otherwise the glass will shatter everywhere and you may be hurt.
5. What does the product look like? Does it look anything like the original reactants? Does it have any of the properties of the reactants (e.g. the magnetism of iron)?

Results:

1. After you removed the test tube from the flame, the mixture glowed a bright red colour. The reaction is exothermic and *produces energy*.
2. The product, iron sulphide, is a dark colour and does not share any of the properties of the original reactants. It is an entirely new product.

Conclusions:

A synthesis reaction has taken place. The equation for the reaction is:

$$Fe + S \rightarrow FeS \tag{7.6}$$

7.3.2.2 Investigation : Physical or chemical change?

Apparatus:
Bunsen burner, 4 test tubes, a test tube rack and a test tube holder, small spatula, pipette, magnet, a birthday candle, $NaCl$ (table salt), $0,1\,M$ $AgNO_3$, $6\,M$ HCl, magnesium ribbon, iron filings, sulphur.

WARNING: $AgNO_3$ stains the skin. Be careful when working with it or use gloves.

Method:

1. Place a small amount of wax from a birthday candle into a test tube and heat it over the bunsen burner until it melts. Leave it to cool.
2. Add a small spatula of $NaCl$ to 5 ml water in a test tube and shake. Then use the pipette to add 10 drops of $AgNO_3$ to the sodium chloride solution. NOTE: Please be careful $AgNO_3$ causes bad stains!!
3. Take a 5 cm piece of magnesium ribbon and tear it into 1 cm pieces. Place two of these pieces into a test tube and add a few drops of $6\,M$ HCl. NOTE: Be very careful when you handle this acid because it can cause major burns.
4. Take about 0,5 g iron filings and 0,5 g sulphur. Test each substance with a magnet. Mix the two samples in a test tube and run a magnet alongside the outside of the test tube.
5. Now heat the test tube that contains the iron and sulphur. What changes do you see? What happens now, if you run a magnet along the outside of the test tube?
6. In each of the above cases, record your observations.

Questions:
Decide whether each of the following changes are physical or chemical and give a reason for your answer in each case. Record your answers in the table below:

Description	Physical or chemical change	Reason
melting candle wax		
dissolving $NaCl$		
mixing $NaCl$ with $AgNO_3$		
tearing magnesium ribbon		
adding HCl to magnesium ribbon		
mixing iron and sulphur		
heating iron and sulphur		

Table 7.1

7.4 Energy changes in chemical reactions

All reactions involve some change in energy. During a *physical* change in matter, such as the evaporation of liquid water to water vapour, the energy of the water molecules increases. However, the change in energy is much smaller than in chemical reactions.

When a chemical reaction occurs, some bonds will *break*, while new bonds may *form*. Energy changes in chemical reactions result from the breaking and forming of bonds. For bonds to *break*, energy must be *absorbed*. When new bonds *form*, energy will be *released* because the new product has a lower energy than the 'in between' stage of the reaction when the bonds in the reactants have just been broken.

In some reactions, the energy that must be *absorbed* to break the bonds in the reactants is less than the total energy that is *released* when new bonds are formed. This means that in the overall reaction, energy is *released*. This type of reaction is known as an **exothermic** reaction. In other reactions, the energy that must be *absorbed* to break the bonds in the reactants is more than the total energy that is *released* when new bonds are formed. This means that in the overall reaction, energy must be *absorbed* from the surroundings. This type of reaction is known as an **endothermic** reaction. Most decomposition reactions are endothermic and heating is needed for the reaction to occur. Most synthesis reactions are exothermic, meaning that energy is given off in the form of heat or light.

More simply, we can describe the energy changes that take place during a chemical reaction as:

Total energy absorbed to break bonds - Total energy released when new bonds form

So, for example, in the reaction...

$2Mg + O_2 \rightarrow 2MgO$

Energy is needed to break the $O-O$ bonds in the oxygen molecule so that new $Mg-O$ bonds can be formed, and energy is released when the product (MgO) forms.

Despite all the energy changes that seem to take place during reactions, it is important to remember that energy cannot be created or destroyed. Energy that enters a system will have come from the surrounding environment and energy that leaves a system will again become part of that environment. This is known as the **conservation of energy** principle.

Definition 7.3: Conservation of energy principle

Energy cannot be created or destroyed. It can only be changed from one form to another.

Chemical reactions may produce some very visible and often violent changes. An explosion, for example, is a sudden increase in volume and release of energy when high temperatures are generated and gases are released. For example, NH_4NO_3 can be heated to generate nitrous oxide. Under these conditions, it is highly sensitive and can detonate easily in an explosive exothermic reaction.

7.5 Conservation of atoms and mass in reactions

The total mass of all the substances taking part in a chemical reaction is conserved during a chemical reaction. This is known as the **law of conservation of mass**. The total number of **atoms** of each element also remains the same during a reaction, although these may be arranged differently in the products.

We will use two of our earlier examples of chemical reactions to demonstrate this:

1. The decomposition of hydrogen peroxide into water and oxygen

$2H_2O_2 \rightarrow 2H_2O + O_2$

Image not finished

Figure 7.7

Left hand side of the equation
Total atomic mass = $(4 \times 1) + (4 \times 16) = 68\,u$
Number of atoms of each element = $(4 \times H) + (4 \times O)$
Right hand side of the equation
Total atomic mass = $(4 \times 1) + (4 \times 16) = 68\,u$
Number of atoms of each element = $(4 \times H) + (4 \times O)$
Both the atomic mass and the number of atoms of each element are conserved in the reaction.

2. The synthesis of magnesium and oxygen to form magnesium oxide

$$2Mg + O_2 \rightarrow 2MgO \tag{7.7}$$

Image not finished

Figure 7.8

Left hand side of the equation
Total atomic mass = $(2 \times 24,3) + (2 \times 16) = 80,6\,u$
Number of atoms of each element = $(2 \times Mg) + (2 \times O)$
Right hand side of the equation
Total atomic mass = $(2 \times 24,3) + (2 \times 16) = 80,6\,u$
Number of atoms of each element = $(2 \times Mg) + (2 \times O)$
Both the atomic mass and the number of atoms of each element are conserved in the reaction.

7.5.1 Activity : The conservation of atoms in chemical reactions

Materials:

1. Coloured marbles or small balls to represent atoms. Each colour will represent a different element.
2. Prestik

Method:

1. Choose a reaction from any that have been used in this chapter or any other *balanced* chemical reaction that you can think of. To help to explain this activity, we will use the decomposition reaction of calcium carbonate to produce carbon dioxide and calcium oxide. $CaCO_3 \rightarrow CO_2 + CaO$
2. Stick marbles together to represent the reactants and put these on one side of your table. In this example you may for example join one red marble (calcium), one green marble (carbon) and three yellow marbles (oxygen) together to form the molecule calcium carbonate ($CaCO_3$).
3. Leaving your reactants on the table, use marbles to make the product molecules and place these on the other side of the table.
4. Now count the number of atoms on each side of the table. What do you notice?
5. Observe whether there is any difference between the molecules in the reactants and the molecules in the products.

Discussion
You should have noticed that the number of atoms in the reactants is the same as the number of atoms in the product. The number of atoms is conserved during the reaction. However, you will also see that the molecules in the reactants and products is not the same. The *arrangement of atoms* is not conserved during the reaction.

7.5.2 Experiment: Conservation of matter

Aim:
To prove the law of conservation of matter experimentally.
Materials:
Test tubes; glass beaker; lead (II) nitrate; sodium iodide; hydrochloric acid; bromothymol blue; Cal-C-Vita tablet, plastic bag; rubber band; mass meter
Method:
Reaction 1

1. Carefully weigh out 5 g of lead (II) nitrate.
2. Dissolve the lead nitrate in 100 ml of water.
3. Weigh the lead nitrate solution.
4. Weigh out 4,5 g of sodium iodide and dissolve this in the lead (II) nitrate solution.
5. Weigh the beaker containing the lead nitrate and sodium iodide mixture.

Reaction 2

1. Measure out 20 ml of sodium hydroxide.
2. Add a few drops of bromothymol blue to the sodium hydroxide.
3. Weigh the sodium hydroxide.
4. Weigh 5 ml of hydrochloric acid.
5. Add 5 ml of hydrochloric acid to the sodium hydroxide. Repeat this step until you observe a colour change (this should occur around 20 ml).
6. Weigh the final solution.

Reaction 3

1. Measure out 100 ml of water into a beaker.
2. Weigh the beaker with water in it.
3. Place the Cal-C-Vita tablet into the plastic bag.
4. Weigh the Cal-C-Vita tablet and the plastic bag.
5. Place the plastic bag over the beaker, being careful to not let the tablet fall into the water
6. Seal the bag around the beaker using the rubber band. Drop the tablet into the water.
7. Observe what happens.
8. Weigh the bag and beaker containing the solution.

Results:
Fill in the following table for reactants (starting materials) and products (ending materials) masses. For the second reaction, you will simply take the mass of 5 ml of hydrochloric acid and multiply it by how many amounts you put in, for example, if you put 4 amounts in, then you would have 20 ml and 4 times the mass of 5 ml.

	Reaction 1	Reaction 2	Reaction 3
Reactants			
Products			

Table 7.2

Add the masses for the reactants for each reaction. Do the same for the products. For each reaction compare the mass of the reactants to the mass of the products. What do you notice? Is the mass conserved?

In the experiment above you should have found that the mass at the start of the reaction is the same as the mass at the end of the reaction. You may have found that these masses differed slightly, but this is due to errors in measurements and in performing experiments (all scientists make some errors in performing experiments).

7.6 Law of constant composition

In any given chemical compound, the elements always combine in the same proportion with each other. This is the **law of constant proportion**.

The **law of constant composition** says that, in any particular chemical compound, all samples of that compound will be made up of the same elements in the same proportion or ratio. For example, any water molecule is always made up of two hydrogen atoms and one oxygen atom in a 2:1 ratio. If we look at the relative masses of oxygen and hydrogen in a water molecule, we see that 94% of the mass of a water molecule is accounted for by oxygen and the remaining 6% is the mass of hydrogen. This mass proportion will be the same for any water molecule.

This does not mean that hydrogen and oxygen always combine in a 2:1 ratio to form H_2O. Multiple proportions are possible. For example, hydrogen and oxygen may combine in different proportions to form H_2O_2 rather than H_2O. In H_2O_2, the H:O ratio is 1:1 and the mass ratio of hydrogen to oxygen is 1:16. This will be the same for any molecule of hydrogen peroxide.

7.7 Volume relationships in gases

In a chemical reaction between gases, the relative volumes of the gases in the reaction are present in a ratio of small whole numbers if all the gases are at the same temperature and pressure. This relationship is also known as **Gay-Lussac's Law**.

For example, in the reaction between hydrogen and oxygen to produce water, two volumes of H_2 react with 1 volume of O_2 to produce 2 volumes of H_2O.

$2H_2 + O_2 \rightarrow 2H_2O$

In the reaction to produce ammonia, one volume of nitrogen gas reacts with three volumes of hydrogen gas to produce two volumes of ammonia gas.

$N_2 + 3H_2 \rightarrow 2NH_3$

This relationship will also be true for all other chemical reactions.

7.8 Summary

The following video provides a summary of the concepts covered in this chapter.

Physical and chemical change

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/QL7V3L3dfDM?fs=1&hl=en_US>

Figure 7.9

1. Matter does not stay the same. It may undergo physical or chemical changes.
2. A **physical change** means that the form of matter may change, but not its identity. For example, when water evaporates, the energy and the arrangement of water molecules will change, but not the structure of the water molecules themselves.
3. During a physical change, the **arrangement of particles** may change but the mass, number of atoms and number of molecules will stay the same.
4. Physical changes involve small changes in **energy** and are easily reversible.
5. A chemical change occurs when one or more substances change into other materials. A chemical reaction involves the formation of new substances with **different properties**. For example, magnesium and oxygen react to form magnesium oxide (MgO)
6. A chemical change may involve a **decomposition** or **synthesis** reaction. During chemical change, the mass and number of atoms is conserved, but the number of molecules is not always the same.
7. Chemical reactions involve larger changes in energy. During a reaction, energy is needed to break bonds in the reactants and energy is released when new products form. If the energy released is greater than the energy absorbed, then the reaction is exothermic. If the energy released is less than the energy absorbed, then the reaction is endothermic. Chemical reactions are not easily reversible.
8. Decomposition reactions are usually **endothermic** and synthesis reactions are usually **exothermic**.
9. The **law of conservation of mass** states that the total mass of all the substances taking part in a chemical reaction is conserved and the number of atoms of each element in the reaction does not change when a new product is formed.
10. The **conservation of energy principle** states that energy cannot be created or destroyed, it can only change from one form to another.
11. The **law of constant composition** states that in any particular compound, all samples of that compound will be made up of the same elements in the same proportion or ratio.
12. **Gay-Lussac's Law** states that in a chemical reaction between gases, the relative volumes of the gases in the reaction are present in a ratio of small whole numbers if all the gases are at the same temperature and pressure.

7.8.1 End of chapter exercises

1. For each of the following definitions give one word or term:
   a. A change that can be seen or felt, where the particles involved are not broken up in any way
   b. The formation of new substances in a chemical reaction
   c. A reaction where a new product is formed from elements or smaller compounds

   Click here for the solution[2]

[2]http://www.fhsst.org/l2z

2. State the conservation of energy principle.
   Click here for the solution[3]
3. Explain how a chemical change differs from a physical change.
   Click here for the solution[4]
4. Complete the following table by saying whether each of the descriptions is an example of a physical or chemical change:

Description	**Physical or chemical**
hot and cold water mix together	
milk turns sour	
a car starts to rust	
food digests in the stomach	
alcohol disappears when it is placed on your skin	
warming food in a microwave	
separating sand and gravel	
fireworks exploding	

Table 7.3

   Click here for the solution[5]
5. For each of the following reactions, say whether it is an example of a synthesis or decomposition reaction:
   a. $(NH_4)_2CO_3 \rightarrow NH_3 + CO_2 + H_2O$
   b. $N_2(g) + 3H_2(g) \rightarrow 2NH_3$
   c. $CaCO_3(s) \rightarrow CaO + CO_2$

   Click here for the solution[6]
6. For the following equation: $CaCO_3(s) \rightarrow CaO + CO_2$ show that the 'law of conservation of mass' applies.
   Click here for the solution[7]

[3] http://www.fhsst.org/l2u
[4] http://www.fhsst.org/l2J
[5] http://www.fhsst.org/l3q
[6] http://www.fhsst.org/l3l
[7] http://www.fhsst.org/l3i

Chapter 8

Representing Chemical Change[1]

8.1 Introduction

As we have already mentioned, a number of changes can occur when elements react with one another. These changes may either be *physical* or *chemical*. One way of representing these changes is through **balanced chemical equations**. A chemical equation describes a chemical reaction by using symbols for the elements involved. For example, if we look at the reaction between iron (Fe) and sulphur (S) to form iron sulphide (FeS), we could represent these changes either in words or using chemical symbols:

iron + sulphur → iron sulphide

or

$Fe + S \rightarrow \text{FeS}$

Another example would be:

ammonia + oxygen → nitric oxide + water

or

$4\text{NH}_3 + 5O_2 \rightarrow 4\text{NO} + 6H_2O$

Compounds on the left of the arrow are called the **reactants** and these are needed for the reaction to take place. In this equation, the reactants are ammonia and oxygen. The compounds on the right are called the **products** and these are what is formed from the reaction.

In order to be able to write a balanced chemical equation, there are a number of important things that need to be done:

1. Know the chemical symbols for the elements involved in the reaction
2. Be able to write the chemical formulae for different reactants and products
3. Balance chemical equations by understanding the laws that govern chemical change
4. Know the state symbols for the equation

We will look at each of these steps separately in the next sections.

8.2 Chemical symbols

It is very important to know the chemical symbols for common elements in the Periodic Table, so that you are able to write chemical equations and to recognise different compounds.

8.2.1 Revising common chemical symbols

- Write down the chemical symbols and names of all the elements that you know.

[1] This content is available online at <http://cnx.org/content/m38139/1.5/>.

- Compare your list with another learner and add any symbols and names that you don't have.
- Spend some time, either in class or at home, learning the symbols for at least the first twenty elements in the periodic table. You should also learn the symbols for other common elements that are not in the first twenty.
- Write a short test for someone else in the class and then exchange tests with them so that you each have the chance to answer one.

8.3 Writing chemical formulae

A **chemical formula** is a concise way of giving information about the atoms that make up a particular chemical compound. A chemical formula shows each element by its symbol and also shows how many atoms of each element are found in that compound. The number of atoms (if greater than one) is shown as a subscript.

Examples:

CH_4 (methane)

Number of atoms: (1 × carbon) + (4 × hydrogen) = 5 atoms in one methane molecule

H_2SO_4 (sulphuric acid)

Number of atoms: (2 × hydrogen) + (1 × sulphur) + (4 × oxygen) = 7 atoms in one molecule of sulphuric acid

A chemical formula may also give information about how the atoms are arranged in a molecule if it is written in a particular way. A molecule of ethane, for example, has the chemical formula C_2H_6. This formula tells us how many atoms of each element are in the molecule, but doesn't tell us anything about how these atoms are arranged. In fact, each carbon atom in the ethane molecule is bonded to three hydrogen atoms. Another way of writing the formula for ethane is CH_3CH_3. The number of atoms of each element has not changed, but this formula gives us more information about how the atoms are arranged in relation to each other.

The slightly tricky part of writing chemical formulae comes when you have to work out the ratio in which the elements combine. For example, you may know that sodium (Na) and chlorine (Cl) react to form sodium chloride, but how do you know that in each molecule of sodium chloride there is only *one* atom of sodium for every *one* atom of chlorine? It all comes down to the **valency** of an atom or group of atoms. Valency is the number of bonds that an element can form with another element. Working out the chemical formulae of chemical compounds using their valency, will be covered in Grade 11. For now, we will use formulae that you already know.

8.4 Balancing chemical equations

8.4.1 The law of conservation of mass

In order to balance a chemical equation, it is important to understand the law of conservation of mass.

Definition 8.1: The law of conservation of mass

The mass of a closed system of substances will remain constant, regardless of the processes acting inside the system. Matter can change form, but cannot be created or destroyed. For any chemical process in a closed system, the mass of the reactants must equal the mass of the products.

In a chemical equation then, the **mass** of the reactants must be equal to the mass of the products. In order to make sure that this is the case, the number of **atoms** of each element in the reactants must be equal to the number of atoms of those same elements in the products. Some examples are shown below:

Example 1:

$$Fe + S \rightarrow FeS \qquad (8.1)$$

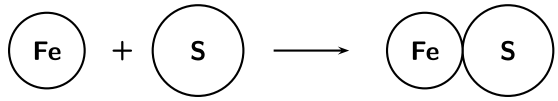


Figure 8.1

Reactants

Atomic mass of reactants $= 55,8\,u + 32,1\,u = 87,9\,u$

Number of atoms of each element in the reactants: $(1 \times Fe)$ and $(1 \times S)$

Products

Atomic mass of products $= 55,8\,u + 32,1\,u = 87,9\,u$

Number of atoms of each element in the products: $(1 \times Fe)$ and $(1 \times S)$

Since the number of atoms of each element is the same in the reactants and in the products, we say that the equation is **balanced**.

Example 2:

$$H_2 + O_2 \rightarrow H_2O \tag{8.2}$$

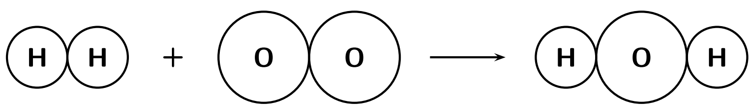


Figure 8.2

Reactants

Atomic mass of reactants $= (1 + 1) + (16 + 16) = 34\,u$

Number of atoms of each element in the reactants: $(2 \times H)$ and $(2 \times O)$

Product

Atomic mass of product $= (1 + 1 + 16) = 18\,u$

Number of atoms of each element in the product: $(2 \times H)$ and $(1 \times O)$

Since the total atomic mass of the reactants and the products is not the same and since there are more oxygen atoms in the reactants than there are in the product, the equation is **not balanced**.

Example 3:

$$NaOH + HCl \rightarrow NaCl + H_2O \tag{8.3}$$

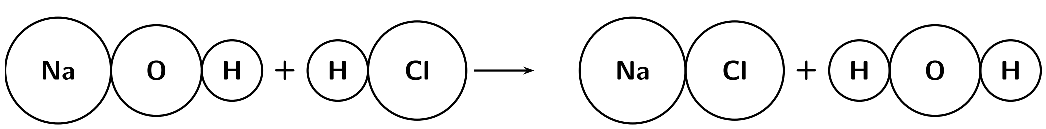


Figure 8.3

Reactants
Atomic mass of reactants $= (23 + 6 + 1) + (1 + 35,4) = 76,4\,u$
Number of atoms of each element in the reactants: $(1 \times Na) + (1 \times O) + (2 \times H) + (1 \times Cl)$
Products
Atomic mass of products $= (23 + 35,4) + (1 + 1 + 16) = 76,4\,u$
Number of atoms of each element in the products: $(1 \times Na) + (1 \times O) + (2 \times H) + (1 \times Cl)$

Since the number of atoms of each element is the same in the reactants and in the products, we say that the equation is **balanced**.

We now need to find a way to balance those equations that are not balanced so that the number of atoms of each element in the reactants is the same as that for the products. This can be done by changing the **coefficients** of the molecules until the atoms on each side of the arrow are balanced. You will see later that these coefficients tell us something about the **mole ratio** in which substances react. They also tell us about the volume relationship between gases in the reactants and products.

TIP: Coefficients

Remember that if you put a number in front of a molecule, that number applies to the *whole* molecule. For example, if you write $2H_2O$, this means that there are 2 molecules of water. In other words, there are 4 hydrogen atoms and 2 oxygen atoms. If we write 3HCl, this means that there are 3 molecules of HCl. In other words there are 3 hydrogen atoms and 3 chlorine atoms in total. In the first example, 2 is the coefficient and in the second example, 3 is the coefficient.

8.4.2 Activity: Balancing chemical equations

You will need: coloured balls (or marbles), prestik, a sheet of paper and coloured pens.

We will try to balance the following equation:

$$Al + O_2 \rightarrow Al_2O_3 \qquad (8.4)$$

Take 1 ball of one colour. This represents a molecule of Al. Take two balls of another colour and stick them together. This represents a molecule of O_2. Place these molecules on your left. Now take two balls of one colour and three balls of another colour to form a molecule of Al_2O_3. Place these molecules on your right. On a piece of paper draw coloured circles to represent the balls. Draw a line down the center of the paper to represent the molecules on the left and on the right.

Count the number of balls on the left and the number on the right. Do you have the same number of each colour on both sides? If not the equation is not balanced. How many balls will you have to add to each side to make the number of balls the same? How would you add these balls?

You should find that you need 4 balls of one colour for Al and 3 pairs of balls of another colour (i.e. 6 balls in total) for O_2 on the left side. On the right side you should find that you need 2 clusters of balls for Al_2O_3. We say that the balanced equation is:

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \qquad (8.5)$$

Repeat this process for the following reactions:

- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
- $2H_2 + O_2 \rightarrow 2H_2O$
- $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

8.4.3 Steps to balance a chemical equation

When balancing a chemical equation, there are a number of steps that need to be followed.

- STEP 1: Identify the reactants and the products in the reaction and write their chemical formulae.
- STEP 2: Write the equation by putting the reactants on the left of the arrow and the products on the right.
- STEP 3: Count the number of atoms of each element in the reactants and the number of atoms of each element in the products.
- STEP 4: If the equation is not balanced, change the coefficients of the molecules until the number of atoms of each element on either side of the equation balance.
- STEP 5: Check that the atoms are in fact balanced.
- STEP 6 (we will look at this a little later): Add any extra details to the equation e.g. phase.

Exercise 8.1: Balancing chemical equations 1 ***(Solution on p. 130.)***
Balance the following equation:

$$Mg + HCl \rightarrow MgCl_2 + H_2 \tag{8.6}$$

Exercise 8.2: Balancing chemical equations 2 ***(Solution on p. 130.)***
Balance the following equation:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O \tag{8.7}$$

Exercise 8.3: Balancing chemical equations 3 ***(Solution on p. 130.)***
Nitrogen gas reacts with hydrogen gas to form ammonia. Write a balanced chemical equation for this reaction.

Exercise 8.4: Balancing chemical equations 4 ***(Solution on p. 131.)***
In our bodies, sugar ($C_6H_{12}O_6$) reacts with the oxygen we breathe in to produce carbon dioxide, water and energy. Write the balanced equation for this reaction.

This simulation allows you to practice balancing simple equations.

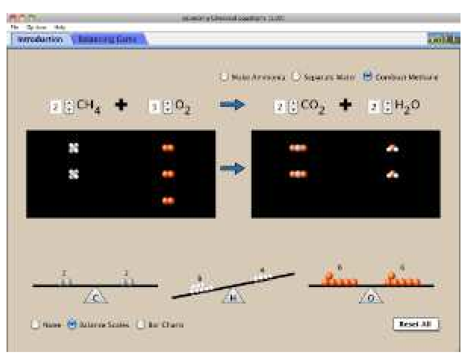


Figure 8.4

run demo[2]

[2] http://phet.colorado.edu/sims/balancing-chemical-equations/balancing-chemical-equations_en.jnlp

8.4.3.1 Balancing simple chemical equations

Balance the following equations:

1. Hydrogen fuel cells are extremely important in the development of alternative energy sources. Many of these cells work by reacting hydrogen and oxygen gases together to form water, a reaction which also produces electricity. Balance the following equation: $H_2 (g) + O_2 (g) \rightarrow H_2O (l)$Click here for the solution[3]
2. The synthesis of ammonia (NH_3), made famous by the German chemist Fritz Haber in the early 20th century, is one of the most important reactions in the chemical industry. Balance the following equation used to produce ammonia: $N_2 (g) + H_2 (g) \rightarrow NH_3 (g)$ Click here for the solution[4]
3. $Mg + P_4 \rightarrow Mg_3P_2$ Click here for the solution[5]
4. $Ca + H_2O \rightarrow Ca (OH)_2 + H_2$Click here for the solution[6]
5. $CuCO_3 + H_2SO_4 \rightarrow CuSO_4 + H_2O + CO_2$ Click here for the solution[7]
6. $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + NaCl$Click here for the solution[8]
7. $C_{12}H_{22}O_{11} + O_2 \rightarrow H_2O + CO_2$ Click here for the solution[9]
8. Barium chloride reacts with sulphuric acid to produce barium sulphate and hydrochloric acid. Click here for the solution[10]
9. Ethane (C_2H_6) reacts with oxygen to form carbon dioxide and steam. Click here for the solution[11]
10. Ammonium carbonate is often used as a smelling salt. Balance the following reaction for the decomposition of ammonium carbonate: $(NH_4)_2CO_3 (s) \rightarrow NH_3 (aq) + CO_2 (g) + H_2O (l)$ Click here for the solution[12]

8.5 State symbols and other information

The state (phase) of the compounds can be expressed in the chemical equation. This is done by placing the correct label on the right hand side of the formula. There are only four labels that can be used:

1. (g) for gaseous compounds
2. (l) for liquids
3. (s) for solid compounds
4. (aq) for an aqueous (water) solution

Occasionally, a catalyst is added to the reaction. A catalyst is a substance that speeds up the reaction without undergoing any change to itself. In a chemical equation, this is shown by using the symbol of the catalyst above the arrow in the equation.

To show that heat is needed for a reaction, a Greek delta (Δ) is placed above the arrow in the same way as the catalyst.

> TIP: You may remember from Physical and chemical change (Chapter 7) that energy cannot be created or destroyed during a chemical reaction but it may change form. In an exothermic reaction, ΔH is less than zero and in an endothermic reaction, ΔH is greater than zero. This value is often written at the end of a chemical equation.

[3] http://www.fhsst.org/lOg
[4] http://www.fhsst.org/lO4
[5] http://www.fhsst.org/lO2
[6] http://www.fhsst.org/lOT
[7] http://www.fhsst.org/lOb
[8] http://www.fhsst.org/lOj
[9] http://www.fhsst.org/lOD
[10] http://www.fhsst.org/lOW
[11] http://www.fhsst.org/lOZ
[12] http://www.fhsst.org/lOB

Exercise 8.5: Balancing chemical equations 5 *(Solution on p. 131.)*
Solid zinc metal reacts with aqueous hydrochloric acid to form an aqueous solution of zinc chloride ($ZnCl_2$)and hydrogen gas. Write a balanced equation for this reaction.

Exercise 8.6: Balancing chemical equations 5 (advanced) *(Solution on p. 132.)*
Balance the following equation:

$$(NH_4)_2SO_4 + NaOH \rightarrow NH_3 + H_2O + Na_2SO_4 \qquad (8.8)$$

In this example, the first two steps are not necessary because the reactants and products have already been given.

The following video explains some of the concepts of balancing chemical equations.

Khan Academy video on balancing chemical equations

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/RnGu3xO2h74&rel=0&hl=en_US&feature=player_embedded&version=3;rel=0>

Figure 8.5

8.5.1 Balancing more advanced chemical equations

Write balanced equations for each of the following reactions:

1. $Al_2O_3(s) + H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$
2. $Mg(OH)_2(aq) + HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$
3. Lead (II) nitrate solution reacts with potassium iodide solution.
4. When heated, aluminium reacts with solid copper oxide to produce copper metal and aluminium oxide (Al_2O_3).
5. When calcium chloride solution is mixed with silver nitrate solution, a white precipitate (solid) of silver chloride appears. Calcium nitrate ($Ca(NO_3)_2$) is also produced in the solution. Click here for the solution[13]

Balanced equations are very important in chemistry. It is only by working with the balanced equations that chemists can perform many different calculations that tell them what quantity of something reacts. In a later chapter we will learn how to work with some of these calculations. We can interpret balanced chemical equations in terms of the conservation of matter, the conservation of mass or the conservation of energy.

This media object is a Flash object. Please view or download it at
<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=reactions-100511040544-phpapp02&stripped_title=reactions-4046986>

Figure 8.6

[13] http://www.fhsst.org/lOK

8.6 Summary

- A **chemical equation** uses symbols to describe a chemical reaction.
- In a chemical equation, **reactants** are written on the left hand side of the equation and the **products** on the right. The arrow is used to show the direction of the reaction.
- When representing chemical change, it is important to be able to write the **chemical formula** of a compound.
- In any chemical reaction, the **law of conservation of mass** applies. This means that the total atomic mass of the reactants must be the same as the total atomic mass of the products. This also means that the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the product.
- If the number of atoms of each element in the reactants is the same as the number of atoms of each element in the product, then the equation is **balanced**.
- If the number of atoms of each element in the reactants is not the same as the number of atoms of each element in the product, then the equation is **not balanced**.
- In order to balance an equation, **coefficients** can be placed in front of the reactants and products until the number of atoms of each element is the same on both sides of the equation.

8.6.1 End of chapter exercises

1. Propane is a fuel that is commonly used as a heat source for engines and homes. Balance the following equation for the combustion of propane: $C_3H_8\left(l\right) + O_2\left(g\right) \rightarrow CO_2\left(g\right) + H_2O\left(l\right)$ Click here for the solution[14]
2. Aspartame, an artificial sweetener, has the formula $C_{14}H_{18}N_2O_2$. Write the balanced equation for its combustion (reaction with O_2) to form CO_2 gas, liquid H_2O, and N_2 gas. Click here for the solution[15]
3. $Fe_2\left(\mathrm{SO}_4\right)_3 + \mathrm{K}\left(\mathrm{SCN}\right) \rightarrow K_3\mathrm{Fe}\left(\mathrm{SCN}\right)_6 + K_2\mathrm{SO}_4$ Click here for the solution[16]
4. Chemical weapons were banned by the Geneva Protocol in 1925. According to this protocol, all chemicals that release suffocating and poisonous gases are not to be used as weapons. White phosphorus, a very reactive allotrope of phosphorus, was recently used during a military attack. Phosphorus burns vigorously in oxygen. Many people got severe burns and some died as a result. The equation for this spontaneous reaction is: $P_4\left(s\right) + O_2\left(g\right) \rightarrow P_2O_5\left(s\right)$
   a. Balance the chemical equation.
   b. Prove that the law of conservation of mass is obeyed during this chemical reaction.
   c. Name the product formed during this reaction.
   d. Classify the reaction as endothermic or exothermic. Give a reason for your answer.
   e. Classify the reaction as a synthesis or decomposition reaction. Give a reason for your answer. Click here for the solution[17]

   (DoE Exemplar Paper 2 2007)
5. Mixing bleach (NaOCl) and ammonia (two common household cleaners) is very dangerous. When these two substances are mixed they produce toxic chloaramine (NH_2Cl) fumes. Balance the following equations that occur when bleach and ammonia are mixed:
   a. $\mathrm{NaOCl}\left(\mathrm{aq}\right) + \mathrm{NH}_3\left(\mathrm{aq}\right) \rightarrow \mathrm{NaONH}_3\left(\mathrm{aq}\right) + Cl_2\left(g\right)$
   b. If there is more bleach than ammonia the following may occur: $\mathrm{NaOCl} + \mathrm{NH}_3 \rightarrow \mathrm{NaOH} + \mathrm{NCl}_3$ Nitrogen trichloride (NCl_3) is highly explosive.

[14] http://www.fhsst.org/lOk
[15] http://www.fhsst.org/lO0
[16] http://www.fhsst.org/lO8
[17] http://www.fhsst.org/lO9

c. If there is more ammonia than bleach the following may occur: $NH_3 + NaOCl \rightarrow NaOH + NH_2Cl$
These two products then react with ammonia as follows:
$NH_3 + NH_2Cl + NaOH \rightarrow N_2H_4 + NaCl + H_2O$
One last reaction occurs to stabilise the hydrazine and chloramine: $NH_2Cl + N_2H_4 \rightarrow NH_4Cl + N_2$
This reaction is highly exothermic and will explode.

Click here for the solution[18]

6. Balance the following chemical equation: $N_2O_5 \rightarrow NO_2 + O_2$
Click here for the solution[19]
7. Sulphur can be produced by the Claus process. This two-step process involves reacting hydrogen sulphide with oxygen and then reacting the sulphur dioxide that is produced with more hydrogen sulphide. The equations for these two reactions are:

$$\begin{aligned} H_2S + O_2 &\rightarrow SO_2 + H_2O \\ H_2S + SO_2 &\rightarrow S + H_2O \end{aligned} \tag{8.9}$$

Balance these two equations.
Click here for the solution[20]

[18]http://www.fhsst.org/l2S
[19]http://www.fhsst.org/l2h/
[20]http://www.fhsst.org/lTq

Solutions to Exercises in Chapter 8

Solution to Exercise 8.1 (p. 125)

Step 1. Reactants: $Mg = 1\,\text{atom}$; $H = 1\,\text{atom}$ and $Cl = 1\,\text{atom}$
Products: $Mg = 1\,\text{atom}$; $H = 2\,\text{atoms}$ and $Cl = 2\,\text{atoms}$

Step 2. The equation is not balanced since there are 2 chlorine atoms in the product and only 1 in the reactants. If we add a coefficient of 2 to the HCl to increase the number of H and Cl atoms in the reactants, the equation will look like this:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \quad (8.10)$$

Step 3. If we count the atoms on each side of the equation, we find the following:
Reactants: $Mg = 1\,\text{atom}$; $H = 2\,\text{atom}$ and $Cl = 2\,\text{atom}$
Products: $Mg = 1\,\text{atom}$; $H = 2\,\text{atom}$ and $Cl = 2\,\text{atom}$
The equation is balanced. The final equation is:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \quad (8.11)$$

Solution to Exercise 8.2 (p. 125)

Step 1. Reactants: $C = 1$; $H = 4$ and $O = 2$
Products: $C = 1$; $H = 2$ and $O = 3$

Step 2. If we add a coefficient of 2 to H_2O, then the number of hydrogen atoms in the reactants will be 4, which is the same as for the reactants. The equation will be:

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O \quad (8.12)$$

Step 3. Reactants: $C = 1$; $H = 4$ and $O = 2$
Products: $C = 1$; $H = 4$ and $O = 4$
You will see that, although the number of *hydrogen* atoms now balances, there are more oxygen atoms in the products. You now need to repeat the previous step. If we put a coefficient of 2 in front of O_2, then we will increase the number of oxygen atoms in the reactants by 2. The new equation is:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \quad (8.13)$$

When we check the number of atoms again, we find that the number of atoms of each element in the reactants is the same as the number in the products. The equation is now balanced.

Solution to Exercise 8.3 (p. 125)

Step 1. The reactants are nitrogen (N_2) and hydrogen (H_2) and the product is ammonia (NH_3).

Step 2. The equation is as follows:

$$N_2 + H_2 \rightarrow NH_3 \quad (8.14)$$

Step 3. Reactants: $N = 2$ and $H = 2$
Products: $N = 1$ and $H = 3$

Step 4. In order to balance the number of nitrogen atoms, we could rewrite the equation as:

$$N_2 + H_2 \rightarrow 2NH_3 \quad (8.15)$$

Step 5. In the above equation, the nitrogen atoms now balance, but the hydrogen atoms don't (there are 2 hydrogen atoms in the reactants and 6 in the product). If we put a coefficient of 3 in front of the hydrogen (H_2), then the hydrogen atoms and the nitrogen atoms balance. The final equation is:

$$N_2 + 3H_2 \rightarrow 2NH_3 \tag{8.16}$$

Solution to Exercise 8.4 (p. 125)

Step 1. Reactants: sugar ($C_6H_{12}O_6$) and oxygen (O_2)
Products: carbon dioxide (CO_2) and water (H_2O)

Step 2.

$$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O \tag{8.17}$$

Step 3. Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 1$; $H = 2$ and $O = 3$

Step 4. It is easier to start with carbon as it only appears once on each side. If we add a 6 in front of CO_2, the equation looks like this:

$$C_6H_{12}O_6 + O_2 \rightarrow 6CO_2 + H_2O \tag{8.18}$$

Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 6$; $H = 2$ and $O = 13$

Step 5. Let's try to get the number of hydrogens the same this time.

$$C_6H_{12}O_6 + O_2 \rightarrow 6CO_2 + 6H_2O \tag{8.19}$$

Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 6$; $H = 12$ and $O = 18$

Step 6.

$$C_6H_{12}O_6 + 12O_2 \rightarrow 6CO_2 + 6H_2O \tag{8.20}$$

Reactants: $C = 6$; $H = 12$ and $O = 18$
Products: $C = 6$; $H = 12$ and $O = 18$

Solution to Exercise 8.5 (p. 127)

Step 1. The reactants are zinc (Zn) and hydrochloric acid (HCl). The products are zinc chloride ($\mathrm{ZnCl_2}$) and hydrogen (H_2).

Step 2.

$$Zn + HCl \rightarrow ZnCl_2 + H_2 \tag{8.21}$$

Step 3. You will notice that the zinc atoms balance but the chlorine and hydrogen atoms don't. Since there are two chlorine atoms on the right and only one on the left, we will give HCl a coefficient of 2 so that there will be two chlorine atoms on each side of the equation.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2 \tag{8.22}$$

Step 4. When you look at the equation again, you will see that all the atoms are now balanced.

Step 5. In the initial description, you were told that zinc was a metal, hydrochloric acid and zinc chloride were in aqueous solutions and hydrogen was a gas.

$$Zn\,(s) + 2HCl\,(aq) \rightarrow ZnCl_2\,(aq) + H_2\,(g) \tag{8.23}$$

Solution to Exercise 8.6 (p. 127)

Step 1. With a complex equation, it is always best to start with atoms that appear only once on each side i.e. Na, N and S atoms. Since the S atoms already balance, we will start with Na and N atoms. There are two Na atoms on the right and one on the left. We will add a second Na atom by giving $NaOH$ a coefficient of two. There are two N atoms on the left and one on the right. To balance the N atoms, $\mathrm{NH_3}$ will be given a coefficient of two. The equation now looks as follows:

$$(NH_4)_2\mathrm{SO}_4 + 2NaOH \rightarrow 2NH_3 + H_2O + Na_2\mathrm{SO}_4 \quad (8.24)$$

Step 2. N, Na and S atoms balance, but O and H atoms do not. There are six O atoms and ten H atoms on the left, and five O atoms and eight H atoms on the right. We need to add one O atom and two H atoms on the right to balance the equation. This is done by adding another H_2O molecule on the right hand side. We now need to check the equation again:

$$(NH_4)_2\mathrm{SO}_4 + 2NaOH \rightarrow 2NH_3 + 2H_2O + Na_2\mathrm{SO}_4 \quad (8.25)$$

The equation is now balanced.

Chapter 9

Reactions in Aqueous Solutions[1]

9.1 Introduction

Many reactions in chemistry and all biological reactions (reactions in living systems) take place in water. We say that these reactions take place in aqueous solution. Water has many unique properties and is plentiful on Earth. For these reasons reactions in aqueous solutions occur frequently. In this chapter we will look at some of these reactions in detail. Almost all the reactions that occur in aqueous solutions involve ions. We will look at three main types of reactions that occur in aqueous solutions, namely precipitation reactions, acid-base reactions and redox reactions. Before we can learn about the types of reactions, we need to first look at ions in aqueous solutions and electrical conductivity.

9.2 Ions in aqueous solution

Water is seldom pure. Because of the structure of the water molecule, substances can dissolve easily in it. This is very important because if water wasn't able to do this, life would not be able to survive. In rivers and the oceans for example, dissolved oxygen means that organisms (such as fish) are still able to respire (breathe). For plants, dissolved nutrients are also available. In the human body, water is able to carry dissolved substances from one part of the body to another.

Many of the substances that dissolve are *ionic* and when they dissolve they form ions in solution. We are going to look at how water is able to dissolve ionic compounds, how these ions maintain a balance in the human body, how they affect water hardness and how they cause acid rain.

9.2.1 Dissociation in water

Water is a **polar molecule** (Figure 9.1). This means that one part of the molecule has a slightly positive charge (positive pole) and the other part has a slightly negative charge (negative pole).

[1]This content is available online at <http://cnx.org/content/m38136/1.6/>.

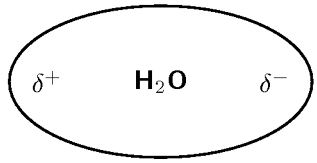


Figure 9.1: Water is a polar molecule

It is the polar nature of water that allows ionic compounds to dissolve in it. In the case of sodium chloride ($NaCl$) for example, the positive sodium ions (Na^+) will be attracted to the negative pole of the water molecule, while the negative chloride ions (Cl^-) will be attracted to the positive pole of the water molecule. In the process, the ionic bonds between the sodium and chloride ions are weakened and the water molecules are able to work their way between the individual ions, surrounding them and slowly dissolving the compound. This process is called **dissociation**. A simplified representation of this is shown in Figure 9.2. We say that dissolution of a substance has occurred when a substance dissociates or dissolves.

Definition 9.1: Dissociation

Dissociation in chemistry and biochemistry is a general process in which ionic compounds separate or split into smaller molecules or ions, usually in a reversible manner.

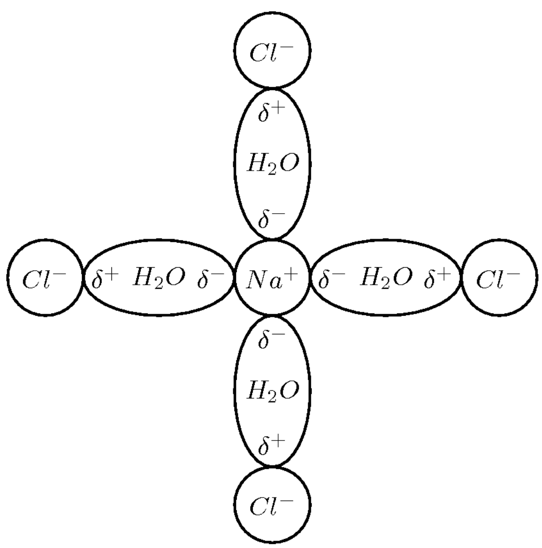


Figure 9.2: Sodium chloride dissolves in water

The dissolution of sodium chloride can be represented by the following equation:

$NaCl\,(s) \rightarrow Na^+(aq) + Cl^-(aq)$

The symbols **s** (solid), **l** (liquid), **g** (gas) and **aq** (material is dissolved in water) are written after the chemical formula to show the state or phase of the material. The dissolution of potassium sulphate into potassium and sulphate ions is shown below as another example:

$K_2SO_4(s) \rightarrow 2K^+(aq) + SO_4^{2-}(aq)$

Remember that **molecular** substances (e.g. covalent compounds) may also dissolve, but most will not form ions. One example is sugar.

$C_6H_{12}O_6(s) \rightleftharpoons C_6H_{12}O_6(aq)$

There are exceptions to this and some molecular substances *will* form ions when they dissolve. Hydrogen chloride for example can ionise to form hydrogen and chloride ions.

$HCl\,(g) \rightarrow H^+(aq) + Cl^-(aq)$

> NOTE: The ability of ionic compounds to dissolve in water is extremely important in the human body! The body is made up of *cells*, each of which is surrounded by a *membrane*. Dissolved ions are found inside and outside of body cells in different concentrations. Some of these ions are positive (e.g. Mg^{2+}) and some are negative (e.g. Cl^-). If there is a difference in the charge that is inside and outside the cell, then there is a *potential difference* across the cell membrane. This is called the **membrane potential** of the cell. The membrane potential acts like a battery and affects the movement of all charged substances across the membrane. Membrane potentials play a role in muscle functioning, digestion, excretion and in maintaining blood pH to name just a few. The movement of ions across the membrane can also be converted into an electric signal that can be transferred along *neurons* (nerve cells), which control body processes. If ionic substances were not able to dissociate in water, then none of these processes would be possible! It is also important to realise that our bodies can *lose* ions such as Na^+, K^+, Ca^{2+}, Mg^{2+}, and Cl^-, for example when we sweat during exercise. Sports drinks such as Lucozade and Powerade are designed to replace these lost ions so that the body's normal functioning is not affected.

Exercise 9.1: Dissociation in water ***(Solution on p. 149.)***

Write a balanced equation to show how silver nitrate ($AgNO_3$) dissociates in water.

9.2.1.1 Ions in solution

1. For each of the following, say whether the substance is ionic or molecular.
   a. potassium nitrate (KNO_3)
   b. ethanol (C_2H_5OH)
   c. sucrose (a type of sugar) ($C_{12}H_{22}O_{11}$)
   d. sodium bromide ($NaBr$)

   Click here for the solution[2]
2. Write a balanced equation to show how each of the following ionic compounds dissociate in water.
   a. sodium sulphate (Na_2SO_4)
   b. potassium bromide (KBr)
   c. potassium permanganate ($KMnO_4$)
   d. sodium phosphate (Na_3PO_4)

   Click here for the solution[3]

[2]See the file at <http://cnx.org/content/m38136/latest/http://www.fhsst.org/l33>
[3]See the file at <http://cnx.org/content/m38136/latest/http://www.fhsst.org/l3g>

9.2.2 Ions and water hardness

This section is not examinable and is included as an example of ions in aqueous solution.

Definition 9.2: Water hardness
Water hardness is a measure of the mineral content of water. Minerals are substances such as calcite, quartz and mica that occur naturally as a result of geological processes.

Hard water is water that has a high mineral content. Water that has a low mineral content is known as **soft water**. If water has a high mineral content, it usually contains high levels of metal ions, mainly calcium (Ca) and magnesium (Mg). The calcium enters the water from either $CaCO_3$ (limestone or chalk) or from mineral deposits of $CaSO_4$. The main source of magnesium is a sedimentary rock called dolomite, $CaMg\,(CO_3)_2$. Hard water may also contain other metals as well as bicarbonates and sulphates.

NOTE: The simplest way to check whether water is hard or soft is to use the lather/froth test. If the water is very soft, soap will lather more easily when it is rubbed against the skin. With hard water this won't happen. Toothpaste will also not froth well in hard water.

A **water softener** works on the principle of **ion exchange**. Hard water passes through a media bed, usually made of resin beads that are supersaturated with sodium. As the water passes through the beads, the hardness minerals (e.g. calcium and magnesium) attach themselves to the beads. The sodium that was originally on the beads is released into the water. When the resin becomes saturated with calcium and magnesium, it must be recharged. A salt solution is passed through the resin. The sodium replaces the calcium and magnesium and these ions are released into the waste water and discharged.

9.3 Acid rain

This section is not examinable and is included as an example of ions in aqueous solution.

The acidity of rainwater comes from the natural presence of three substances (CO_2, NO, and SO_2) in the lowest layer of the atmosphere. These gases are able to dissolve in water and therefore make rain more acidic than it would otherwise be. Of these gases, carbon dioxide (CO_2) has the highest concentration and therefore contributes the most to the natural acidity of rainwater. We will look at each of these gases in turn.

Definition 9.3: Acid rain
Acid rain refers to the deposition of acidic components in rain, snow and dew. Acid rain occurs when sulphur dioxide and nitrogen oxides are emitted into the atmosphere, undergo chemical transformations and are absorbed by water droplets in clouds. The droplets then fall to earth as rain, snow, mist, dry dust, hail, or sleet. This increases the acidity of the soil and affects the chemical balance of lakes and streams.

1. **Carbon dioxide** Carbon dioxide reacts with water in the atmosphere to form **carbonic acid** (H_2CO_3).
$$CO_2 + H_2O \rightarrow H_2CO_3 \tag{9.1}$$
The carbonic acid dissociates to form hydrogen and hydrogen carbonate ions. It is the presence of hydrogen ions that lowers the pH of the solution making the rain acidic.
$$H_2CO_3 \rightarrow H^+ + \mathrm{HCO}_3^- \tag{9.2}$$
2. **Nitric oxide** Nitric oxide (NO) also contributes to the natural acidity of rainwater and is formed during lightning storms when nitrogen and oxygen react. In air, NO is oxidised to form nitrogen dioxide (NO_2). It is the nitrogen dioxide which then reacts with water in the atmosphere to form **nitric acid** (HNO_3).
$$3NO_2\,(g) + H_2O\,(l) \rightarrow 2HNO_3\,(aq) + NO\,(g) \tag{9.3}$$

The nitric acid dissociates in water to produce hydrogen ions and nitrate ions. This again lowers the pH of the solution making it acidic.

$$HNO_3 \rightarrow H^+ + \mathrm{NO}_3^- \quad (9.4)$$

3. **Sulphur dioxide** Sulphur dioxide in the atmosphere first reacts with oxygen to form sulphur trioxide, before reacting with water to form **sulphuric acid**.

$$2SO_2 + O_2 \rightarrow 2SO_3 \quad (9.5)$$

$$SO_3 + H_2O \rightarrow H_2SO_4 \quad (9.6)$$

Sulphuric acid dissociates in a similar way to the previous reactions.

$$H_2SO_4 \rightarrow HSO_4^- + H^+ \quad (9.7)$$

Although these reactions do take place naturally, human activities can greatly increase the concentration of these gases in the atmosphere, so that rain becomes far more acidic than it would otherwise be. The burning of fossil fuels in industries, vehicles etc is one of the biggest culprits. If the acidity of the rain drops to below 5, it is referred to as **acid rain**.

Acid rain can have a very damaging effect on the environment. In rivers, dams and lakes, increased acidity can mean that some species of animals and plants will not survive. Acid rain can also degrade soil minerals, producing metal ions that are washed into water systems. Some of these ions may be toxic e.g. Al^{3+}. From an economic perspective, altered soil pH can drastically affect agricultural productivity.

Acid rain can also affect buildings and monuments, many of which are made from marble and limestone. A chemical reaction takes place between $CaCO_3$ (limestone) and sulphuric acid to produce aqueous ions which can be easily washed away. The same reaction can occur in the lithosphere where limestone rocks are present e.g. limestone caves can be eroded by acidic rainwater.

$$H_2SO_4 + CaCO_3 \rightarrow CaSO_4 \cdot H_2O + CO_2 \quad (9.8)$$

9.3.1 Investigation : Acid rain

You are going to test the effect of 'acid rain' on a number of substances.
Materials needed:

samples of chalk, marble, zinc, iron, lead, dilute sulphuric acid, test tubes, beaker, glass dropper
Method:

1. Place a small sample of each of the following substances in a separate test tube: chalk, marble, zinc, iron and lead
2. To each test tube, add a few drops of dilute sulphuric acid.
3. Observe what happens and record your results.

Discussion questions:

- In which of the test tubes did reactions take place? What happened to the sample substances?
- What do your results tell you about the effect that acid rain could have on each of the following: buildings, soils, rocks and geology, water ecosystems?
- What precautions could be taken to reduce the potential impact of acid rain?

9.4 Electrolytes, ionisation and conductivity

Conductivity in aqueous solutions, is a measure of the ability of water to conduct an electric current. The more **ions** there are in the solution, the higher its conductivity.

Definition 9.4: Conductivity
Conductivity is a measure of a solution's ability to conduct an electric current.

9.4.1 Electrolytes

An **electrolyte** is a material that *increases* the conductivity of water when dissolved in it. Electrolytes can be further divided into **strong electrolytes** and **weak electrolytes**.

Definition 9.5: Electrolyte
An electrolyte is a substance that contains free ions and behaves as an electrically conductive medium. Because they generally consist of ions in solution, electrolytes are also known as ionic solutions.

1. **Strong electrolytes** A strong electrolyte is a material that ionises completely when it is dissolved in water:

$$AB\ (s,\ l,\ g) \rightarrow A^{+}(\mathrm{aq}) + B^{-}(\mathrm{aq}) \tag{9.9}$$

   This is a **chemical change** because the original compound has been split into its component ions and bonds have been broken. In a strong electrolyte, we say that the *extent of ionisation* is high. In other words, the original material dissociates completely so that there is a high concentration of ions in the solution. An example is a solution of potassium nitrate:

$$KNO_3(\mathrm{s}) \rightarrow K^{+}(\mathrm{aq}) + \mathrm{NO}_3^{-}(\mathrm{aq}) \tag{9.10}$$

2. **Weak electrolytes** A weak electrolyte is a material that goes into solution and will be surrounded by water molecules when it is added to water. However, not *all* of the molecules will dissociate into ions. The *extent of ionisation* of a weak electrolyte is low and therefore the concentration of ions in the solution is also low.

$$AB\,(s,l,g) \rightarrow AB\,(aq) \rightleftharpoons A^{+}\,(\mathrm{aq}) + B^{-}\,(\mathrm{aq}) \tag{9.11}$$

   The following example shows that in the final solution of a weak electrolyte, some of the original compound *plus* some dissolved ions are present.

$$C_2H_3O_2H\,(l) \rightarrow C_2H_3O_2H \rightleftharpoons C_2H_3O_2^{-}\,(\mathrm{aq}) + H^{+}\,(\mathrm{aq}) \tag{9.12}$$

9.4.2 Non-electrolytes

A **non-electrolyte** is a material that does not increase the conductivity of water when dissolved in it. The substance goes into solution and becomes surrounded by water molecules, so that the molecules of the chemical become separated from each other. However, although the substance does dissolve, it is not changed in any way and no chemical bonds are broken. The change is a **physical change**. In the oxygen example below, the reaction is shown to be reversible because oxygen is only partially soluble in water and comes out of solution very easily.

$$C_2H_5\mathrm{OH}\,(l) \rightarrow C_2H_5\mathrm{OH}\,(\mathrm{aq}) \tag{9.13}$$

$$O_2\,(g) \rightleftharpoons O_2\,(\mathrm{aq}) \tag{9.14}$$

9.4.3 Factors that affect the conductivity of water

The conductivity of water is therefore affected by the following factors:

- The **type of substance** that dissolves in water. Whether a material is a strong electrolyte (e.g. potassium nitrate, KNO_3), a weak electrolyte (e.g. acetate, CH_3COOH) or a non-electrolyte (e.g. sugar, alcohol, oil) will affect the conductivity of water because the concentration of ions in solution will be different in each case.
- The **concentration of ions** in solution. The higher the concentration of ions in solution, the higher its conductivity will be.
- **Temperature.** The warmer the solution, the higher the solubility of the material being dissolved and therefore the higher the conductivity as well.

9.4.3.1 Experiment : Electrical conductivity

Aim:

To investigate the electrical conductivities of different substances and solutions.
Apparatus:

Solid salt ($NaCl$) crystals; different liquids such as distilled water, tap water, seawater, benzene and alcohol; solutions of salts e.g. $NaCl$, KBr; a solution of an acid (e.g. HCl) and a solution of a base (e.g. $NaOH$); torch cells; ammeter; conducting wire, crocodile clips and 2 carbon rods.
Method:

Set up the experiment by connecting the circuit as shown in the diagram below. In the diagram, 'X' represents the substance or solution that you will be testing. When you are using the solid crystals, the crocodile clips can be attached directly to each end of the crystal. When you are using solutions, two carbon rods are placed into the liquid and the clips are attached to each of the rods. In each case, complete the circuit and allow the current to flow for about 30 seconds. Observe whether the ammeter shows a reading.

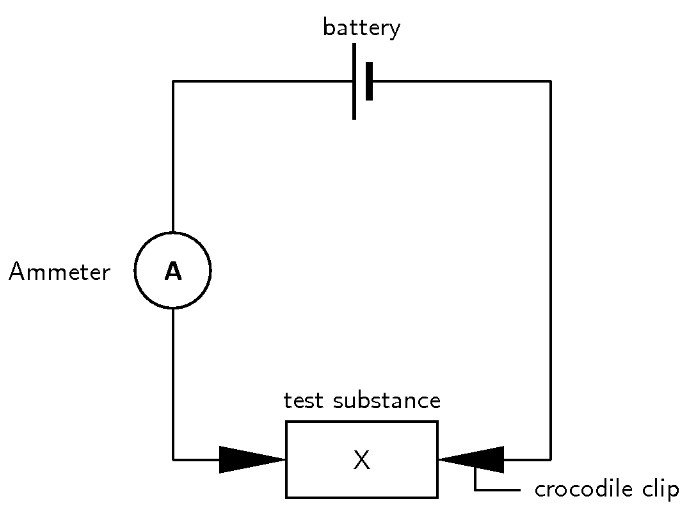


Figure 9.3

Results:

Record your observations in a table similar to the one below:

Test substance	Ammeter reading

Table 9.1

What do you notice? Can you explain these observations?

Remember that for electricity to flow, there needs to be a movement of charged particles e.g. ions. With the solid $NaCl$ crystals, there was no flow of electricity recorded on the ammeter. Although the solid is made up of ions, they are held together very tightly within the crystal lattice and therefore no current will flow. Distilled water, benzene and alcohol also don't conduct a current because they are *covalent compounds* and therefore do not contain ions.

The ammeter should have recorded a current when the salt solutions and the acid and base solutions were connected in the circuit. In solution, salts *dissociate* into their ions, so that these are free to move in the solution. Acids and bases behave in a similar way and dissociate to form hydronium and oxonium ions. Look at the following examples:

$$\mathrm{KBr} \rightarrow K^+ + \mathrm{Br}^- \qquad (9.15)$$

$$\mathrm{NaCl} \rightarrow \mathrm{Na}^+ + \mathrm{Cl}^- \qquad (9.16)$$

$$\mathrm{HCl} + H_2O \rightarrow H_3O^+ + \mathrm{Cl}^- \qquad (9.17)$$

$$\mathrm{NaOH} \rightarrow \mathrm{Na}^+ + \mathrm{OH}^- \qquad (9.18)$$

Conclusions:

Solutions that contain free-moving ions are able to conduct electricity because of the movement of charged particles. Solutions that do not contain free-moving ions do not conduct electricity.

> NOTE: Conductivity in streams and rivers is affected by the geology of the area where the water is flowing through. Streams that run through areas with granite bedrock tend to have lower conductivity because granite is made of materials that do not ionise when washed into the water. On the other hand, streams that run through areas with clay soils tend to have higher conductivity because the materials ionise when they are washed into the water. Pollution can also affect conductivity. A failing sewage system or an inflow of fertiliser runoff would raise the conductivity because of the presence of chloride, phosphate, and nitrate ions, while an oil spill (non-ionic) would lower the conductivity. It is very important that conductivity is kept within a certain acceptable range so that the organisms living in these water systems are able to survive.

9.5 Types of reactions

We will look at three types of reactions that occur in aqueous solutions. These are precipitation reactions, acid-base reactions and redox reactions. Precipitation and acid-base reactions are sometimes called ion exchange reactions. Redox reactions are electron transfer reactions. It is important to remember the difference between these two types of reactions. In ion exchange reactions ions are exchanged, in electron transfer reactions electrons are transferred. These terms will be explained further in the following sections.

Ion exchange reactions can be represented by:

$$AB\,(aq) + CD\,(aq) \rightarrow AD + CB \tag{9.19}$$

Either AD or CB may be a solid or a gas. When a solid forms this is known as a precipitation reaction. If a gas is formed then this may be called a gas forming reaction. Acid-base reactions are a special class of ion exchange reactions and we will look at them seperately.

The formation of a precipitate or a gas helps to make the reaction happen. We say that the reaction is driven by the formation of a precipitate or a gas. All chemical reactions will only take place if there is something to make them happen. For some reactions this happens easily and for others it is harder to make the reaction occur.

Definition 9.6: Ion exchange reaction
A type of reaction where the positive ions exchange their respective negative ions due to a driving force.

NOTE: Ion exchange reactions are used in ion exchange chromatography. Ion exchange chromatography is used to purify water and as a means of softening water. Often when chemists talk about ion exchange, they mean ion exchange chromatography.

9.5.1 Precipitation reactions

Sometimes, ions in solution may react with each other to form a new substance that is *insoluble*. This is called a **precipitate**.

Definition 9.7: Precipitate
A precipitate is the solid that forms in a solution during a chemical reaction.

9.5.1.1 Demonstration : The reaction of ions in solution

Apparatus and materials:

4 test tubes; copper(II) chloride solution; sodium carbonate solution; sodium sulphate solution

Figure 9.4

Method:

1. Prepare 2 test tubes with approximately 5 ml of dilute Cu(II) chloride solution in each
2. Prepare 1 test tube with 5 ml sodium carbonate solution

3. Prepare 1 test tube with 5 ml sodium sulphate solution
4. Carefully pour the sodium carbonate solution into one of the test tubes containing copper(II) chloride and observe what happens
5. Carefully pour the sodium sulphate solution into the second test tube containing copper(II) chloride and observe what happens

Results:

1. A light blue precipitate forms when sodium carbonate reacts with copper(II) chloride
2. No precipitate forms when sodium sulphate reacts with copper(II) chloride

It is important to understand what happened in the previous demonstration. We will look at what happens in each reaction, step by step.

1. **Reaction 1:** Sodium carbonate reacts with copper(II) chloride.
   When these compounds react, a number of ions are present in solution: Cu^{2+}, Cl^{-}, Na^{+} and CO_3^{2-}. Because there are lots of ions in solution, they will collide with each other and may recombine in different ways. The product that forms may be insoluble, in which case a precipitate will form, or the product will be soluble, in which case the ions will go back into solution. Let's see how the ions in this example could have combined with each other:

$$Cu^{2+} + CO_3^{2-} \rightarrow \mathrm{CuCO_3} \tag{9.20}$$

$$Cu^{2+} + 2Cl^{-} \rightarrow \mathrm{CuCl_2} \tag{9.21}$$

$$Na^{+} + Cl^{-} \rightarrow \mathrm{NaCl} \tag{9.22}$$

$$2Na^{+} + CO_3^{2-} \rightarrow Na_2CO_3 \tag{9.23}$$

   You can automatically exclude the reactions where sodium carbonate and copper(II) chloride are the products because these were the initial reactants. You also know that sodium chloride ($NaCl$) is soluble in water, so the remaining product (copper carbonate) must be the one that is insoluble. It is also possible to look up which salts are soluble and which are insoluble. If you do this, you will find that most carbonates are insoluble, therefore the precipitate that forms in this reaction must be $\mathrm{CuCO_3}$. The reaction that has taken place between the ions in solution is as follows:

$$2Na^{+} + CO_3^{2-} + Cu^{2+} + 2Cl^{-} \rightarrow CuCO_3 + 2Na^{+} + 2Cl^{-} \tag{9.24}$$

2. **Reaction 2:** Sodium sulphate reacts with copper(II) chloride.
   The ions that are present in solution are Cu^{2+}, Cl^{-}, Na^{+} and SO_4^{2-}. The ions collide with each other and may recombine in different ways. The possible combinations of the ions are as follows:

$$Cu^{2+} + SO_4^{2-} \rightarrow \mathrm{CuSO_4} \tag{9.25}$$

$$Cu^{2+} + 2Cl^{-} \rightarrow \mathrm{CuCl_2} \tag{9.26}$$

$$Na^{+} + Cl^{-} \rightarrow \mathrm{NaCl} \tag{9.27}$$

$$Na^{+} + SO_4^{2-} \rightarrow Na_2SO_4 \tag{9.28}$$

If we look up which of these salts are soluble and which are insoluble, we see that most chlorides and most sulphates are soluble. This is why no precipitate forms in this second reaction. Even when the ions recombine, they immediately separate and go back into solution. The reaction that has taken place between the ions in solution is as follows:

$$2Na^{+}+SO_4^{2-}+Cu^{2+}+2Cl^{-}\rightarrow 2Na^{+}+SO_4^{2-}+Cu^{2+}+2Cl^{-} \quad (9.29)$$

Table 9.2 shows some of the general rules about the solubility of different salts based on a number of investigations:

Salt	**Solubility**
Nitrates	All are soluble
Potassium, sodium and ammonium salts	All are soluble
Chlorides, bromides and iodides	All are soluble except silver, lead(II) and mercury(II) salts (e.g. silver chloride)
Sulphates	All are soluble except lead(II) sulphate, barium sulphate and calcium sulphate
Carbonates	All are insoluble except those of potassium, sodium and ammonium
Compounds with fluorine	Almost all are soluble except those of magnesium, calcium, strontium (II), barium (II) and lead (II)
Perchlorates and acetates	All are soluble
Chlorates	All are soluble except potassium chlorate
Metal hydroxides and oxides	Most are insoluble

Table 9.2: General rules for the solubility of salts

Salts of carbonates, phosphates, oxalates, chromates and sulphides are generally insoluble.

9.5.2 Testing for common anions in solution

It is also possible to carry out tests to determine which ions are present in a solution. You should try to do each of these tests in class.

9.5.2.1 Test for a chloride

Prepare a solution of the unknown salt using distilled water and add a small amount of **silver nitrate** solution. If a white precipitate forms, the salt is either a chloride or a carbonate.

$$Cl^{-}+Ag^{+}+\mathrm{NO}_3^{-}\rightarrow \mathrm{AgCl}+\mathrm{NO}_3^{-} \quad (9.30)$$

(AgCl is white precipitate)

$$\mathrm{CO}_3^{2-}+2Ag^{+}+2\mathrm{NO}_3^{-}\rightarrow Ag_2\mathrm{CO}_3+2\mathrm{NO}_3^{-} \quad (9.31)$$

($Ag_2\mathrm{CO}_3$ is white precipitate)

The next step is to treat the precipitate with a small amount of **concentrated nitric acid**. If the precipitate remains unchanged, then the salt is a chloride. If carbon dioxide is formed, and the precipitate disappears, the salt is a carbonate.

$\mathrm{AgCl}+\mathrm{HNO}_3\rightarrow$ (no reaction; precipitate is unchanged)

$\mathrm{Ag}_2\mathrm{CO}_3+2\mathrm{HNO}_3\rightarrow 2\mathrm{AgNO}_3+H_2O+\mathrm{CO}_2$ (precipitate disappears)

9.5.2.2 Test for a sulphate

Add a small amount of barium chloride solution to a solution of the test salt. If a white precipitate forms, the salt is either a sulphate or a carbonate.

$SO_4^{2-} + Ba^{2+} + Cl^- \rightarrow BaSO_4 + Cl^-$ ($BaSO_4$ is a white precipitate)

$CO_3^{2-} + Ba^{2+} + Cl^- \rightarrow BaCO_3 + Cl^-$ ($BaCO_3$ is a white precipitate)

If the precipitate is treated with nitric acid, it is possible to distinguish whether the salt is a sulphate or a carbonate (as in the test for a chloride).

$BaSO_4 + HNO_3 \rightarrow$ (no reaction; precipitate is unchanged)

$BaCO_3 + 2HNO_3 \rightarrow Ba(NO_3)_2 + H_2O + CO_2$ (precipitate disappears)

9.5.2.3 Test for a carbonate

If a sample of the dry salt is treated with a small amount of acid, the production of carbon dioxide is a positive test for a carbonate.

$\mathrm{Acid} + CO_3^{2-} \rightarrow CO_2$

If the gas is passed through limewater and the solution becomes milky, the gas is carbon dioxide.

$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$ (It is the insoluble $CaCO_3$ precipitate that makes the limewater go milky)

9.5.2.4 Test for bromides and iodides

As was the case with the chlorides, the bromides and iodides also form precipitates when they are reacted with silver nitrate. Silver chloride is a white precipitate, but the silver bromide and silver iodide precipitates are both pale yellow. To determine whether the precipitate is a bromide or an iodide, we use chlorine water and carbon tetrachloride (CCl_4).

Chlorine water frees bromine gas from the bromide and colours the carbon tetrachloride a reddish brown.

Chlorine water frees iodine gas from an iodide and colours the carbon tetrachloride purple.

9.5.2.4.1 Precipitation reactions and ions in solution

1. Silver nitrate ($AgNO_3$) reacts with potassium chloride (KCl) and a white precipitate is formed.
   a. Write a balanced equation for the reaction that takes place.
   b. What is the name of the insoluble salt that forms?
   c. Which of the salts in this reaction are soluble?

   Click here for the solution[4]
2. Barium chloride reacts with sulphuric acid to produce barium sulphate and hydrochloric acid.
   a. Write a balanced equation for the reaction that takes place.
   b. Does a precipitate form during the reaction?
   c. Describe a test that could be used to test for the presence of barium sulphate in the products.

   Click here for the solution[5]
3. A test tube contains a clear, colourless salt solution. A few drops of silver nitrate solution are added to the solution and a pale yellow precipitate forms. Which one of the following salts was dissolved in the original solution?
   a. NaI
   b. KCl
   c. K_2CO_3
   d. Na_2SO_4

[4] See the file at <http://cnx.org/content/m38136/latest/http://www.fhsst.org/l3c>
[5] See the file at <http://cnx.org/content/m38136/latest/http://www.fhsst.org/l3O>

(IEB Paper 2, 2005) Click here for the solution[6]

9.5.3 Other reactions in aqueous solutions

There are many types of reactions that can occur in aqueous solutions. In this section we will look at two of them: acid-base reactions and redox reactions. These reactions will be covered in more detail in Grade 11.

9.5.3.1 Acid-base reactions

Acid base reactions take place between acids and bases. In general, the products will be water and a salt (i.e. an ionic compound). An example of this type of reaction is:

$$NaOH\,(aq) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O\ \ (l) \qquad (9.32)$$

This is an special case of an ion exchange reaction since the sodium in the sodium hydroxide swaps places with the hydrogen in the hydrogen chloride forming sodium chloride. At the same time the hydroxide and the hydrogen combine to form water.

9.5.3.2 Redox reactions

Redox reactions involve the exchange of electrons. One ion loses electrons and becomes more positive, while the other ion gains electrons and becomes more negative. To decide if a redox reaction has occurred we look at the charge of the atoms, ions or molecules involved. If one of them has become more positive and the other one has become more negative then a redox reaction has occurred. For example, sodium metal is oxidised to form sodium oxide (and sometimes sodium peroxide as well). The balanced equation for this is:

$$4Na + O_2 \rightarrow 2Na_2O \qquad (9.33)$$

In the above reaction sodium and oxygen are both neutral and so have no charge. In the products however, the sodium atom has a charge of $+1$ and the oxygen atom has a charge of -2. This tells us that the sodium has lost electrons and the oxygen has gained electrons. Since one species has become more positive and one more negative we can conclude that a redox reaction has occurred. We could also say that electrons have been transferred from one species to the other. (In this case the electrons were transferred from the sodium to the oxygen).

9.5.3.2.1 Demonstration: Oxidation of sodium metal

You will need a bunsen burner, a small piece of sodium metal and a metal spatula. Light the bunsen burner. Place the sodium metal on the spatula. Place the sodium in the flame. When the reaction finishes, you should observe a white powder on the spatula. This is a mixture of sodium oxide (Na_2O) and sodium peroxide (Na_2O_2).

> WARNING: Sodium metal is very reactive. Sodium metal reacts vigourously with water and should never be placed in water. Be very careful when handling sodium metal.

[6]See the file at <http://cnx.org/content/m38136/latest/http://www.fhsst.org/l3x>

9.5.4 Experiment: Reaction types

Aim:

To use experiments to determine what type of reaction occurs.
Apparatus:

Soluble salts (e.g. potassium nitrate, ammonium chloride, sodium carbonate, silver nitrate, sodium bromide); hydrochloric acid (HCl); sodium hydroxide(NaOH); bromothymol blue; zinc metal; copper (II) sulphate; beakers; test-tubes
Method:

- For each of the salts, dissolve a small amount in water and observe what happens.
- Try dissolving pairs of salts (e.g. potassium nitrate and sodium carbonate) in water and observe what happens.
- Dissolve some sodium carbonate in hydrochloric acid and observe what happens.
- Carefully measure out $20cm^3$ of sodium hydroxide into a beaker.
- Add some bromothymol blue to the sodium hydroxide
- Carefully add a few drops of hydrochloric acid to the sodium hydroxide and swirl. Repeat until you notice the colour change.
- Place the zinc metal into the copper sulphate solution and observe what happens.

Results:

Answer the following questions:

- What did you observe when you dissolved each of the salts in water?
- What did you observe when you dissolved pairs of salts in the water?
- What did you observe when you dissolved sodium carbonate in hydrochloric acid?
- Why do you think we used bromothymol blue when mixing the hydrochloric acid and the sodium hydroxide? Think about the kind of reaction that occurred.
- What did you observe when you placed the zinc metal into the copper sulphate?
- Classify each reaction as either precipitation, gas forming, acid-base or redox.
- What makes each reaction happen (i.e. what is the driving force)? Is it the formation of a precipitate or something else?
- What criteria would you use to determine what kind of reaction occurs?
- Try to write balanced chemical equations for each reaction

Conclusion:

We can see how we can classify reactions by performing experiments.

In the experiment above, you should have seen how each reaction type differs from the others. For example, a gas forming reaction leads to bubbles in the solution, a precipitation reaction leads to a precipitate forming, an acid-base reaction can be seen by adding a suitable indicator and a redox reaction can be seen by one metal disappearing and a deposit forming in the solution.

9.6 Summary

- The **polar** nature of water means that **ionic compounds** dissociate easily in aqueous solution into their component ions.
- **Ions** in solution play a number of roles. In the human body for example, ions help to regulate the internal environment (e.g. controlling muscle function, regulating blood pH). Ions in solution also determine water hardness and pH.

- **Water hardness** is a measure of the mineral content of water. Hard water has a high mineral concentration and generally also a high concentration of metal ions e.g. calcium and magnesium. The opposite is true for soft water.
- **Conductivity** is a measure of a solution's ability to conduct an electric current.
- An **electrolyte** is a substance that contains free ions and is therefore able to conduct an electric current. Electrolytes can be divided into **strong** and **weak** electrolytes, based on the extent to which the substance ionises in solution.
- A **non-electrolyte** cannot conduct an electric current because it dooes not contain free ions.
- The **type of substance**, the **concentration of ions** and the **temperature** of the solution affect its conductivity.
- There are three main types of reactions that occur in aqueous solutions. These are precipitation reactions, acid-base reactions and redox reactions.
- Precipitation and acid-base reactions are sometimes known as ion exchange reactions. Ion exchange reactions also include gas forming reactions.
- A **precipitate** is formed when ions in solution react with each other to form an insoluble product. Solubility 'rules' help to identify the precipitate that has been formed.
- A number of tests can be used to identify whether certain **anions** are present in a solution.
- An acid-base reaction is one in which an acid reacts with a base to form a salt and water.
- A redox reaction is one in which electrons are transferred from one substance to another.

9.7 End of chapter exercises

1. Give one word for each of the following descriptions:
   a. the change in phase of water from a gas to a liquid
   b. a charged atom
   c. a term used to describe the mineral content of water
   d. a gas that forms sulphuric acid when it reacts with water

   Click here for the solution[7]
2. Match the information in column A with the information in column B by writing only the letter (A to I) next to the question number (1 to 7)

Column A	**Column B**
1. A polar molecule	A. H_2SO_4
2. molecular solution	B. $CaCO_3$
3. Mineral that increases water hardness	C. $NaOH$
4. Substance that increases the hydrogen ion concentration	D. salt water
5. A strong electrolyte	E. calcium
6. A white precipitate	F. carbon dioxide
7. A non-conductor of electricity	G. potassium nitrate
	H. sugar water
	I. O_2

Table 9.3

Click here for the solution[8]

[7] http://www.fhsst.org/l3a
[8] http://www.fhsst.org/l3C

3. For each of the following questions, choose the one correct answer from the list provided.
   a. Which one of the following substances does not conduct electricity in the solid phase but is an electrical conductor when molten?
      i. Cu
      ii. $PbBr_2$
      iii. H_2O
      iv. I_2

      (IEB Paper 2, 2003) Click here for the solution[9]
   b. The following substances are dissolved in water. Which one of the solutions is basic?
      i. sodium nitrate
      ii. calcium sulphate
      iii. ammonium chloride
      iv. potassium carbonate

      (IEB Paper 2, 2005) Click here for the solution[10]
4. Explain the difference between a weak electrolyte and a strong electrolyte. Give a generalised equation for each.
   Click here for the solution[11]
5. What factors affect the conductivity of water? How do each of these affect the conductivity?
   Click here for the solution[12]
6. For each of the following substances state whether they are molecular or ionic. If they are ionic, give a balanced reaction for the dissociation in water.
   a. Methane (CH_4)
   b. potassium bromide
   c. carbon dioxide
   d. hexane (C_6H_{14})
   e. lithium fluoride (LiF)
   f. magnesium chloride

   Click here for the solution[13]
7. Three test tubes (X, Y and Z) each contain a solution of an unknown potassium salt. The following observations were made during a practical investigation to identify the solutions in the test tubes: A: A white precipitate formed when silver nitrate ($AgNO_3$) was added to test tube Z. B: A white precipitate formed in test tubes X and Y when barium chloride ($BaCl_2$) was added. C: The precipitate in test tube X dissolved in hydrochloric acid (HCl) and a gas was released. D: The precipitate in test tube Y was insoluble in hydrochloric acid.
   a. Use the above information to identify the solutions in each of the test tubes X, Y and Z.
   b. Write a chemical equation for the reaction that took place in test tube X before hydrochloric acid was added.

   (DoE Exemplar Paper 2 2007) Click here for the solution[14]

[9] http://www.fhsst.org/l31
[10] http://www.fhsst.org/l3r
[11] http://www.fhsst.org/lTl
[12] http://www.fhsst.org/lTi
[13] http://www.fhsst.org
[14] http://www.fhsst.org/l3Y

Solutions to Exercises in Chapter 9

Solution to Exercise 9.1 (p. 135)

Step 1. The cation is: Ag^{+} and the anion is: NO_3^-

Step 2. Since we know both the anion and the cation that silver nitrate dissociates into we can write the following equation:

$$\mathrm{AgNO}_3(s) \rightarrow Ag^{+}(aq) + \mathrm{NO}_3^-(aq) \tag{9.34}$$

Chapter 10

Quantitative Aspects of Chemical Change[1]

10.1 Quantitative Aspects of Chemical Change

An equation for a chemical reaction can provide us with a lot of useful information. It tells us what the reactants and the products are in the reaction, and it also tells us the ratio in which the reactants combine to form products. Look at the equation below:

$Fe + S \rightarrow \text{FeS}$

In this reaction, every atom of iron (Fe) will react with a single atom of sulphur (S) to form one molecule of iron sulphide (FeS). However, what the equation doesn't tell us, is the **quantities** or the **amount** of each substance that is involved. You may for example be given a small sample of iron for the reaction. How will you know how many atoms of iron are in this sample? And how many atoms of sulphur will you need for the reaction to use up all the iron you have? Is there a way of knowing what mass of iron sulphide will be produced at the end of the reaction? These are all very important questions, especially when the reaction is an industrial one, where it is important to know the quantities of reactants that are needed, and the quantity of product that will be formed. This chapter will look at how to quantify the changes that take place in chemical reactions.

10.2 The Mole

Sometimes it is important to know exactly how many particles (e.g. atoms or molecules) are in a sample of a substance, or what quantity of a substance is needed for a chemical reaction to take place.

You will remember from Relative atomic mass (Section 3.7.2: Relative atomic mass) that the **relative atomic mass** of an element, describes the mass of an atom of that element relative to the mass of an atom of carbon-12. So the mass of an atom of carbon (relative atomic mass is $12\,u$) for example, is twelve times greater than the mass of an atom of hydrogen, which has a relative atomic mass of $1\,u$. How can this information be used to help us to know what mass of each element will be needed if we want to end up with the same number of *atoms* of carbon and hydrogen?

Let's say for example, that we have a sample of $12\,g$ carbon. What mass of *hydrogen* will contain the same number of atoms as $12\,g$ carbon? We know that each atom of carbon weighs twelve times more than an atom of hydrogen. Surely then, we will only need $1\,g$ of hydrogen for the number of atoms in the two samples to be the same? You will notice that the number of particles (in this case, *atoms*) in the two substances is the same when the ratio of their sample masses (12 g carbon: 1g hydrogen = 12:1) is the same as the ratio of their relative atomic masses (12 u: 1 u = 12:1).

[1] This content is available online at <http://cnx.org/content/m38155/1.4/>.

To take this a step further, if you were to weigh out samples of a number of elements so that the mass of the sample was the same as the relative atomic mass of that element, you would find that the number of particles in each sample is $6,022 \times 10^{23}$. These results are shown in Table 10.1 below for a number of different elements. So, $24,31\,g$ of magnesium (relative atomic mass = $24,31\,u$) for example, has the same number of atoms as $40,08\,g$ of calcium (relative atomic mass = $40,08\,u$).

Element	**Relative atomic mass (u)**	**Sample mass (g)**	**Atoms in sample**
Hydrogen (H)	1	1	$6,022 \times 10^{23}$
Carbon (C)	12	12	$6,022 \times 10^{22}$
Magnesium (Mg)	24.31	24.31	$6,022 \times 10^{23}$
Sulphur (S)	32.07	32.07	$6,022 \times 10^{23}$
Calcium (Ca)	40.08	40.08	$6,022 \times 10^{23}$

Table 10.1: Table showing the relationship between the sample mass, the relative atomic mass and the number of atoms in a sample, for a number of elements.

This result is so important that scientists decided to use a special unit of measurement to define this quantity: the **mole** or 'mol'. A **mole** is defined as being an amount of a substance which contains the same number of particles as there are atoms in 12 g of carbon. In the examples that were used earlier, $24,31\,g$ magnesium is *one mole* of magnesium, while $40,08\,g$ of calcium is *one mole* of calcium. A mole of any substance always contains the same number of particles.

Definition 10.1: Mole

The mole (abbreviation 'n') is the SI (Standard International) unit for 'amount of substance'. It is defined as an amount of substance that contains the same number of particles (atoms, molecules or other particle units) as there are atoms in $12\,g$ carbon.

In one mole of any substance, there are $6,022 \times 10^{23}$ particles.

Definition 10.2: Avogadro's number

The number of particles in a mole, equal to $6,022 \times 10^{23}$. It is also sometimes referred to as the number of atoms in $12\,g$ of carbon-12.

If we were to write out Avogadro's number then it would look like: 602 200 000 000 000 000 000 000. This is a very large number. If we had this number of cold drink cans, then we could cover the surface of the earth to a depth of over $300\,km$! If you could count atoms at a rate of 10 million per second, then it would take you 2 billion years to count the atoms in one mole!

We can build up to the idea of Avogadro's number. For example, if you have 12 eggs then you have a dozen eggs. After this number we get a gross of eggs, which is 144 eggs. Finally if we wanted one mole of eggs this would be $6,022 \times 10^{23}$. That is a lot of eggs!

NOTE: The original hypothesis that was proposed by Amadeo Avogadro was that *'equal volumes of gases, at the same temperature and pressure, contain the same number of molecules'*. His ideas were not accepted by the scientific community and it was only four years after his death, that his original hypothesis was accepted and that it became known as 'Avogadro's Law'. In honour of his contribution to science, the number of particles in one mole was named *Avogadro's number*.

10.2.1 Moles and mass

1. Complete the following table:

Element	Relative atomic mass (u)	Sample mass (g)	Number of moles in the sample
Hydrogen	1.01	1.01	
Magnesium	24.31	24.31	
Carbon	12.01	24.02	
Chlorine	35.45	70.9	
Nitrogen		42.08	

Table 10.2

Click here for the solution[2]

2. How many atoms are there in...
   a. 1 mole of a substance
   b. 2 moles of calcium
   c. 5 moles of phosphorus
   d. $24,31\,g$ of magnesium
   e. $24,02\,g$ of carbon

   Click here for the solution[3]

10.3 Molar Mass

Definition 10.3: Molar mass
Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or $g \cdot \mathrm{mol}^{-1}$.

Refer to Table 10.1. You will remember that when the mass, in grams, of an element is equal to its relative atomic mass, the sample contains one mole of that element. This mass is called the **molar mass** of that element.

You may sometimes see the molar mass written as M_m. We will use M in this book, but you should be aware of the alternate notation.

It is worth remembering the following: On the periodic table, the relative atomic mass that is shown can be interpreted in two ways.

1. The mass of a *single, average atom* of that element relative to the mass of an atom of carbon.
2. The mass of *one mole of the element*. This second use is the molar mass of the element.

Element	Relative atomic mass (u)	Molar mass ($g \cdot \mathrm{mol}^{-1}$)	Mass of one mole of the element (g)
Magnesium	24,31	24,31	24,31
Lithium	6,94	6,94	6,94

continued on next page

[2]http://www.fhsst.org/lgl
[3]http://www.fhsst.org/lgi

Oxygen	16	16	16
Nitrogen	14,01	14,01	14,01
Iron	55,85	55,85	55,85

Table 10.3: The relationship between relative atomic mass, molar mass and the mass of one mole for a number of elements.

Exercise 10.1: Calculating the number of moles from mass ***(Solution on p. 167.)***

Calculate the number of moles of iron (Fe) in a $11,7\,g$ sample.

Exercise 10.2: Calculating mass from moles ***(Solution on p. 167.)***

You have a sample that contains 5 moles of zinc.

1. What is the mass of the zinc in the sample?
2. How many atoms of zinc are in the sample?

10.3.1 Moles and molar mass

1. Give the molar mass of each of the following elements:
   a. hydrogen
   b. nitrogen
   c. bromine

   Click here for the solution[4]
2. Calculate the number of moles in each of the following samples:
   a. $21,62\,g$ of boron (B)
   b. $54,94\,g$ of manganese (Mn)
   c. $100,3\,g$ of mercury (Hg)
   d. $50\,g$ of barium (Ba)
   e. $40\,g$ of lead (Pb)

   Click here for the solution[5]

10.4 An equation to calculate moles and mass in chemical reactions

The calculations that have been used so far, can be made much simpler by using the following equation:

$$\mathbf{n}\,(\text{number of moles}) = \frac{\mathbf{m}\,(\text{mass of substance in } g)}{\mathbf{M}\left(\text{molar mass of substance in } g\cdot\text{mol}^{-1}\right)} \quad (10.1)$$

TIP: Remember that when you use the equation $n = \frac{m}{M}$, the mass is always in grams (g) and molar mass is in grams per mol ($g\cdot\text{mol}^{-1}$).

The equation can also be used to calculate mass and molar mass, using the following equations:

$$m = n \times M \quad (10.2)$$

[4]http://www.fhsst.org/lg3
[5]http://www.fhsst.org/lgO

and

$$M = \frac{m}{n} \tag{10.3}$$

The following diagram may help to remember the relationship between these three variables. You need to imagine that the horizontal line is like a 'division' sign and that the vertical line is like a 'multiplication' sign. So, for example, if you want to calculate 'M', then the remaining two letters in the triangle are 'm' and 'n' and 'm' is above 'n' with a division sign between them. In your calculation then, 'm' will be the numerator and 'n' will be the denominator.

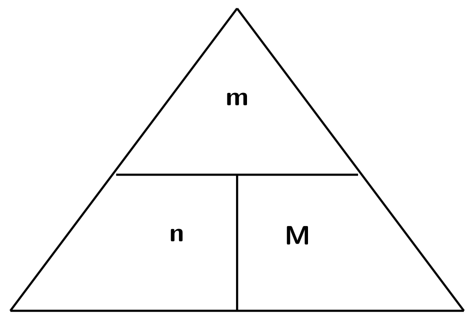


Figure 10.1

Exercise 10.3: Calculating moles from mass ***(Solution on p. 167.)***

Calculate the number of moles of copper there are in a sample that weighs $127\,g$.

Exercise 10.4: Calculating mass from moles ***(Solution on p. 167.)***

You are given a 5 mol sample of sodium. What mass of sodium is in the sample?

Exercise 10.5: Calculating atoms from mass ***(Solution on p. 167.)***

Calculate the number of atoms there are in a sample of aluminium that weighs $80,94\,g$.

10.4.1 Some simple calculations

1. Calculate the number of moles in each of the following samples:
   a. $5,6\,g$ of calcium
   b. $0,02\,g$ of manganese
   c. $40\,g$ of aluminium

   Click here for the solution[6]
2. A lead sinker has a mass of $5\,g$.
   a. Calculate the number of moles of lead the sinker contains.
   b. How many lead atoms are in the sinker?

   Click here for the solution[7]
3. Calculate the mass of each of the following samples:
   a. $2,5\,mol$ magnesium
   b. $12\,mol$ lithium

[6] http://www.fhsst.org/lgc
[7] http://www.fhsst.org/lga

c. $4,5 \times 10^{25}$ atoms of silicon

Click here for the solution[8]

10.5 Molecules and compounds

So far, we have only discussed moles, mass and molar mass in relation to *elements*. But what happens if we are dealing with a molecule or some other chemical compound? Do the same concepts and rules apply? The answer is 'yes'. However, you need to remember that all your calculations will apply to the *whole molecule*. So, when you calculate the molar mass of a molecule, you will need to add the molar mass of each atom in that compound. Also, the number of moles will also apply to the whole molecule. For example, if you have one mole of nitric acid (HNO_3), it means you have $6,022 \times 10^{23}$ **molecules** of nitric acid in the sample. This also means that there are $6,022 \times 10^{23}$ **atoms** of hydrogen, $6,022 \times 10^{23}$ **atoms** of nitrogen and $(3 \times 6,022 \times 10^{23})$ **atoms** of oxygen in the sample.

In a balanced chemical equation, the number that is written in front of the element or compound, shows the **mole ratio** in which the reactants combine to form a product. If there are no numbers in front of the element symbol, this means the number is '1'.

e.g. $N_2 + 3H_2 \rightarrow 2NH_3$

In this reaction, 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia.

Exercise 10.6: Calculating molar mass ***(Solution on p. 167.)***

Calculate the molar mass of H_2SO_4.

Exercise 10.7: Calculating moles from mass ***(Solution on p. 168.)***

Calculate the number of moles there are in $1\,kg$ of $MgCl_2$.

Exercise 10.8: Calculating the mass of reactants and products ***(Solution on p. 168.)***

Barium chloride and sulphuric acid react according to the following equation to produce barium sulphate and hydrochloric acid.

$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

If you have $2\,g$ of $BaCl_2$...

1. What quantity (in g) of H_2SO_4 will you need for the reaction so that all the barium chloride is used up?
2. What mass of HCl is produced during the reaction?

10.5.1 Group work : Understanding moles, molecules and Avogadro's number

Divide into groups of three and spend about 20 minutes answering the following questions together:

1. What are the units of the mole? Hint: Check the definition of the mole.
2. You have a $56\,g$ sample of iron sulphide (FeS)
   a. How many **moles** of FeS are there in the sample?
   b. How many **molecules** of FeS are there in the sample?
   c. What is the difference between a mole and a molecule?
3. The exact size of **Avogadro's number** is sometimes difficult to imagine.
   a. Write down Avogadro's number without using scientific notation.
   b. How long would it take to count to Avogadro's number? You can assume that you can count two numbers in each second.

[8] http://www.fhsst.org/lgx

Khan academy video on the mole - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/AsqEkF7hcII&rel=0>

Figure 10.2

10.5.2 More advanced calculations

1. Calculate the molar mass of the following chemical compounds:
   a. KOH
   b. $FeCl_3$
   c. $Mg(OH)_2$

   Click here for the solution[9]
2. How many moles are present in:
   a. $10\,g$ of Na_2SO_4
   b. $34\,g$ of $Ca(OH)_2$
   c. $2,45 \times 10^{23}$ molecules of CH_4?

   Click here for the solution[10]
3. For a sample of $0,2\,moles$ of potassium bromide (KBr), calculate...
   a. the number of moles of K^+ ions
   b. the number of moles of Br^- ions

   Click here for the solution[11]
4. You have a sample containing $3\,moles$ of calcium chloride.
   a. What is the chemical formula of calcium chloride?
   b. How many calcium atoms are in the sample?

   Click here for the solution[12]
5. Calculate the mass of:
   a. $3\,moles$ of NH_4OH
   b. $4,2\,moles$ of $Ca(NO_3)_2$

   Click here for the solution[13]
6. $96,2\,g$ sulphur reacts with an unknown quantity of zinc according to the following equation: $Zn + S \rightarrow ZnS$
   a. What mass of zinc will you need for the reaction, if all the sulphur is to be used up?
   b. What mass of zinc sulphide will this reaction produce?

   Click here for the solution[14]
7. Calcium chloride reacts with carbonic acid to produce calcium carbonate and hydrochloric acid according to the following equation: $CaCl_2 + H_2CO_3 \rightarrow CaCO_3 + 2HCl$ If you want to produce $10\,g$ of calcium carbonate through this chemical reaction, what quantity (in g) of calcium chloride will you

[9]http://www.fhsst.org/lgC
[10]http://www.fhsst.org/lgr
[11]http://www.fhsst.org/lg1
[12]http://www.fhsst.org/lgY
[13]http://www.fhsst.org/lgg
[14]http://www.fhsst.org/lg4

need at the start of the reaction?
Click here for the solution[15]

10.6 The Composition of Substances

The **empirical formula** of a chemical compound is a simple expression of the relative number of each type of atom in that compound. In contrast, the **molecular formula** of a chemical compound gives the actual number of atoms of each element found in a molecule of that compound.

Definition 10.4: Empirical formula
The empirical formula of a chemical compound gives the relative number of each type of atom in that compound.

Definition 10.5: Molecular formula
The molecular formula of a chemical compound gives the exact number of atoms of each element in one molecule of that compound.

The compound *ethanoic acid* for example, has the molecular formula CH_3COOH or simply $C_2H_4O_2$. In one molecule of this acid, there are two carbon atoms, four hydrogen atoms and two oxygen atoms. The ratio of atoms in the compound is 2:4:2, which can be simplified to 1:2:1. Therefore, the empirical formula for this compound is CH_2O. The empirical formula contains the smallest whole number ratio of the elements that make up a compound.

Knowing either the empirical or molecular formula of a compound, can help to determine its composition in more detail. The opposite is also true. Knowing the *composition* of a substance can help you to determine its formula. There are four different types of composition problems that you might come across:

1. Problems where you will be given the formula of the substance and asked to calculate the percentage by mass of each element in the substance.
2. Problems where you will be given the percentage composition and asked to calculate the formula.
3. Problems where you will be given the products of a chemical reaction and asked to calculate the formula of one of the reactants. These are often referred to as combustion analysis problems.
4. Problems where you will be asked to find number of moles of waters of crystallisation.

Exercise 10.9: Calculating the percentage by mass of elements in a compound ***(Solution on p. 168.)***
Calculate the percentage that each element contributes to the overall mass of sulphuric acid (H_2SO_4).

Exercise 10.10: Determining the empirical formula of a compound ***(Solution on p. 169.)***
A compound contains 52.2% carbon (C), 13.0% hydrogen (H) and 34.8% oxygen (O). Determine its empirical formula.

Exercise 10.11: Determining the formula of a compound ***(Solution on p. 169.)***
$207\,g$ of lead combines with oxygen to form $239\,g$ of a lead oxide. Use this information to work out the formula of the lead oxide (Relative atomic masses: $Pb = 207\,u$ and $O = 16\,u$).

Exercise 10.12: Empirical and molecular formula ***(Solution on p. 170.)***
Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: 39,9% carbon, 6,7% hyrogen and 53,4% oxygen.

1. Determine the empirical formula of acetic acid.
2. Determine the molecular formula of acetic acid if the molar mass of acetic acid is $60\,g \cdot \text{mol}^{-1}$.

[15] http://www.fhsst.org/lg2

Exercise 10.13: Waters of crystallisation *(Solution on p. 170.)*
Aluminium trichloride ($AlCl_3$) is an ionic substance that forms crystals in the solid phase. Water molecules may be trapped inside the crystal lattice. We represent this as: $AlCl_3 \cdot nH_2O$. A learner heated some aluminium trichloride crystals until all the water had evaporated and found that the mass after heating was $2,8\,g$. The mass before heating was $5\,g$. What is the number of moles of water molecules in the aluminium trichloride?

Khan academy video on molecular and empirical formulae - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/gfBcM3uvWfs&rel=0>

Figure 10.3

Khan academy video on mass composition - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/xatVrAh2U0E&rel=0>

Figure 10.4

10.6.1 Moles and empirical formulae

1. Calcium chloride is produced as the product of a chemical reaction.
   a. What is the formula of calcium chloride?
   b. What percentage does each of the elements contribute to the mass of a molecule of calcium chloride?
   c. If the sample contains $5\,g$ of calcium chloride, what is the mass of calcium in the sample?
   d. How many moles of calcium chloride are in the sample?

   Click here for the solution[16]
2. $13\,g$ of zinc combines with $6,4\,g$ of sulphur. What is the empirical formula of zinc sulphide?
   a. What mass of zinc sulphide will be produced?
   b. What percentage does each of the elements in zinc sulphide contribute to its mass?
   c. Determine the formula of zinc sulphide.

   Click here for the solution[17]
3. A calcium mineral consisted of 29,4% calcium, 23,5% sulphur and 47,1% oxygen by mass. Calculate the empirical formula of the mineral.
   Click here for the solution[18]
4. A chlorinated hydrocarbon compound was analysed and found to consist of 24,24% carbon, 4,04% hydrogen and 71,72% chlorine. From another experiment the molecular mass was found to be $99g \cdot$

[16] http://www.fhsst.org/lgT
[17] http://www.fhsst.org/lgb
[18] http://www.fhsst.org/lgj

mol^{-1}. Deduce the empirical and molecular formula.
Click here for the solution[19]

10.7 Molar Volumes of Gases

It is possible to calculate the volume of one mole of gas at STP using what we know about gases.

1. *Write down the ideal gas equation* pV = nRT, therefore $V = \frac{nRT}{p}$
2. *Record the values that you know, making sure that they are in SI units* You know that the gas is under STP conditions. These are as follows: $p = 101,3\,\text{kPa} = 101300\,Pa$ $n = 1\,\text{mol}$ $R = 8,31\,J \cdot K^{-1} \cdot \text{mol}^{-1}$ $T = 273\,K$
3. *Substitute these values into the original equation.*

$$V = \frac{nRT}{p} \tag{10.4}$$

$$V = \frac{1\text{mol} \times 8,31 J \cdot K^{-1} \cdot \text{mol}^{-1} \times 273K}{101300 Pa} \tag{10.5}$$

4. *Calculate the volume of* 1 *mole of gas under these conditions* The volume of 1 *mole* of gas at STP is $22,4 \times 10^{-3}\,m^3 = 22,4\text{dm}^3$.

TIP: The standard units used for this equation are P in Pa, V in m^3 and T in K. Remember also that $1\,000\,\text{cm}^3 = 1\text{dm}^3$ and $1\,000\,\text{dm}^3 = 1\,m^3$.

Exercise 10.14: Ideal Gas ***(Solution on p. 170.)***
A sample of gas occupies a volume of $20\,\text{dm}^3$, has a temperature of $200\,K$ and has a pressure of $105\,Pa$. Calculate the number of moles of gas that are present in the sample.

10.7.1 Using the combined gas law

1. An enclosed gas(i.e. one in a sealed container) has a volume of $300\,\text{cm}^3$ and a temperature of $300\,K$. The pressure of the gas is $50\,kPa$. Calculate the number of moles of gas that are present in the container.
   Click here for the solution[20]
2. What pressure will $3\,mol$ of gaseous nitrogen exert if it is pumped into a container that has a volume of $25\,\text{dm}^3$ at a temperature of $29\,^0\,C$?
   Click here for the solution[21]
3. The volume of air inside a tyre is 19 litres and the temperature is $290\,K$. You check the pressure of your tyres and find that the pressure is $190\,kPa$. How many moles of air are present in the tyre?
   Click here for the solution[22]
4. Compressed carbon dioxide is contained within a gas cylinder at a pressure of 700 kPa. The temperature of the gas in the cylinder is $310\,K$ and the number of moles of gas is $13\,mols$ of carbon dioxide. What is the volume of the gas inside the cylinder?
   Click here for the solution[23]

[19] http://www.fhsst.org/lgD
[20] http://www.fhsst.org/lgW
[21] http://www.fhsst.org/lgZ
[22] http://www.fhsst.org/lgB
[23] http://www.fhsst.org/lgK

10.8 Molar concentrations of liquids

A typical solution is made by dissolving some solid substance in a liquid. The amount of substance that is dissolved in a given volume of liquid is known as the **concentration** of the liquid. Mathematically, concentration (C) is defined as moles of solute (n) per unit volume (V) of solution.

$$C = \frac{n}{V} \tag{10.6}$$

For this equation, the units for volume are dm^3. Therefore, the unit of concentration is $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. When concentration is expressed in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$ it is known as the **molarity** (M) of the solution. Molarity is the most common expression for concentration.

TIP: Do not confuse molarity (M) with molar mass (M). Look carefully at the question in which the M appears to determine whether it is concentration or molar mass.

Definition 10.6: Concentration
Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. Another term that is used for concentration is **molarity (M)**

Exercise 10.15: Concentration Calculations 1 ***(Solution on p. 170.)***
If $3,5\,g$ of sodium hydroxide (NaOH) is dissolved in $2,5\,\mathrm{dm}^3$ of water, what is the concentration of the solution in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$?

Exercise 10.16: Concentration Calculations 2 ***(Solution on p. 171.)***
You have a $1\,\mathrm{dm}^3$ container in which to prepare a solution of potassium permanganate ($KMnO_4$). What mass of $KMnO_4$ is needed to make a solution with a concentration of $0,2\,M$?

Exercise 10.17: Concentration Calculations 3 ***(Solution on p. 171.)***
How much sodium chloride (in g) will one need to prepare $500\,\mathrm{cm}^3$ of solution with a concentration of $0,01\,M$?

10.8.1 Molarity and the concentration of solutions

1. $5,95\,g$ of potassium bromide was dissolved in $400\,\mathrm{cm}^3$ of water. Calculate its molarity.
   Click here for the solution[24]
2. $100\,g$ of sodium chloride (NaCl) is dissolved in $450\,\mathrm{cm}^3$ of water.
   a. How many moles of NaCl are present in solution?
   b. What is the volume of water (in dm^3)?
   c. Calculate the concentration of the solution.
   d. What mass of sodium chloride would need to be added for the concentration to become $5,7\,\mathrm{mol} \cdot \mathrm{dm}^{-3}$?

   Click here for the solution[25]
3. What is the molarity of the solution formed by dissolving $80\,g$ of sodium hydroxide ($NaOH$) in $500\,\mathrm{cm}^3$ of water?
   Click here for the solution[26]
4. What mass (g) of hydrogen chloride (HCl) is needed to make up $1000\,\mathrm{cm}^3$ of a solution of concentration $1\,\mathrm{mol} \cdot \mathrm{dm}^{-3}$?
   Click here for the solution[27]

[24] http://www.fhsst.org/lgk
[25] http://www.fhsst.org/lg0
[26] http://www.fhsst.org/lg8
[27] http://www.fhsst.org/lg9

5. How many moles of H_2SO_4 are there in $250\,cm^3$ of a $0,8\,M$ sulphuric acid solution? What mass of acid is in this solution?
Click here for the solution[28]

10.9 Stoichiometric calculations

Stoichiometry is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products. In representing chemical change showed how to write balanced chemical equations. By knowing the ratios of substances in a reaction, it is possible to use stoichiometry to calculate the amount of either reactants or products that are involved in the reaction. The examples shown below will make this concept clearer.

Exercise 10.18: Stoichiometric calculation 1 *(Solution on p. 171.)*
What volume of oxygen at S.T.P. is needed for the complete combustion of $2\,dm^3$ of propane (C_3H_8)? (Hint: CO_2 and H_2O are the products in this reaction (and in all combustion reactions))

Exercise 10.19: Stoichiometric calculation 2 *(Solution on p. 171.)*
What mass of iron (II) sulphide is formed when $5,6\,g$ of iron is completely reacted with sulphur?

When we are given a known mass of a reactant and are asked to work out how much product is formed, we are working out the theoretical yield of the reaction. In the laboratory chemists never get this amount of product. In each step of a reaction a small amount of product and reactants is 'lost' either because a reactant did not completely react or some of the product was left behind in the original container. Think about this. When you make your lunch or supper, you might be a bit hungry, so you eat some of the food that you are preparing. So instead of getting the full amount of food out (theoretical yield) that you started preparing, you lose some along the way.

Exercise 10.20: Industrial reaction to produce fertiliser *(Solution on p. 171.)*
Sulphuric acid (H_2SO_4) reacts with ammonia (NH_3) to produce the fertiliser ammonium sulphate ($(NH_4)_2SO_4$) according to the following equation:

$H_2SO_4\,(aq) + 2NH_3\,(g) \rightarrow (NH_4)_2SO_4\,(aq)$

What is the maximum mass of ammonium sulphate that can be obtained from $2,0\,kg$ of sulphuric acid?

Khan academy video on stoichiometry - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/jFv6k2OV7IU&rel=0>

Figure 10.5

10.9.1 Stoichiometry

1. Diborane, B_2H_6, was once considered for use as a rocket fuel. The combustion reaction for diborane is: $B_2H_6\,(g) + 3O_2\,(g) \rightarrow 2HBO_2\,(g) + 2H_2O\,(l)$ If we react $2,37\,grams$ of diborane, how many grams of water would we expect to produce?
Click here for the solution[29]

[28] http://www.fhsst.org/lgX
[29] http://www.fhsst.org/lgI

2. Sodium azide is a commonly used compound in airbags. When triggered, it has the following reaction: $2\mathrm{NaN}_3\,(s) \rightarrow 2Na\,(s) + 3N_2\,(g)$ If $23,4\,grams$ of sodium azide is used, how many moles of nitrogen gas would we expect to produce?
   Click here for the solution[30]
3. Photosynthesis is a chemical reaction that is vital to the existence of life on Earth. During photosynthesis, plants and bacteria convert carbon dioxide gas, liquid water, and light into glucose ($C_6H_{12}O_6$) and oxygen gas.
   a. Write down the equation for the photosynthesis reaction.
   b. Balance the equation.
   c. If $3\,moles$ of carbon dioxide are used up in the photosynthesis reaction, what mass of glucose will be produced?

   Click here for the solution[31]

This media object is a Flash object. Please view or download it at <http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=moles-100512054508-phpapp02&stripped_title=moles-4065434&userName=kwarne>

Figure 10.6

10.10 Summary

- It is important to be able to quantify the changes that take place during a chemical reaction.
- The **mole (n)** is a SI unit that is used to describe an amount of substance that contains the same number of particles as there are atoms in $12\,g$ of carbon.
- The number of particles in a mole is called the **Avogadro constant** and its value is $6,022 \times 10^{23}$. These particles could be atoms, molecules or other particle units, depending on the substance.
- The **molar mass (M)** is the mass of one mole of a substance and is measured in grams per mole or $g \cdot \mathrm{mol}^{-1}$. The numerical value of an element's molar mass is the same as its relative atomic mass. For a compound, the molar mass has the same numerical value as the molecular mass of that compound.
- The relationship between moles (n), mass in grams (m) and molar mass (M) is defined by the following equation:

$$n = \frac{m}{M} \tag{10.7}$$

- In a balanced chemical equation, the number in front of the chemical symbols describes the **mole ratio** of the reactants and products.
- The **empirical formula** of a compound is an expression of the relative number of each type of atom in the compound.
- The **molecular formula** of a compound describes the actual number of atoms of each element in a molecule of the compound.
- The formula of a substance can be used to calculate the **percentage by mass** that each element contributes to the compound.
- The **percentage composition** of a substance can be used to deduce its chemical formula.
- One mole of gas occupies a volume of $22,4\,\mathrm{dm}^3$.

[30]http://www.fhsst.org/lg5
[31]http://www.fhsst.org/lgN

- The **concentration** of a solution can be calculated using the following equation,

$$C = \frac{n}{V} \tag{10.8}$$

  where C is the concentration (in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$), n is the number of moles of solute dissolved in the solution and V is the volume of the solution (in dm^{-3}).
- **Molarity** is a measure of the concentration of a solution, and its units are $\mathrm{mol} \cdot \mathrm{dm}^{-3}$.
- **Stoichiometry** is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products.
- The theoretical yield of a reaction is the maximum amount of product that we expect to get out of a reaction

10.11 End of chapter exercises

1. Write only the word/term for each of the following descriptions:
   a. the mass of one mole of a substance
   b. the number of particles in one mole of a substance

   Click here for the solution[32]
2. Multiple choice: Choose the one correct answer from those given.
   a. $5\,g$ of magnesium chloride is formed as the product of a chemical reaction. Select the **true** statement from the answers below:
      a. 0.08 moles of magnesium chloride are formed in the reaction
      b. the number of atoms of Cl in the product is $0,6022 \times 10^{23}$
      c. the number of atoms of Mg is 0,05
      d. the atomic ratio of Mg atoms to Cl atoms in the product is 1:1

      Click here for the solution[33]
   b. 2 moles of oxygen gas react with hydrogen. What is the mass of oxygen in the reactants?
      a. 32 g
      b. 0,125 g
      c. 64 g
      d. 0,063 g

      Click here for the solution[34]
   c. In the compound potassium sulphate ($K_2\mathrm{SO}_4$), oxygen makes up x% of the mass of the compound. x = ...
      a. 36.8
      b. 9,2
      c. 4
      d. 18,3

      Click here for the solution[35]
   d. The molarity of a $150\,\mathrm{cm}^3$ solution, containing 5 g of $NaCl$ is...
      a. $0,09\,M$
      b. $5,7 \times 10^{-4}\,M$
      c. $0,57\,M$
      d. $0,03\,M$

[32]http://www.fhsst.org/lgR
[33]http://www.fhsst.org/lgn
[34]http://www.fhsst.org/lgQ
[35]http://www.fhsst.org/lgU

Click here for the solution[36]

3. Calculate the number of moles in:
   a. 5 g of methane (CH_4)
   b. 3,4 g of hydrochloric acid
   c. 6,2 g of potassium permanganate ($KMnO_4$)
   d. 4 g of neon
   e. 9,6 kg of titanium tetrachloride ($TiCl_4$)

   Click here for the solution[37]
4. Calculate the mass of:
   a. 0,2 mols of potassium hydroxide (KOH)
   b. 0,47 mols of nitrogen dioxide
   c. 5,2 mols of helium
   d. 0,05 mols of copper (II) chloride ($CuCl_2$)
   e. $31,31 \times 10^{23}$ molecules of carbon monoxide (CO)

   Click here for the solution[38]
5. Calculate the percentage that each element contributes to the overall mass of:
   a. Chloro-benzene (C_6H_5Cl)
   b. Lithium hydroxide (LiOH)

   Click here for the solution[39]
6. CFC's (chlorofluorocarbons) are one of the gases that contribute to the depletion of the ozone layer. A chemist analysed a CFC and found that it contained 58,64% chlorine, 31,43% fluorine and 9,93% carbon. What is the empirical formula?
   Click here for the solution[40]
7. 14 g of nitrogen combines with oxygen to form 46 g of a nitrogen oxide. Use this information to work out the formula of the oxide.
   Click here for the solution[41]
8. Iodine can exist as one of three oxides (I_2O_4; I_2O_5; I_4O_9). A chemist has produced one of these oxides and wishes to know which one they have. If he started with 508 g of iodine and formed 652 g of the oxide, which form has he produced?
   Click here for the solution[42]
9. A fluorinated hydrocarbon (a hydrocarbon is a chemical compound containing hydrogen and carbon.) was analysed and found to contain 8,57% H, 51,05% C and 40,38% F.
   a. What is its empirical formula?
   b. What is the molecular formula if the molar mass is $94,1 g \cdot \text{mol}^{-1}$?

   Click here for the solution[43]
10. Copper sulphate crystals often include water. A chemist is trying to determine the number of moles of water in the copper sulphate crystals. She weighs out 3 g of copper sulphate and heats this. After heating, she finds that the mass is 1,9 g. What is the number of moles of water in the crystals? (Copper sulphate is represented by $CuSO_4 \cdot xH_2O$).Click here for the solution[44]
11. $300\,\text{cm}^3$ of a $0,1\,\text{mol} \cdot \text{dm}^{-3}$ solution of sulphuric acid is added to $200\,\text{cm}^3$ of a $0,5\,\text{mol} \cdot \text{dm}^{-3}$ solution of sodium hydroxide.

[36]http://www.fhsst.org/lgy
[37]http://www.fhsst.org/lT3/
[38]http://www.fhsst.org/lTO
[39]http://www.fhsst.org/lTc
[40]http://www.fhsst.org/lTx
[41]http://www.fhsst.org/lTa
[42]http://www.fhsst.org/lTC
[43]http://www.fhsst.org/lT1
[44]http://www.fhsst.org/lTr

a. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.
b. Calculate the number of moles of sulphuric acid which were added to the sodium hydroxide solution.
c. Is the number of moles of sulphuric acid enough to fully neutralise the sodium hydroxide solution? Support your answer by showing all relevant calculations. (IEB Paper 2 2004)

Click here for the solution[45]

12. A learner is asked to make $200\,\mathrm{cm}^3$ of sodium hydroxide ($NaOH$) solution of concentration $0,5\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$.
    a. Determine the mass of sodium hydroxide pellets he needs to use to do this.
    b. Using an accurate balance the learner accurately measures the correct mass of the $NaOH$ pellets. To the pellets he now adds exactly $200\,\mathrm{cm}^3$ of pure water. Will his solution have the correct concentration? Explain your answer.
    c. The learner then takes $300\,\mathrm{cm}^3$ of a $0,1\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$ solution of sulphuric acid ($H_2\mathrm{SO}_4$) and adds it to $200\,\mathrm{cm}^3$ of a $0,5\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$ solution of $NaOH$ at 25^0C.
    d. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.
    e. Calculate the number of moles of $H_2\mathrm{SO}_4$ which were added to the NaOH solution.
    f. Is the number of moles of $H_2\mathrm{SO}_4$ calculated in the previous question enough to fully neutralise the $NaOH$ solution? Support your answer by showing all the relevant calculations. (IEB Paper 2, 2004)

    Click here for the solution[46]

[45] http://www.fhsst.org/lgP
[46] http://www.fhsst.org/lgm

Solutions to Exercises in Chapter 10

Solution to Exercise 10.1 (p. 154)

Step 1. If we look at the periodic table, we see that the molar mass of iron is $55,85\,g \cdot \text{mol}^{-1}$. This means that 1 mole of iron will have a mass of $55,85\,g$.

Step 2. If 1 mole of iron has a mass of $55,85\,g$, then: the number of moles of iron in $111,7\,g$ must be:

$$\frac{111,7g}{55,85g \cdot \text{mol}^{-1}} = 2\,\text{mol} \qquad (10.9)$$

There are 2 moles of iron in the sample.

Solution to Exercise 10.2 (p. 154)

Step 1. Molar mass of zinc is $65,38\,g \cdot \text{mol}^{-1}$, meaning that 1 mole of zinc has a mass of $65,38\,g$.

Step 2. If 1 mole of zinc has a mass of $65,38\,g$, then 5 moles of zinc has a mass of: $65,38\,g \times 5\,\text{mol} = 326,9\,g$ (answer to a)

Step 3.

$$5 \times 6,022 \times 10^{23} = 30,115 \times 10^{23} \qquad (10.10)$$

(answer to b)

Solution to Exercise 10.3 (p. 155)

Step 1.

$$n = \frac{m}{M} \qquad (10.11)$$

Step 2.

$$n = \frac{127}{63,55} = 2 \qquad (10.12)$$

There are 2 moles of copper in the sample.

Solution to Exercise 10.4 (p. 155)

Step 1.

$$m = n \times M \qquad (10.13)$$

Step 2. $M_{Na} = 22,99\,g \cdot \text{mol}^{-1}$
Therefore,

$$m = 5 \times 22,99 = 114,95\,g \qquad (10.14)$$

The sample of sodium has a mass of $114,95\,g$.

Solution to Exercise 10.5 (p. 155)

Step 1.

$$n = \frac{m}{M} = \frac{80,94}{26,98} = 3\,\text{moles} \qquad (10.15)$$

Step 2. Number of atoms in 3 mol aluminium $= 3 \times 6,022 \times 10^{23}$
There are $18,069 \times 10^{23}$ aluminium atoms in a sample of $80,94\,g$.

Solution to Exercise 10.6 (p. 156)

Step 1. Hydrogen $= 1,008\,g \cdot \text{mol}^{-1}$; Sulphur $= = 32,07\,g \cdot \text{mol}^{-1}$; Oxygen $= 16\,g \cdot \text{mol}^{-1}$

Step 2.

$$M_{(H_2SO_4)} = (2 \times 1,008) + (32,07) + (4 \times 16) = 98,09\,g \cdot \text{mol}^{-1} \qquad (10.16)$$

Solution to Exercise 10.7 (p. 156)

Step 1.

$$n = \frac{m}{M} \tag{10.17}$$

Step 2. a. Convert mass into grams

$$m = 1\text{kg} \times 1000 = 1000\,g \tag{10.18}$$

b. Calculate the molar mass of $MgCl_2$.

$$M_{(\text{MgCl}_2)} = 24,31 + (2 \times 35,45) = 95,21\,g \cdot \text{mol}^{-1} \tag{10.19}$$

Step 3.

$$n = \frac{1000}{95,21} = 10,5\,\text{mol} \tag{10.20}$$

There are $10,5$ moles of magnesium chloride in a $1\,kg$ sample.

Solution to Exercise 10.8 (p. 156)

Step 1.

$$n = \frac{m}{M} = \frac{2}{208,24} = 0,0096\,\text{mol} \tag{10.21}$$

Step 2. According to the balanced equation, 1 mole of $BaCl_2$ will react with 1 mole of H_2SO_4. Therefore, if $0,0096$ mol of $BaCl_2$ react, then there must be the same number of moles of H_2SO_4 that react because their mole ratio is 1:1.

Step 3.

$$m = n \times M = 0,0096 \times 98,086 = 0,94\,g \tag{10.22}$$

(answer to 1)

Step 4. According to the balanced equation, 2 moles of HCl are produced for every 1 mole of the two reactants. Therefore the number of moles of HCl produced is $(2 \times 0,0096)$, which equals $0,0096$ moles.

Step 5.

$$m = n \times M = 0,0192 \times 35,73 = 0,69\,g \tag{10.23}$$

(answer to 2)

Solution to Exercise 10.9 (p. 158)

Step 1. Hydrogen $= 1,008 \times 2 = 2,016\,u$
Sulphur $= 32,07\,u$
Oxygen $= 4 \times 16 = 64\,u$

Step 2. Use the calculations in the previous step to calculate the molecular mass of sulphuric acid.

$$\text{Mass} = 2,016 + 32,07 + 64 = 98,09\,u \tag{10.24}$$

Step 3. Use the equation:
Percentage by mass $= \frac{\text{atomic mass}}{\text{molecular mass of H}_2\text{SO}_4} \times 100\%$
Hydrogen

$$\frac{2,016}{98,09} \times 100\% = 2,06\% \tag{10.25}$$

Sulphur

$$\frac{32,07}{98,09} \times 100\% = 32,69\% \tag{10.26}$$

Oxygen

$$\frac{64}{98,09} \times 100\% = 65,25\% \tag{10.27}$$

(You should check at the end that these percentages add up to 100%!)
In other words, in one molecule of sulphuric acid, hydrogen makes up 2,06% of the mass of the compound, sulphur makes up 32,69% and oxygen makes up 65,25%.

Solution to Exercise 10.10 (p. 158)

Step 1. Carbon $= 52,2\,g$, hydrogen $= 13,0\,g$ and oxygen $= 34,8\,g$
Step 2.

$$n = \frac{m}{M} \tag{10.28}$$

Therefore,

$$n\left(\text{Carbon}\right) = \frac{52,2}{12,01} = 4,35\text{mol} \tag{10.29}$$

$$n\left(\text{Hydrogen}\right) = \frac{13,0}{1,008} = 12,90\text{mol} \tag{10.30}$$

$$n\left(\text{Oxygen}\right) = \frac{34,8}{16} = 2,18\text{mol} \tag{10.31}$$

Step 3. In this case, the smallest number of moles is 2.18. Therefore...
Carbon

$$\frac{4,35}{2,18} = 2 \tag{10.32}$$

Hydrogen

$$\frac{12,90}{2,18} = 6 \tag{10.33}$$

Oxygen

$$\frac{2,18}{2,18} = 1 \tag{10.34}$$

Therefore the empirical formula of this substance is: C_2H_6O. Do you recognise this compound?

Solution to Exercise 10.11 (p. 158)

Step 1.

$$239 - 207 = 32\,g \tag{10.35}$$

Step 2.

$$n = \frac{m}{M} \tag{10.36}$$

Lead

$$\frac{207}{207} = 1\text{ mol} \tag{10.37}$$

Oxygen

$$\frac{32}{16} = 2\text{ mol} \tag{10.38}$$

Step 1.

Step 3. The mole ratio of $Pb : O$ in the product is 1:2, which means that for every atom of lead, there will be two atoms of oxygen. The formula of the compound is PbO_2.

Solution to Exercise 10.12 (p. 158)

In $100\,g$ of acetic acid, there is $39,9\,g\,C$, $6,7\,g\,H$ and $53,4\,g\,O$

Step 2. $n = \frac{m}{M}$

$$\begin{aligned} n_C &= \tfrac{39,9}{12} = 3,33\,\mathrm{mol} \\ n_H &= \tfrac{6,7}{1} = 6,7\,mol \\ n_O &= \tfrac{53,4}{16} = 3,34\,mol \end{aligned} \quad (10.39)$$

Step 3. Empirical formula is CH_2O

Step 4. The molar mass of acetic acid using the empirical formula is $30\,g \cdot \mathrm{mol}^{-1}$. Therefore the actual number of moles of each element must be double what it is in the empirical formula.
The molecular formula is therefore $C_2H_4O_2$ or CH_3COOH

Solution to Exercise 10.13 (p. 159)

Step 1. We first need to find n, the number of water molecules that are present in the crystal. To do this we first note that the mass of water lost is $5 - 2,8 = 2,2$.

Step 2. The next step is to work out the mass ratio of aluminium trichloride to water and the mole ratio. The mass ratio is:

$$2,8 : 2,2 \quad (10.40)$$

To work out the mole ratio we divide the mass ratio by the molecular mass of each species:

$$\frac{2,8}{133} : \frac{2,2}{18} = 0,021 : 0,12 \quad (10.41)$$

Next we do the following:

$$0,021 \frac{1}{0,021} = 1 \quad (10.42)$$

and

$$\frac{0,12}{0,021} = 6 \quad (10.43)$$

So the mole ratio of aluminium trichloride to water is:

$$1 : 6 \quad (10.44)$$

Step 3. And now we know that there are 6 moles of water molecules in the crystal.

Solution to Exercise 10.14 (p. 160)

Step 1. The only value that is not in SI units is volume. $V = 0,02\,m^3$.

Step 2. We know that pV = nRT
Therefore,

$$n = \frac{pV}{RT} \quad (10.45)$$

Step 3.

$$n = \frac{105 \times 0,02}{8,31 \times 280} = \frac{2,1}{2326,8} = 0,0009\,\mathrm{moles} \quad (10.46)$$

Solution to Exercise 10.15 (p. 161)

Step 1.

$$n = \frac{m}{M} = \frac{3,5}{40} = 0,0875\,\text{mol} \tag{10.47}$$

Step 2.

$$C = \frac{n}{V} = \frac{0,0875}{2,5} = 0,035 \tag{10.48}$$

The concentration of the solution is $0,035\,\text{mol}\cdot\text{dm}^{-3}$ or $0,035\,M$

Solution to Exercise 10.16 (p. 161)

Step 1.

$$C = \frac{n}{V} \tag{10.49}$$

therefore

$$n = C \times V = 0,2 \times 1 = 0,2\,\text{mol} \tag{10.50}$$

Step 2.

$$m = n \times M = 0,2 \times 158,04 = 31,61\,g \tag{10.51}$$

The mass of KMnO_4 that is needed is $31,61\,g$.

Solution to Exercise 10.17 (p. 161)

Step 1.

$$V = \frac{500}{1000} = 0,5\,\text{dm}^3 \tag{10.52}$$

Step 2.

$$n = C \times V = 0,01 \times 0,5 = 0,005\,\text{mol} \tag{10.53}$$

Step 3.

$$m = n \times M = 0,005 \times 58,45 = 0,29\,g \tag{10.54}$$

The mass of sodium chloride needed is $0,29\,g$

Solution to Exercise 10.18 (p. 162)

Step 1. $C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)$
Step 2. From the balanced equation, the ratio of oxygen to propane in the reactants is 5:1.
Step 3. 1 volume of propane needs 5 volumes of oxygen, therefore $2\,\text{dm}^3$ of propane will need $10\,\text{dm}^3$ of oxygen for the reaction to proceed to completion.

Solution to Exercise 10.19 (p. 162)

Step 1. $Fe\,(s) + S\,(s) \rightarrow \text{FeS}\,(s)$
Step 2.

$$n = \frac{m}{M} = \frac{5,6}{55,85} = 0,1\text{mol} \tag{10.55}$$

Step 3. From the equation $1\,mole$ of Fe gives $1\,mole$ of FeS. Therefore, $0,1\,moles$ of iron in the reactants will give $0,1\,moles$ of iron sulphide in the product.
Step 4.

$$m = n \times M = 0,1 \times 87,911 = 8,79\,g \tag{10.56}$$

The mass of iron (II) sulphide that is produced during this reaction is $8,79\,g$.

Solution to Exercise 10.20 (p. 162)

Step 1.

$$n\left(H_2\mathrm{SO}_4\right) = \frac{m}{M} = \frac{2000\,g}{98,078\,g\cdot\mathrm{mols}^{-1}} = 20,39\,\mathrm{mols} \tag{10.57}$$

Step 2. From the balanced equation, the mole ratio of $H_2\mathrm{SO}_4$ in the reactants to $(\mathrm{NH}_4)_2\mathrm{SO}_4$ in the product is 1:1. Therefore, $20,39\,mols$ of $H_2\mathrm{SO}_4$ of $(\mathrm{NH}_4)_2\mathrm{SO}_4$.

The maximum mass of ammonium sulphate that can be produced is calculated as follows:

$$m = n \times M = 20,41\mathrm{mol} \times 132\,g\cdot\mathrm{mol}^{-1} = 2694\,g \tag{10.58}$$

The maximum amount of ammonium sulphate that can be produced is $2,694\,kg$.

Chapter 11

The Hydrosphere[1]

11.1 Introduction

As far as we know, the Earth we live on is the only planet that is able to support life. Amongst other factors, the Earth is just the right distance from the sun to have temperatures that are suitable for life to exist. Also, the Earth's atmosphere has exactly the right type of gases in the right amounts for life to survive. Our planet also has **water** on its surface, which is something very unique. In fact, Earth is often called the 'Blue Planet' because most of it is covered in water. This water is made up of *freshwater* in rivers and lakes, the *saltwater* of the oceans and estuaries, *groundwater* and *water vapour*. Together, all these water bodies are called the **hydrosphere**.

> NOTE: The total mass of the hydrosphere is approximately $1,4 \times 10^{18}$ tonnes! (The volume of one tonne of water is approximately 1 cubic metre.)

11.2 Interactions of the hydrosphere

It is important to realise that the hydrosphere is not an isolated system, but rather interacts with other global systems, including the *atmosphere*, *lithosphere* and *biosphere*. These interactions are sometimes known collectively as the water cycle.

- *Atmosphere* When water is heated (e.g. by energy from the sun), it evaporates and forms water vapour. When water vapour cools again, it condenses to form liquid water which eventually returns to the surface by precipitation e.g. rain or snow. This cycle of water moving through the atmosphere and the energy changes that accompany it, is what drives weather patterns on earth.
- *Lithosphere* In the lithosphere (the ocean and continental crust at the Earth's surface), water is an important *weathering* agent, which means that it helps to break rock down into rock fragments and then soil. These fragments may then be transported by water to another place, where they are deposited. These two processes (weathering and the transporting of fragments) are collectively called *erosion*. Erosion helps to shape the earth's surface. For example, you can see this in rivers. In the upper streams, rocks are eroded and sediments are transported down the river and deposited on the wide flood plains lower down. On a bigger scale, river valleys in mountains have been carved out by the action of water, and cliffs and caves on rocky beach coastlines are also the result of weathering and erosion by water. The processes of weathering and erosion also increase the content of dissolved minerals in the water. These dissolved minerals are important for the plants and animals that live in the water.
- *Biosphere* In the biosphere, land plants absorb water through their roots and then transport this through their vascular (transport) system to stems and leaves. This water is needed in *photosynthesis*,

[1] This content is available online at <http://cnx.org/content/m38138/1.4/>.

the food production process in plants. Transpiration (evaporation of water from the leaf surface) then returns water back to the atmosphere.

11.3 Exploring the Hydrosphere

The large amount of water on our planet is something quite unique. In fact, about 71% of the earth is covered by water. Of this, almost 97% is found in the oceans as saltwater, about 2.2% occurs as a solid in ice sheets, while the remaining amount (less than 1%) is available as freshwater. So from a human perspective, despite the vast amount of water on the planet, only a very small amount is actually available for human consumption (e.g. drinking water). In Reactions in aqueous solutions (Chapter 9) we looked at some of the reactions that occur in aqueous solution and saw some of the chemistry of water, in this section we are going to spend some time exploring a part of the hydrosphere in order to start appreciating what a complex and beautiful part of the world it is. After completing the following investigation, you should start to see just how important it is to know about the chemistry of water.

11.3.1 Investigation : Investigating the hydrosphere

1. For this exercise, you can choose any part of the hydrosphere that you would like to explore. This may be a rock pool, a lake, river, wetland or even just a small pond. The guidelines below will apply best to a river investigation, but you can ask similar questions and gather similar data in other areas. When choosing your study site, consider how accessible it is (how easy is it to get to?) and the problems you may experience (e.g. tides, rain).
2. Your teacher will provide you with the equipment you need to collect the following data. You should have at least one study site where you will collect data, but you might decide to have more if you want to compare your results in different areas. This works best in a river, where you can choose sites down its length.
   a. *Chemical data* Measure and record data such as temperature, pH, conductivity and dissolved oxygen at each of your sites. You may not know exactly what these measurements mean right now, but it will become clearer later.
   b. *Hydrological data* Measure the water velocity of the river and observe how the volume of water in the river changes as you move down its length. You can also collect a water sample in a clear bottle, hold it to the light and see whether the water is clear or whether it has particles in it.
   c. *Biological data* What types of animals and plants are found in or near this part of the hydrosphere? Are they specially adapted to their environment?

   Record your data in a table like the one shown below:

	Site 1	**Site 2**	**Site 3**
Temperature			
pH			
Conductivity			
Dissolved oxygen			
Animals and plants			

Table 11.1

3. **Interpreting the data** Once you have collected and recorded your data, think about the following questions:
   - How does the data you have collected vary at different sites?

- Can you explain these differences?
- What effect do you think *temperature*, *dissolved oxygen* and *pH* have on animals and plants that are living in the hydrosphere?
- Water is seldom 'pure'. It usually has lots of things dissolved (e.g. Mg^{2+}, Ca^{2+} and NO_3^- ions) or suspended (e.g. soil particles, debris) in it. Where do these substances come from?
- Are there any human activities near this part of the hydrosphere? What effect could these activities have on the hydrosphere?

11.4 The Importance of the Hydrosphere

It is so easy sometimes to take our hydrosphere for granted and we seldom take the time to really think about the role that this part of the planet plays in keeping us alive. Below are just some of the very important functions of water in the hydrosphere:

- *Water is a part of living cells* Each cell in a living organism is made up of almost 75% water, and this allows the cell to function normally. In fact, most of the chemical reactions that occur in life, involve substances that are dissolved in water. Without water, cells would not be able to carry out their normal functions and life could not exist.
- *Water provides a habitat* The hydrosphere provides an important place for many animals and plants to live. Many gases (e.g. CO_2, O_2), nutrients e.g. nitrate (NO_3^-), nitrite (NO_2^-) and ammonium (NH_4^+) ions, as well as other ions (e.g. Mg^{2+} and Ca^{2+}) are dissolved in water. The presence of these substances is critical for life to exist in water.
- *Regulating climate* One of water's unique characteristics is its high *specific heat*. This means that water takes a long time to heat up and also a long time to cool down. This is important in helping to regulate temperatures on earth so that they stay within a range that is acceptable for life to exist. *Ocean currents* also help to disperse heat.
- *Human needs* Humans use water in a number of ways. Drinking water is obviously very important, but water is also used domestically (e.g. washing and cleaning) and in industry. Water can also be used to generate electricity through hydropower.

These are just a few of the very important functions that water plays on our planet. Many of the functions of water relate to its chemistry and to the way in which it is able to dissolve substances in it.

11.5 Threats to the Hydrosphere

It should be clear by now that the hydrosphere plays an extremely important role in the survival of life on Earth and that the unique properties of water allow various important chemical processes to take place which would otherwise not be possible. Unfortunately for us however, there are a number of factors that threaten our hydrosphere and most of these threats are because of human activities. We are going to focus on two of these issues: **overuse** and **pollution** and look at ways in which these problems can possibly be overcome.

1. **Pollution**
   Pollution of the hydrosphere is also a major problem. When we think of pollution, we sometimes only think of things like plastic, bottles, oil and so on. But any chemical that is present in the hydrosphere in an amount that is not what it should be is a pollutant. Animals and plants that live in the Earth's water bodies are specially adapted to surviving within a certain range of conditions. If these conditions are changed (e.g. through pollution), these organisms may not be able to survive. Pollution then, can affect entire aquatic ecosystems. The most common forms of pollution in the hydrosphere are *waste products* from humans and from industries, *nutrient pollution* e.g. fertiliser runoff which causes eutrophication (an excess of nutrients in the water leading to excessive plant growth) and toxic

trace elements such as aluminium, mercury and copper to name a few. Most of these elements come from mines or from industries.

2. **Overuse of water**
   We mentioned earlier that only a very small percentage of the hydrosphere's water is available as freshwater. However, despite this, humans continue to use more and more water to the point where water *consumption* is fast approaching the amount of water that is *available*. The situation is a serious one, particularly in countries such as South Africa which are naturally dry and where water resources are limited. It is estimated that between 2020 and 2040, water supplies in South Africa will no longer be able to meet the growing demand for water in this country. This is partly due to population growth, but also because of the increasing needs of industries as they expand and develop. For each of us, this should be a very scary thought. Try to imagine a day without water... difficult isn't it? Water is so much a part of our lives, that we are hardly aware of the huge part that it plays in our daily lives.

11.5.1 Discussion : Creative water conservation

As populations grow, so do the demands that are placed on dwindling water resources. While many people argue that building dams helps to solve this water-shortage problem, there is evidence that dams are only a temporary solution and that they often end up doing far more ecological damage than good. The only sustainable solution is to reduce the *demand* for water, so that water supplies are sufficient to meet this. The more important question then is how to do this.

Discussion:

Divide the class into groups, so that there are about five people in each. Each group is going to represent a different sector within society. Your teacher will tell you which sector you belong to from the following: Farming, industry, city management or civil society (i.e. you will represent the ordinary 'man on the street'). In your groups, discuss the following questions as they relate to the group of people you represent: (Remember to take notes during your discussions, and nominate a spokesperson to give feedback to the rest of the class on behalf of your group)

- What steps could be taken by your group to conserve water?
- Why do you think these steps are *not* being taken?
- What incentives do you think could be introduced to encourage this group to conserve water more efficiently?

11.5.2 Investigation: Building of dams

In the previous discussion, we mentioned that there is evidence that dams are only a temporary solution to the water crisis. In this investigation you will look at why dams are a potentially bad solution to the problem.

For this investigation you will choose a dam that has been built in your area, or an area close to you. Make a note of which rivers are in the area. Try to answer the following questions:

- If possible talk to people who have lived in the area for a long time and try to get their opinion on how life changed since the dam was built. If it is not possible to talk to people in the area, then look for relevant literature on the area.
- Try to find out if any environmental impact assessments (this is where people study the environmnent and see what effect the proposed project has on the environment) were done before the dam was built. Why do you think this is important? Why do you think companies do not do these assessments?
- Look at how the ecology has changed. What was the ecology of the river? What is the current ecology? Do you think it has changed in a good way or a bad way?

Write a report or give a presentation in class on your findings from this investigation. Critically examine your findings and draw your own conclusion as to whether or not dams are only a short term solution to the growing water crisis.

It is important to realise that our hydrosphere exists in a delicate balance with other systems and that disturbing this balance can have serious consequences for life on this planet.

11.5.3 Group Project : School Action Project

There is a lot that can be done within a school to save water. As a class, discuss what actions could be taken by your class to make people more aware of how important it is to conserve water. Also consider what ways your school can save water. If possible, try to put some of these ideas into action and see if they really do conserve water.

11.6 How pure is our water?

When you drink a glass of water you are not just drinking water, but many other substances that are dissolved into the water. Some of these come from the process of making the water safe for humans to drink, while others come from the environment. Even if you took water from a mountain stream (which is often considered pure and bottled for people to consume), the water would still have impurities in it. Water pollution increases the amount of impurities in the water and sometimes makes the water unsafe for drinking. In this section we will look at a few of the substances that make water impure and how we can make pure water. We will also look at the pH of water.

In Reactions in aqueous solutions (Chapter 9) we saw how compounds can dissolve in water. Most of these compounds (e.g. Na^{+}, Cl^{-}, Ca^{2+}, Mg^{2+}, etc.) are safe for humans to consume in the small amounts that are naturally present in water. It is only when the amounts of these ions rise above the safe levels that the water is considered to be polluted.

You may have noticed sometimes that when you pour a glass a water straight from the tap, it has a sharp smell. This smell is the same smell that you notice around swimming pools and is due to chlorine in the water. Chlorine is the most common compound added to water to make it safe for humans to use. Chlorine helps to remove bacteria and other biological contaminants in the water. Other methods to purify water include filtration (passing the water through a very fine mesh) and flocculation (a process of adding chemicals to the water to help remove small particles).

pH of water is also important. Water that is to basic (pH greater than 7) or to acidic (pH less than 7) may present problems when humans consume the water. If you have ever noticed after swimming that your eyes are red or your skin is itchy, then the pH of the swimming pool was probably to basic or to acidic. This shows you just how sensitive we are to the smallest changes in our environment. The pH of water depends on what ions are dissolved in the water. Adding chlorine to water often lowers the pH. You will learn more about pH in grade 11.

11.6.1 Experiment: Water purity

Aim:

To test the purity and pH of water samples
Apparatus:

pH test strips (you can find these at pet shops, they are used to test pH of fish tanks), microscope (or magnifying glass), filter paper, funnel, silver nitrate, concentrated nitric acid, barium chloride, acid, chlorine water (a solution of chlorine in water), carbon tetrachloride, some test-tubes or beakers, water samples from different sources (e.g. a river, a dam, the sea, tap water, etc.).

Method:

1. Look at each water sample and note if the water is clear or cloudy.
2. Examine each water sample under a microscope and note what you see.
3. Test the pH of each of the water samples.
4. Pour some of the water from each sample through filter paper.
5. Refer to Testing for common anions in solutions (Section 9.5.2: Testing for common anions in solution) for the details of common anion tests. Test for chloride, sulphate, carbonate, bromide and iodide in each of the water samples.

Results:

Write down what you saw when you just looked at the water samples. Write down what you saw when you looked at the water samples under a microscope. Where there any dissolved particles? Or other things in the water? Was there a difference in what you saw with just looking and with looking with a a microscope? Write down the pH of each water sample. Look at the filter paper from each sample. Is there sand or other particles on it? Which anions did you find in each sample?

Discussion:

Write a report on what you observed. Draw some conclusions on the purity of the water and how you can tell if water is pure or not.

Conclusion:

You should have seen that water is not pure, but rather has many substances dissolved in it.

11.6.2 Project: water purification

Prepare a presentation on how water is purified. This can take the form of a poster, or a presentation or a project. Things that you should look at are:

- Water for drinking (potable water)
- Distilled water and its uses
- Deionised water and its uses
- What methods are used to prepare water for various uses
- What regulations govern drinking water
- Why water needs to be purified
- How safe are the purification methods

11.7 Summary

- The **hydrosphere** includes all the water that is on Earth. Sources of water include freshwater (e.g. rivers, lakes), saltwater (e.g. oceans), groundwater (e.g. boreholes) and water vapour. Ice (e.g. glaciers) is also part of the hydrosphere.
- The hydrosphere interacts with other **global systems**, including the atmosphere, lithosphere and biosphere.
- The hydrosphere has a number of important **functions**. Water is a part of all living cells, it provides a habitat for many living organisms, it helps to regulate climate and it is used by humans for domestic, industrial and other use.
- Despite the importance of the hydrosphere, a number of factors threaten it. These include **overuse** of water, and **pollution**.
- Water is not pure, but has many substances dissolved in it.

11.8 End of chapter exercises

1. What is the hydrosphere? How does it interact with other global systems?
   Click here for the solution[2]
2. Why is the hydrosphere important?
   Click here for the solution[3]

[2]http://www.fhsst.org/lgp
[3]http://www.fhsst.org/lgd

Glossary

A

Acid rain

Acid rain refers to the deposition of acidic components in rain, snow and dew. Acid rain occurs when sulphur dioxide and nitrogen oxides are emitted into the atmosphere, undergo chemical transformations and are absorbed by water droplets in clouds. The droplets then fall to earth as rain, snow, mist, dry dust, hail, or sleet. This increases the acidity of the soil and affects the chemical balance of lakes and streams.

Atomic mass number (A)

The number of protons and neutrons in the nucleus of an atom

Atomic number (Z)

The number of protons in an atom

Atomic orbital

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

Avogadro's number

The number of particles in a mole, equal to $6,022 \times 10^{23}$. It is also sometimes referred to as the number of atoms in $12\,g$ of carbon-12.

B

Boiling point

The temperature at which a *liquid* changes its phase to become a *gas*. The process is called evaporation and the reverse process is called condensation

C

Chemical bond

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

Chemical change

The formation of new substances in a chemical reaction. One type of matter is changed into something different.

Compound

A compound is a group of two or more different atoms that are attracted to each other by relatively strong forces or bonds.

Compound

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

Concentration

Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in $\mathrm{mol \cdot dm^{-3}}$. Another term that is used for concentration is **molarity (M)**

Conductivity

Conductivity is a measure of a solution's ability to conduct an electric current.

Conductors and insulators

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

Conservation of energy principle

Energy cannot be created or destroyed. It can only be changed from one form to another.

Core electrons

All the electrons in an atom, excluding the valence electrons

Covalent bond

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

D **Density**

Density is a measure of the mass of a substance per unit volume.

Dissociation

Dissociation in chemistry and biochemistry is a general process in which ionic compounds separate or split into smaller molecules or ions, usually in a reversible manner.

E **Electrolyte**

An electrolyte is a substance that contains free ions and behaves as an electrically conductive medium. Because they generally consist of ions in solution, electrolytes are also known as ionic solutions.

Electron configuration

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

Element

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

Empirical formula

The empirical formula of a chemical compound gives the relative number of each type of atom in that compound.

Empirical formula

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

H **Heterogeneous mixture**

A heterogeneous mixture is one that is non-uniform and the different components of the mixture can be seen.

Homogeneous mixture

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

I **Intermolecular force**

A force between molecules, which holds them together.

Intramolecular force

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

Ion

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

Ion exchange reaction

A type of reaction where the positive ions exchange their respective negative ions due to a driving force.

Ionic bond

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

Isotope

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

M **Magnetism**

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

Melting point

The temperature at which a *solid* changes its phase or state to become a *liquid*. The process is called melting and the reverse process (change in phase from liquid to solid) is called **freezing**.

Metallic bond

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

Mixture

A **mixture** is a combination of two or more substances, where these substances are not bonded (or joined) to each other.

Model

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

Molar mass

Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or $g \cdot \mathrm{mol}^{-1}$.

Mole

The mole (abbreviation 'n') is the SI (Standard International) unit for 'amount of substance'. It is defined as an amount of substance that contains the same number of particles (atoms, molecules or other particle units) as there are atoms in $12\,g$ carbon.

Molecular formula

The molecular formula of a chemical compound gives the exact number of atoms of each element in one molecule of that compound.

Molecular formula

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

P **Physical change**

A change that can be seen or felt, but that doesn't involve the break up of the particles in the reaction. During a physical change, the *form* of matter may change, but not its *identity*. A change in temperature is an example of a physical change.

Precipitate

A precipitate is the solid that forms in a solution during a chemical reaction.

R **Relative atomic mass**

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

T **The law of conservation of mass**

The mass of a closed system of substances will remain constant, regardless of the processes acting inside the system. Matter can change form, but cannot be created or destroyed. For any chemical process in a closed system, the mass of the reactants must equal the mass of the products.

V **Valence electrons**

The electrons in the outer energy level of an atom

Valency

The number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.

Viscosity

Viscosity is a measure of how resistant a liquid is to flowing (in other words, how easy it is to pour the liquid from one container to another).

W **Water hardness**

Water hardness is a measure of the mineral content of water. Minerals are substances such as calcite, quartz and mica that occur naturally as a result of geological processes.

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